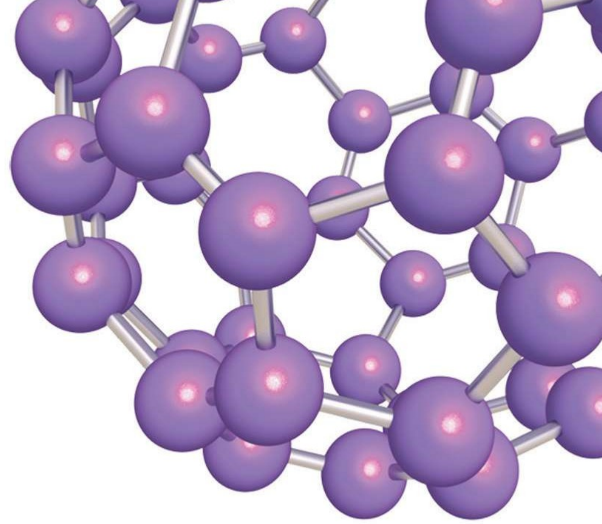
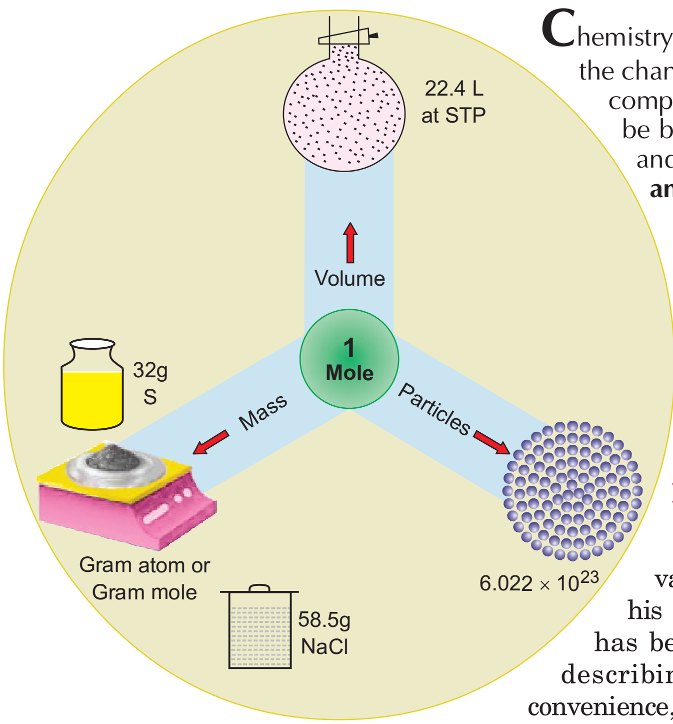
UNIT 1

SOME BASIC CONCEPTS OF CHEMISTRY



OBJECTIVES



Chemistry is the study of the materials that make up the universe and the changes which these materials undergo. The study deals with the composition, structure and properties of matter. These aspects can be best understood in terms of the constituents of matter: **atoms** and **molecules.** In fact, chemistry is called the **science of atoms and molecules**.

Can we see, weigh and perceive the atoms and molecules ?

Can we count the number of atoms and molecules in a given mass of matter ?

These are some curious questions which come to our mind. In the present unit, we will learn the answers of some of these questions.

IMPORTANCE OF CHEMISTRY

The material world around us consists of an enormous variety of substances. Man has been curious to know about his surroundings ever since he came into existence. Science has been helping the man to systematize his knowledge for describing and understanding the nature. For the sake of convenience, science is subdivided into various disciplines. The main disciplines of science are : chemistry, physics, biology, geology, etc.

Chemistry is the branch of science which deals with the study of composition, properties and structures of matter and the changes which the matter undergoes under different conditions and the laws which govern these changes.

Chemistry plays a central role in science and is often interlinked with other branches of science. It also plays an important role in our daily life.

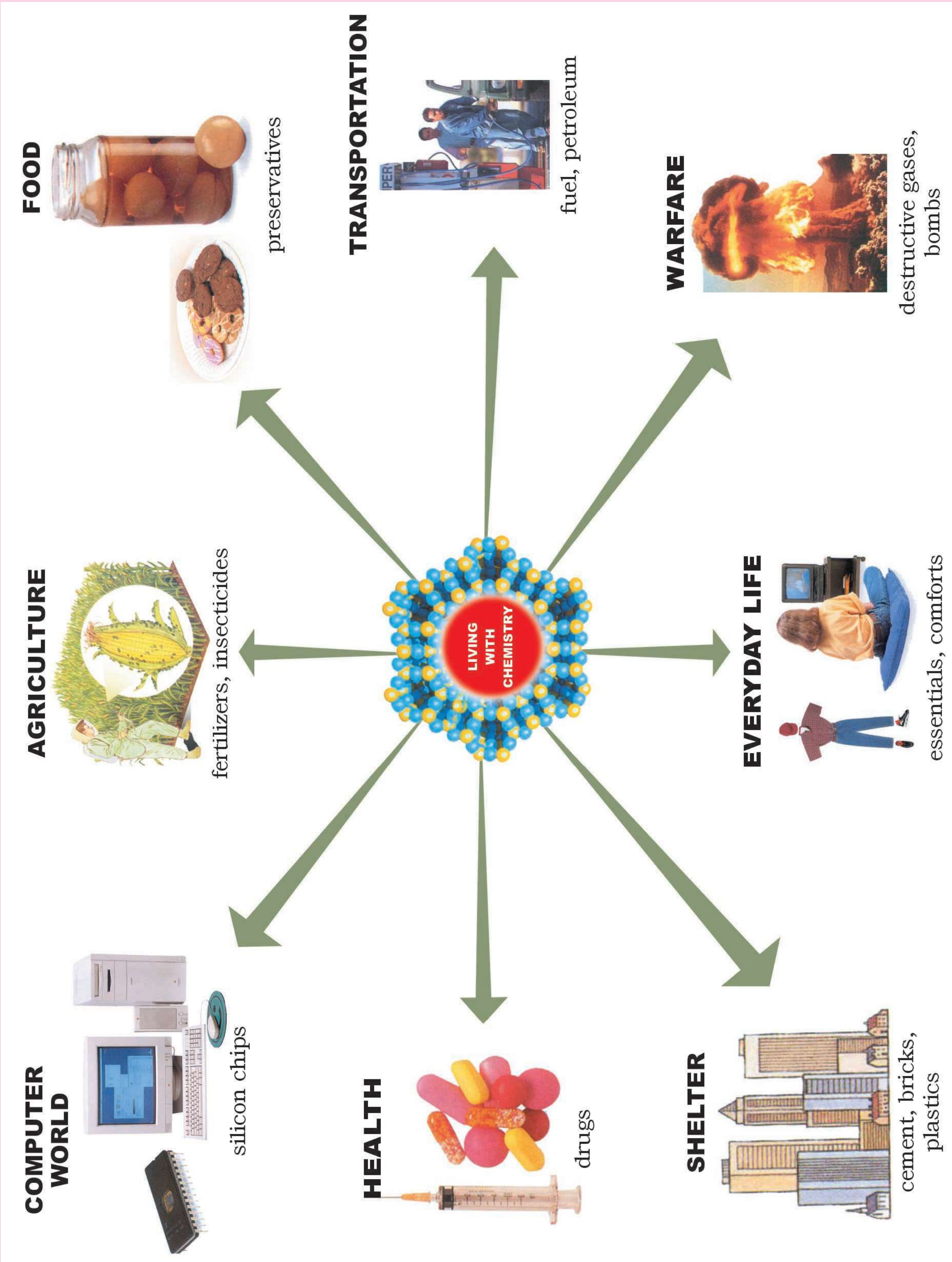


Fig. 1 Applications of chemistry in different fields.

Scope of Chemistry

In chemistry, we study the composition of materials to find out what they are made of. The marvellous thing about chemistry is that all the matter in the universe is made up of tiny smallest particles called **atoms** or **molecules.** It is very interesting that only about 118 types of atoms make up whole of the matter in the universe. We also learn how their composition affects their characteristics and behaviour so that we can plan to make new materials with properties of our interests. For this, we have to learn how substances undergo changes in composition and properties. These changes in the language of chemistry are called **chemical reactions.** Thus, studying chemistry is essential for us to understand better the world in which we live.

Since there is a large variety of substances in the universe, the scope of chemistry is immense. Whether we are concerned with living systems in biology and medicine, with materials such as iron, steel and concrete as in engineering or with manufacture of computer chips, we deal with chemistry. Progress in modern society is completely based on advances in chemistry. In today's technological age, the importance of chemistry is increasing. Chemical principles are playing key role in diverse areas such as weather patterns, functioning of brain and operation of computers.

It is difficult today to imagine our life without chemistry.

Main Applications of Chemistry

Chemistry has helped in agriculture, food, medicine, warfare, transportation, computers and in our everyday life. The important applications of chemistry in different fields are shown in Fig. 1. and are discussed below:

Chemistry for meeting our basic needs and necessities of life.

Chemistry has helped significantly in meeting human needs for food, health care products and other necessities of life.

- Chemistry has provided ***chemical fertilizers,*** improved varieties of ***insecticides, fungicides*** and ***pesticides*** to increase the yield of crops and fruits.
- The use of ***preservatives*** has helped to keep food materials for longer periods.
- Chemistry has also helped for **better health and sanitation**. The epidemics such as cholera, small pox, plague have now become things of the past.
- The discovery of **anaesthetics** has made surgical operations more and more successful.
- Chemistry has also given us a variety of drugs such as **antipyretics** (to lower body temperature in high fever), **analgesics** (to relieve pain), **tranquilizers** (for treatment of stress and mental diseases), **antimicrobials** (to cure infections), **antimalarials** (to treat malaria) etc. **Antibiotics** such as penicillin, amoxycillin, streptomycin and ***broad spectrum antibiotics*** such as tetracycline, chloramphenicol, etc. have cured a variety of diseases due to harmful micro-organisms. ***Antifertility drugs*** have been world-wide used for birth control methods.

Chemistry has given us a large number of **life *saving drugs***.

> Life saving drugs like ***cisplatin*** and ***taxol*** for cancer therapy and **AZT** (azidothymidine) used for helping AIDS victims are latest contributions of chemistry in medicines.

Chemistry for our comforts, pleasure and luxuries

Chemistry has also pioneer contribution towards our comforts, pleasure and luxuries.

- It has given us building materials, synthetic fibres and variety of articles of domestic use.
- Chemical industries manufacturing fertilizers, acids, alkalies, salts, dyes, polymers, drugs, soaps, detergents, metals, alloys and other organic and inorganic chemicals including new materials have improved our national economy.

Chemistry for giving us new materials

In recent years, chemistry has given us new materials such as super-conducting ceramics, conducting polymers, optical fibres, microalloys, carbon fibres, etc.

Chemistry in war

Chemistry has also increased the striking power of a country in war times. It is responsible for the discovery of highly explosive substances such as dynamite, TNT (trinitrotoluene), nitroglycerine, poisonous gases such as mustard gas, phosgene, lewisite and many deadly weapons such as atom bomb and hydrogen bomb.

Future Goals and Challenges for Chemistry

In recent years, chemistry has solved with a fair degree of success some of the challenging aspects of environmental degradation. So far alternatives to environmentally hazardous refrigerants like CFCs (chlorofluoro carbons), responsible for ozone depletion in the stratosphere has been successfully developed. Still many goals are there for the chemists to achieve successfully. Environmental problems, management of Green house gases (like methane, carbon dioxide, etc.) understanding bio-chemical processes, use of enzymes for large scale production of chemicals, synthesis of exotic materials are some main challenges for the future chemists. A developing country like India looks forward towards intelligent, talented and creative chemists for accepting these challenges.

NATURE OF MATTER

The entire universe is made up of matter and energy.

Anything that has mass and occupies space is called **matter.**

The presence of matter can be felt by one or more of our senses. Everything around us, for example, book, pen, table, sugar cubes, iron rod, water, air are composed of matter because they have mass and they occupy space.

Matter can be **classified** in two ways :

A : *Physical classification of matter*

B : *Chemical classification of matter.*

A : Physical Classification of Matter

Depending upon the physical state of matter, it can be classified into three states, namely, *solid, liquid* and *gaseous state.*

(*i*) Solid state. *A solid has a definite shape and definite volume.* Thus, solids are generally hard and rigid. For example, wood, table, copper rod, common salt, etc.

(*ii*) Liquid state. *A liquid has a definite volume but not definite shape.* A liquid takes the shape of the container in which it is placed. For example, water, milk, oil, etc.

(*iii*) Gaseous state. *A gas neither possesses a definite volume nor a definite shape.* It occupies the whole of the volume of the vessel in which it is placed. For example, air, oxygen, hydrogen, carbon dioxide, etc.

Particle Nature of Matter

As we know matter is made up of particles. In terms of particle concept of matter, in **solids**, the particles are closely packed and the empty spaces between the particles are very small. Due to close packing of the particles, they can only vibrate about their fixed positions i.e. *can have only vibratory motion.* Due to fixed positions of the particles in solids, they have highly ordered arrangement and this gives definite regular shape to the crystals. The regular ordered arrangement of particles in solids is called **lattice.**

The attractive forces among the particles called interparticle forces are strong in solids. Thus, **solids have a definite shape and a definite volume.**

In **liquids**, the particles are loosely packed and empty spaces between them are relatively large. Due to loose packing, the attractive forces between them are relatively weak. However, these are not weak enough to allow the particles to separate from one another. Thus, **liquids have definite volume but donot have fixed shape.** They take the shape of the container.

On the other hand, in **gases**, the particles are very loosely packed and empty spaces between them are very large. As a result, the attractive forces between the particles are very very small so that their movement is easy and fast. They can move to each corner of the vessel and therefore **gases donot have definite shape and definite volume.**

Because of such arrangement of particles, different states of matter exhibit different characteristics. The arrangement of particles in solids, liquids and gaseous state is shown in Fig. 2.

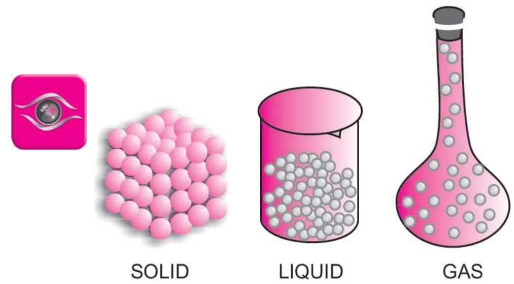


Fig. 2. Arrangement of particles in solid, liquid and gaseous state.

These three states of matter are interconvertible by changing the conditions of temperature and pressure. On heating, a solid changes to a liquid and on further heating the liquid changes to the gaseous (or vapour) state. In the reverse process, a gas on cooling liquefies to the liquid and liquid on further cooling freezes to the solid.

$$\text{SOLID} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{LIQUID} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{GAS}$$

B. Chemical Classification of Matter

The chemical classification of matter is based upon its composition. Different substances differ from each other in their constituents composition. On the basis of chemical composition, matter can be classified as :

A : Pure substances

B : Mixtures

A. Pure substances

Pure substances *consist of single type of particles.* All the constituent particles of a pure substance are same in their chemical nature. For example, copper, silver, gold, water, glucose, sodium chloride (common salt) are some examples of pure substances. As we know, water contains hydrogen and oxygen but they are always present in a fixed ratio and therefore, like all other pure substances behave as single substance having fixed chemical composition. Similarly, glucose contains carbon, hydrogen and oxygen in a fixed ratio and thus behaves as a pure substance. The constituents of pure substances cannot be separated by simple physical methods (like filtration, evaporation, distillation, sublimation, mechanical separation, etc.). These can only be separated by chemical or electrochemical methods.

Pure substances can be further classified into **elements** and **compounds**.

1. Element

An element is the simplest form of a pure substance. It may be defined as :

> ***the simplest form of a pure substance which can neither be decomposed into nor built from simpler substances by ordinary physical or chemical methods.***

The common examples of elements are hydrogen, oxygen, nitrogen, sulphur, iron, lead, gold, mercury, etc. There are about 118 elements known at present. Out of these, 92 have so far been found to occur naturally and the remaining have been prepared in the laboratory. All the elements do not occur in the crust of earth in equal proportions. About twenty elements make up 99% of the earth's crust. The most abundant elements in the earth's crust are oxygen, silicon, aluminium, iron, calcium, sodium, potassium, hydrogen, chlorine, carbon, etc.

Elements are further classified into following types :

(*a*) Metals. These elements are generally solids and possess characteristics such as *bright lustre, hardness and ability to conduct electricity and heat.* These are generally malleable (can be beaten into thin sheets) and ductile (can be drawn into wires). Some common examples of metals are copper, iron, silver, gold, aluminium, etc. About 80% of the known elements are metals.

(*b*) Non-metals. These elements are generally *non lustrous, brittle and poor conductors of heat and electricity.* The common examples of non-metals are carbon, hydrogen, oxygen, nitrogen, etc.

(*c*) Metalloids. *These are elements which have characteristics common to both metals and non-metals.* The common examples of metalloids are silicon, arsenic, bismuth, antimony, etc.

2. Compound

Compounds are pure substances containing more than one kind of element. A compound may be defined as :

> ***a pure substance containing two or more than two elements combined together in a definite proportion by mass and which can be decomposed into its constituent elements by suitable chemical methods.***

A compound always contains the same elements in *definite* or *fixed proportion* by weight. For example, water always contains hydrogen and oxygen in the ratio of 1 : 8 by mass.

Similarly, carbon dioxide always contains carbon and oxygen in the ratio of 3 : 8.

The properties of the compounds are totally different from the elements from which these are formed. For example, *hydrogen is combustible while oxygen is supporter of combustion but water (a compound of hydrogen and oxygen) is normally used for extinguishing fire*.

The compounds may be classified into two types as :

(*a*) Inorganic compounds. These are the compounds which are obtained from non-living sources such as rocks, minerals, etc. For example, common salt, marble, washing soda, etc.

(*b*) Organic compounds. These are the compounds which are obtained from living sources such as plants and animals. All these contain carbon. For example, carbohydrates, oils, fats, waxes, proteins, etc.

B. Mixtures

Mixtures contain more than one kind of pure form of matter, known as **substance.** It is a simple combination of two or more substances in which the constituent substances retain their identities. The composition of the mixture may be varied to any extent. The substances present in a mixture are called its **components**. Therefore, a mixture may contain two or more substances in any ratio. Thus, a mixture may be defined as :

> ***a combination of two or more elements or compounds in any proportion so that the components do not lose their identity.***

Many of the substances present around us are mixtures. For example, sugar solution in water, air, tea, brass (an alloy of copper and zinc), soft drink, soil, etc are all mixtures.

Mixtures are of two types :

(*a*) *Homogeneous mixture*
(*b*) *Heterogeneous mixture*

(*a*) Homogeneous mixture. *A mixture is said to be homogeneous if it has a uniform composition throughout.* The components of a mixture cannot be seen even under a powerful microscope. Homogeneous mixtures are also called **solutions.** Some examples of homogeneous mixtures are air, gasoline, sea water, brass, etc. Air is a homogeneous mixture of a number of gases such as oxygen, nitrogen, carbon dioxide, water vapour, etc. The composition of air is the same everywhere.

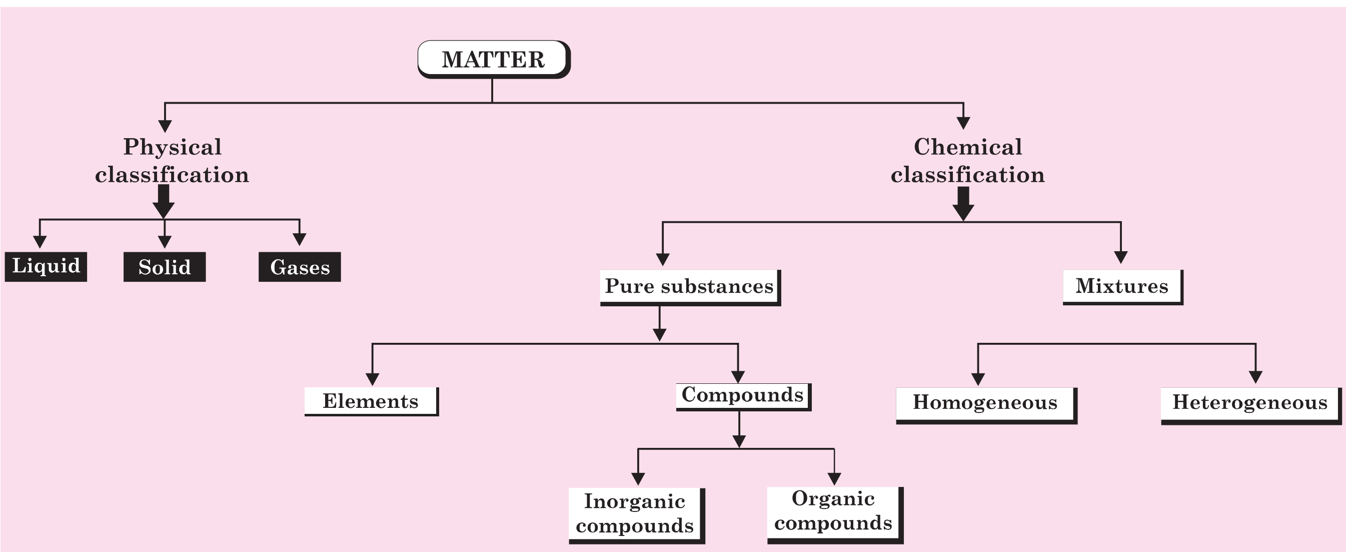
(*b*) Heterogeneous mixture. *A mixture is said to be heterogeneous if its composition is not uniform throughout.* These mixtures consist of two or more parts (called phases) which have different compositions. These have visible boundaries of separation between the different constituents and can be easily seen even with naked eye. For example, a mixture of iron filings, common salt and sulphur gives a heterogeneous mixture in which the components can be seen lying side by side very easily.

Differences between Compound and Mixture

Let us sum up important differences between compounds and mixtures.

Compounds	Mixtures
1. In a compound, two are more elements are combined chemically.	1. In a mixture, two or more elements or compounds just mix together.
2. The compounds contains two or more elements in a fixed ratio by mass. Its composition is always fixed.	2. The components of a mixture may be present in any ratio. Its composition is variable.
3. A compound has a definite formula.	3. A mixture does not have a definite formula.
4. A compound is always homogeneous i.e., has the same composition throughout.	4. A mixture may be homogeneous or heterogeneous.
5. A chemical reaction takes place and therefore, the formation of a compound takes place with absorption or evolution of energy.	5. No chemical reaction takes place and therefore, the formation of mixture is not accompained by any energy change.
6. The properties of a compound are entirely different from those of its constituents.	6. A mixture shows the properties of its constituents.
7. A compound cannot be separated into its constituents by ordinary physical methods. These can be separated by chemical or electrochemical reactions.	7. A mixture can be separated into its constituents by physical methods (like filtration, evaporation, distillation, sublimation, mechanical separation etc.)
8. A compound has a fixed melting point, boilling point, etc.	8. A mixture does not have fixed melting point, boiling point, etc.

The classification of matter has been summed up below :



PROPERTIES OF MATTER AND THEIR MEASUREMENT

Every substance has unique or characteristic properties. These properties may be classified as **physical properties** and **chemical properties**.

> **Physical properties** are those **properties which can be measured or observed without changing the identity or composition of the substance**.

For example, colour, odour, melting point, boiling point, density, etc. On the other hand,

> **chemical properties are those properties in which a chemical change in the substance takes place**.

The observation or measurement of chemical properties involve chemical reactions. The examples of chemical properties are characteristic reactions of different substances such as acidity, basicity, combustibility and reactions with other elements and compounds.

Many physical properties of matter such as mass, length, volume, temperature, pressure, density, etc. are quantitative in nature. Therefore, these properties are also called physical quantities.

Expressing a Physical Quantity

The value of a physical quantity is always expressed in two parts :

(*i*) *numerical value* and (*ii*) *unit*.

For example, we may express correctly the weight of a box as 5700 g or 5.7 kg. The simple numerical figure 5700 or 5.7 does not convey any meaningful information. Thus, it is essential to include units with every experimental value. A **unit** is defined as

the standard of reference chosen to measure any physical quantity.

Naturally, we must have a convenient system of units for assigning numerical values to the measured or calculated values. Fortunately, it turns out that it is sufficient to define **some basic units** such as units of **mass**, **length** and **time**. Since *these are independent units and cannot be derived from any other unit,* these are also called **fundamental units**.

The units for other quantities can be derived from these units and hence are called **derived units.**

There are many different systems of units. For the first time in 1791, a study committee of the **French Academy of Science** devised a system called *"the metric system"* which became popular in the scientific community throughout the world. The fundamental units of metric system are the **grams** for *mass,* the **metre** for *length* and the **litre** for the *volume***.** In India, the metric system was adopted in 1957. The metric system is a decimal system and the different units for a physical quantity are related in powers of ten. The different powers are generally indicated by a prefix attached to the unit.

SI Units

In recent years, the scientists have generally agreed to use the **International System of Units** abbreviated as SI units. These units were adopted by *General Conference of Weights and Measures in 1960.* These are abbreviated as **SI** and the designation comes from its French name **Systeme Internationale**.

The SI system has ***seven base units*** and are listed in Table 1. These base units pertain to seven fundamental scientific quantities and all other units can be derived from these.

Table. 1. Seven Base SI Units.

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

It may be noted that the unit of temperature according to SI system is Kelvin but still celsius scale (°C) of temperature is commonly used in our daily life. These two units are related as :

Kelvin temperature (K) = °C + 273.15

Similarly, **Angstrom** (Å) is commonly used as a unit of length in chemistry. It is almost of the same size as the size of an atom. It is equal to 10^{-10} m.

$$1\ \text{Å} = 10^{-10}\ \text{m}$$

However, it must be kept in mind that it is not SI unit. In SI units, we may use *nanometre (nm)* or *picometre (pm) :*

$$1\ \text{nm} = 10^{-9}\ \text{m}$$
$$1\ \text{pm} = 10^{-12}\ \text{m}$$

The definitions of the SI base units are given in Table 2.

Table 2. Definitions of seven SI base units.

Measured Quantity	Unit	Definition
Length	**metre**	The ***metre*** is the length of the path travelled by light in vacuum during a time interval of 1/299, 792,458 of a second.
Mass	**kilogram**	The ***kilogram*** is the unit of mass and is equal to the mass of the international prototype of the kilogram. It is also defined as the mass of a platinum block stored at the International Bureau of Weights and Measures in France.
Time	**second**	The ***second*** is the duration of 9, 192, 631, 770 periods of the radiation corresponding to the transition between the two hyperfine levels of the ground state of the cesium- 133 atom.
Electric current	**ampere**	The ***ampere*** is that constant current which, if maintained in two straight parallel conductors of infinite length, of negligible cross-section, and placed one metre apart in vacuum, would produce between these conductors a force equal to 2×10^{-7} newton per metre of length.
Thermodynamic temperature	**kelvin**	The ***kelvin*** is the unit of thermodynamic temperature and equals exactly 1/273.16 of the thermodynamic temperature of the triple point of water.
Amount of substance	**mole**	The ***mole*** is the amount of substance of a system which contains as many elementary entities as there are atoms in 0.012 kilogram of carbon–12.
Luminous intensity or Luminosity	**candela**	The ***candela*** is the luminous intensity, in a given direction, of a source that emits monochromatic radiation of frequency 540×10^{12} hertz and that has a radiant intensity in that direction of 1/683 watt per steradian.

PREFIXES

The SI units of some of the physical quantities are either too small or too large. To change the order of magnitude, these are expressed by using *prefixes* before the name of the base units. The various perfixes are listed in Table 3.

Table 3. List of common prefixes used in SI system.

Multiple	Prefix	Symbol	Multiple	Prefix	Symbol
10^{-1}	deci	d	10^{1}	deka	da
10^{-2}	centi	c	10^{2}	hecto	h
10^{-3}	milli	m	10^{3}	kilo	k
10^{-6}	micro	μ	10^{6}	mega	M
10^{-9}	nano	n	10^{9}	giga	G
10^{-12}	pico	p	10^{12}	tera	T
10^{-15}	femto	f	10^{15}	peta	P
10^{-18}	atto	a	10^{18}	exa	E
10^{-21}	zepto	z	10^{21}	zetta	Z
10^{-24}	yocto	y	10^{24}	yotta	Y

SOME COMMONLY USED QUANTITIES

Let us learn about some common quantities which we often come across.

1. Mass and weight.

Mass of a substance *is the amount of matter present in it.*

Weight is *the force exerted on an object by the pull of gravity.* The mass of a substance is constant and is independent of its location. For example, your body has same mass whether you are on Earth or on the moon. However, *weight of an object may change from one place to another due to change in gravity*. The mass of a substance can be measured very accurately in the laboratory by using an analytical balance. An analytical balance commonly used in laboratories is shown in Fig. 3 (*a*). These days, many types of electronic balances are available (Fig. 3 *b*). The analytical balances usually weigh to ± 0.0001 g, although some balances are available which can weigh to ± 0.00001 g.

The SI unit of mass is **kilogram**, (kg). Because kilogram is too large unit for many purposes, its fraction, gram, g, is commonly used and 1 kg = 1000 g. The smaller units such as milligram, mg (1mg = 10^{-3} g), microgram, μg (1 μg = 10^{-3} mg or = 10^{-6} g) are also commonly used.

$$1 \text{ kg} = 10^3 \text{ g} = 10^6 \text{ mg} = 10^9 \text{ μg}$$
$$1 \text{g} = 10^3 \text{ mg} = 10^6 \text{ μg}$$
$$1 \text{ mg} = 10^3 \text{ μg}$$

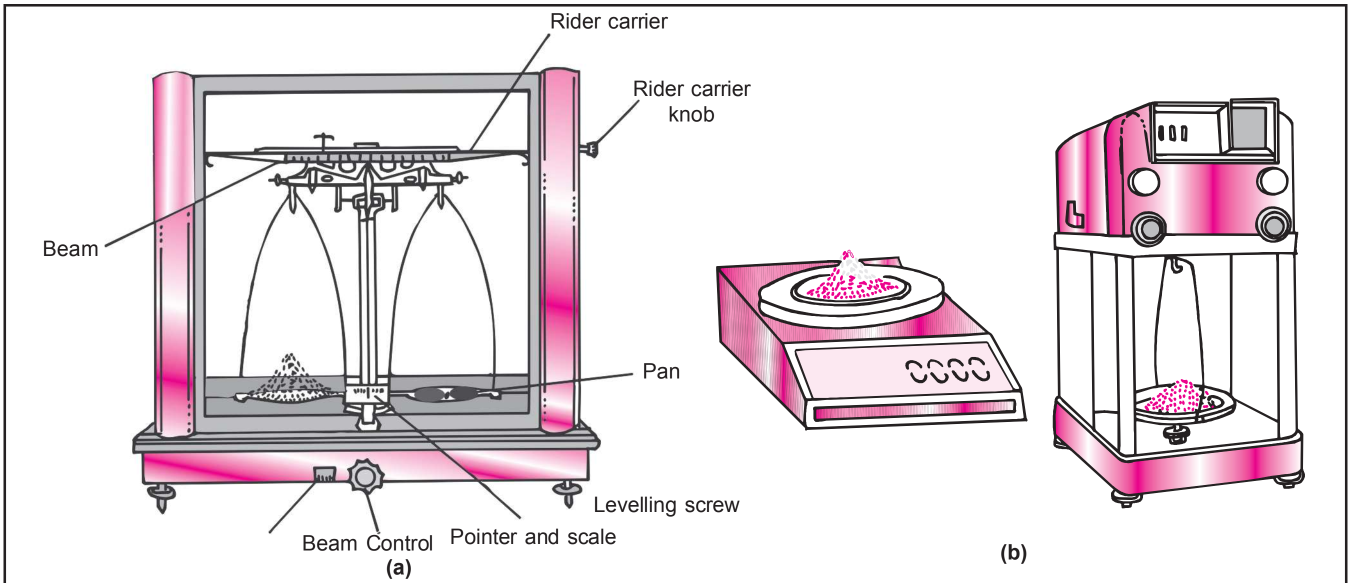


Fig. 3. (a) Analytical balance (b) Electronic balances.

2. Volume

Volume is the *amount of space occupied by an object.* It has the units of (length)^3. So in SI system, volume has units cubic metre, m^3 and is defined as ***the amount of space occupied by a cube having each edge of 1 metre.*** But in laboratories smaller volumes are commonly used. Hence, smaller units such as cubic decimetre, dm^3 (1dm^3 = 10^{-3} m^3) and cubic centimetre, cm^3 (1cm^3 = 0.001 dm^3 or = 10^{-6}m^3) are commonly used.

However, it may be noted that in normal use, volume of liquids are commonly expressed in more familiar units such as litre, L and millilitre, mL. Litre is equal in size to dm^3 and millilitre is equal in size to cm^3.

$$1L = 1dm^3 \text{ and } 1mL = 1\ cm^3$$
$$\text{and } 1L = 1000\ mL = 1000\ cm^3 = 1dm^3$$

These relationships can be understood from Fig. 4.

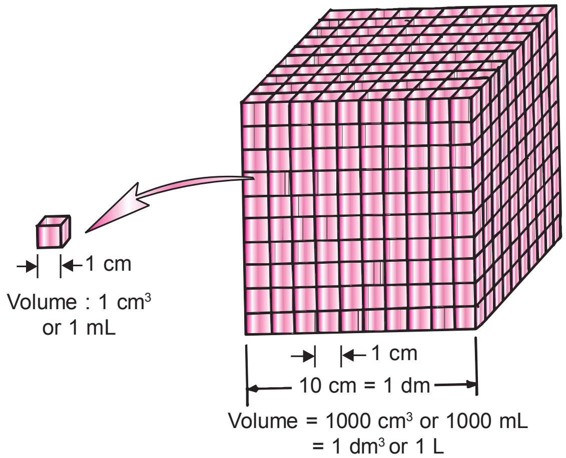


Fig. 4. Comparison between the relative sizes of two units of volume.

The volume of liquids or solutions are commonly measured in the laboratory by *graduated cylinder*, *burette* or *pipette*. A volumetric flask is used to prepare a known volume of a solution. Volumetric flasks of different volumes are available. These devices for measuring volume are shown in Fig. 5. A precise volume of a liquid can also be measured with the help of *graduated syringe*.

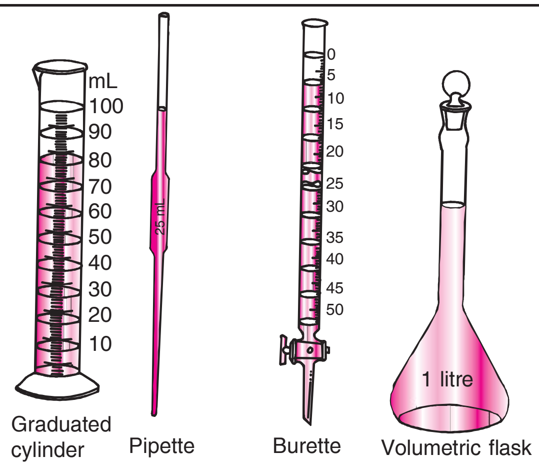


Fig. 5. Some volume measuring devices.

3. Measuring Density

Density of a substance is ***its amount of mass per unit volume***. So, density is the property which relates the mass of an object to its volume. It is simply the ***mass of an object divided by its volume.*** The units of density may be obtained as :

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

$$\text{SI unit of density} = \frac{\text{SI unit of mass}}{\text{SI unit of volume}}$$

$$= \frac{kg}{m^3} \text{ or } kg\ m^{-3}$$

This unit is quite large and it is commonly expressed as $g\ cm^{-3}$ for solids and $g\ mL^{-1}$ for liquids, where mass is expressed in gram and volume in cm^3 or mL. As we know that volume of a substance changes when heated or cooled, therefore, densities depend upon temperature. Hence, *while reporting a density, the temperature must also be mentioned.*

4. Temperature

Temperature *measures the degree of hotness or coldness of an object.* There are three common scales to measure temperature

1. **The SI scale or Kelvin scale measured in Kelvin (K)**
2. **Celsius scale measured in degree Celsius (°C)**
3. **Fahrenheit scale measured in degrees Fahrenheit (°F)**

The Fahrenheit scale has been commonly used in United States and is slowly been replaced by Celsius scale and Kelvin scale. The thermometers with celsius scale are calibrated from 0° to 100° where these two temperatures are the freezing point and boiling point of water respectively.

The Fahrenheit scale is represented between 32° to 212° which represent freezing point and boiling point of water respectively. There are 100 degrees between these two points on the celsius scale but 180° degrees between them in Fahrenheit giving rise to exact relationship as

$$180°\ F = 100°C \quad \text{or } 1.8°F = 1°C$$

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

$$°F = \frac{9}{5}(°C) + 32°$$

So to convert **Celsius to Fahrenheit**, we can use the formula

$$°F = \frac{9}{5}(°C) + 32°$$

and to convert **Fahrenheit to Celsius**, we can use the formula,

$$°C = \frac{5}{9}[°F - 32°]$$

According to SI system, the unit of temperature is **Kelvin.** One kelvin is the same size as one degree celsius. On Kelvin scale, the lowest possible temperature is equal to 0 Kelvin or 0 K. It may be noted that **kelvin temperature is written without a degree sign.** On this scale, the freezing point of water is – 273.15 K and boiling point of water is 373.15 K. The coldest possible temperature 0 K becomes –273.15°C and it is also sometimes called **absolute zero.**

The Kelvin scale is related to Celsius scale as :

K = °C + 273.15

or K = °C + 273 (for simplicity)

so that

Temperature in Kelvin = Temperature in °C + 273.15
Temperature in °C = Temperature in K – 273.15

For example, the human body temperature and room temperature on these temperature scales are :

Human body temperature = 98.6°F = 37°C = 310.15 K

Room temperature = 77°F = 25°C = 298.15K

The three types of thermometers using different scales are shown in Fig. 6.

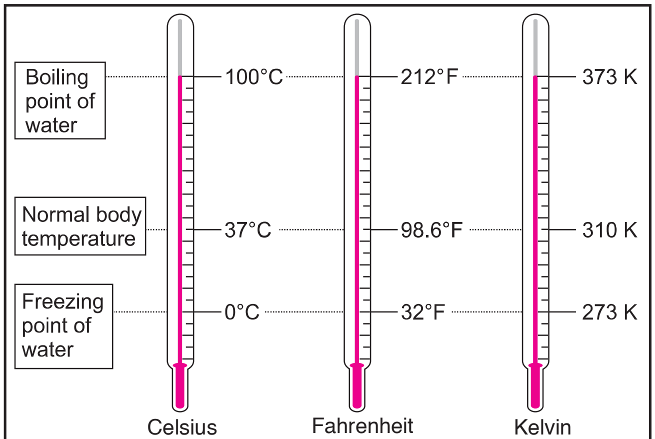


Fig. 6. Thermometers using different temperature scales.

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

DERIVED UNITS

The units of different physical quantities can be derived from the seven base units. These are called **derived units** because these are derived from the base units. For deriving these units, we can multiply or divide the symbols for units as if they are algebraic quantities. For example, to determine the volume of a box, we multiply length, breadth and height as :

Volume = Length × Breadth × Height

If units of length are m,

Volume = m × m × m = m^3

Therefore, **the SI units of volume are m^3**. If length, breadth and height of the box are 1.0 m, 2.0 m and 3.0 m respectively, then volume in terms of SI units may be expressed as :

Volume = (1.0 m) × (2.0 m) × (3.0 m) = 6.0 m^3

The other units may also be built up in this way from the definition of the physical quantity. For example,

(*i*) **Area**. The units of area can be easily derived as :

$$\text{Area} = \text{Length} \times \text{Breadth} = \text{m} \times \text{m} = \text{m}^2$$

(Units of length are m)

(*ii*) **Density.** The density of a substance is mass per unit volume, *i.e.*

$$\text{Density} = \frac{\text{Mass of sample}}{\text{Volume of sample}} = \frac{\text{kg}}{\text{m}^3} = \mathbf{kg\ m^{-3}}$$

(*iii*) **Velocity.** It is the change in distance per unit time, *i.e.*

$$\text{Velocity} = \frac{\text{Distance}}{\text{Time}} = \frac{\text{m}}{\text{s}} = \mathbf{m\ s^{-1}}$$

(*iv*) **Acceleration.** It is change in velocity per unit time, *i.e.*

$$\text{Acceleration} = \frac{\text{Velocity}}{\text{Time}} = \frac{(\text{m s}^{-1})}{(\text{s})} = \mathbf{m\ s^{-2}}$$

(*v*) **Force.** It is mass multiplied by acceleration.

$$\text{Force} = \text{Mass} \times \text{Acceleration} = (\text{kg}) \times (\text{m s}^{-2}) = \mathbf{kg\ m\ s^{-2}}$$

(*vi*) **Pressure.** Pressure is force per unit area, *i.e.*

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

But force has the derived units as kg m s^{-2}.

$$\therefore \quad \text{Units of pressure} = \frac{\text{kg m s}^{-2}}{\text{m}^2} = \mathbf{kg\ m^{-1}\ s^{-2}}$$

Some common **derived units** are given in Table 4.

Table 4. Some common SI derived units.

Physical quantity	Relation with other basic quantities	Unit	Symbol and definition
Area	Length square	—	m^2
Volume	Length cube	—	m^3
Density	Mass per unit volume	—	Kg/m^3 or $kg\ m^{-3}$
Speed	Distance travelled per unit time	—	m/s or $m\ s^{-1}$
Acceleration	Speed change per unit time	—	m/s × 1/s or $m\ s^{-2}$
Force	Mass × acceleration	Newton*	$N = kg\ m\ s^{-2}$
Pressure	Force per unit area	Pascal	$Pa = kg\ m^{-1}\ s^{-2}$
Energy, work	Force × distance travelled	Joule**	$J = kg\ m\ s^{-2} \times m$ or $kg\ m^2\ s^{-2}$
Power	Energy/Time	Watt	$W = kg\ m^2\ s^{-3}$ or Js^{-1}
Electric charge or potential difference	Current × time	Coulomb	C = As
Electric resistance	Potential difference/current	Ohm	$\Omega = VA^{-1}$
Electrical conductance	Reciprocal of resistance	Ohm^{-1} or Siemen	$S = AV^{-1}$
Frequency	Cycles/sec	Hertz	$Hz = s^{-1}$

* Newton is defined as the force which gives an acceleration of 1 m s^{-2} to a mass of 1 kg; so that

$$f = ma = 1\ (1\ kg)\ (1\ m\ s^{-2}) = 1\ kg\ m\ s^{-2} = 1\ N$$

** Joule is the work done when a displacement of 1 metre takes place by a force of 1 newton so that

$$w = f \times d = (1N) \times (1\ m) = (1\ kg\ m\ s^{-2}) \times (1m) = 1\ kg\ m^2\ s^{-2} = 1\ J$$

KEY NOTES

SOME IMPORTANT NOTES ON THE USE OF SI UNITS

In order to avoid confusion in the SI units, it is useful to remember the following points :

- Unit combinations should be indicated by means of either a dot or leaving space in between, *e.g.* metre kelvin may be written as :

  metre Kelvin = m.K or m K

  It should not be written as mK which stands for milli kelvin.
- Words and symbols should not be used in mixed forms. Thus, it is not proper to use J per mole. It should be written either as joule per mole or J mol^{-1}. *If mathematical operations are shown, only symbols should be used e.g.,* J mol^{-1} and not joule mol^{-1}.
- Symbols of the units do not have a plural ending like 's'. For example, we should write 10 cm and not 10 cms.
- A unit written with a prefix and a power is a power for the complete unit *e.g.,* cm^3 means $(centimetre)^3$ and not centi $(metre)^3$.
- Exponents also operate on prefixes. *e.g.,*

  $1\ cm^2 = (10^{-2}\ m)^2 = 10^{-4}\ m^2$

  ($\because$ 1 cm = 10^{-2} m)

It is not equal to 10^{-2} m²

Similarly,

$1\ mm^3 = (10^{-3}\ m)^3 = 10^{-9}\ m^3$

($\because$ 1 mm = 10^{-3} m)

It is not equal to 10^{-3} m³

It may be noted that some of the units are named after the name of the scientist who did pioneering work in the related field. For example, the unit of energy is joule abbreviated as 'J' and the unit of electric charge is coulomb abbreviated as C. *The symbols of the units which bear the names of the scientists are generally written in capital letters.* For example, the unit 'joule' is adopted after the name James P. Joule and is denoted by 'J' (capital J).

MEASUREMENT AND SIGNIFICANT FIGURES

Chemistry is largely an experimental science and deals with things that can be measured. A number of common devices are used in laboratories to make simple measurements. For example, a meter rod measures *length*, balance measures *mass;* burette, pipette, graduated cylinder and volumetric flask measure *volume;* thermometer measures *temperature,* etc. Every scientific measurement is limited by the reliability of the measuring instrument and the skill of the observer. In each measurement, uncertainty should be reported carefully. If we repeat a particular measurement, we usually do not obtain the same result because each measurement is subject to experimental error. Therefore, different measured values vary slightly from one another. To express the results of different measurements two terms ; accuracy and precision are commonly used.

Accuracy and Precision

The term accuracy denotes the closeness of an experimental value or the mean value of a set of measurements to the true value. Thus,

accuracy is *a measure of the difference between the experimental value or the mean value of a set of measurements and the true value.*

If the mean value of different measurements is close to the true value, the measurement is said to be **accurate**. Smaller the difference between the mean value and the true value, the larger is the accuracy. Suppose a sample of a compound contains 20.30% of chloride. The chemical analysis of the sample by one method (say method A) gives mean value of 20.60% chloride while another method (say method B) gives mean value of 20.46% chloride. Then,

Difference between the true value and mean value in method A = 20.60 – 20.30 = 0.30%

Difference between the true value and mean value in method B = 20.46 – 20.30 = 0.16%

Therefore, the method B gives more accurate data than method A. Thus, **accuracy also expresses the correctness of measurement.**

However, it may be noted that quite often, the true value of a quantity is not known. Therefore, it becomes very difficult to calculate the accuracy. In such cases, precision of the measurements is calculated.

Precision *refers to how closely two or more measurements of the same quantity agree with one another. Thus,*

precision is expressed as the difference between a measured value and the arithmetic mean value for a series of measurements.

In other words, precision gives the extent of agreement of the individual values among themselves *i.e.,* between the repeated measurements of the same quantity. *This means that smaller the difference between the individual values of repeated measurements, the greater is the precision.* For example, the analysis of an element gave the following results by two methods.

Method A : %Cl = 20.60 ± 0.05

Method B : %Cl = 20.46 ± 0.15

Thus, method A gives more precise data because the results differ only by ± 0.05%.

The terms accuracy and precision can be illustrated by the following example. Consider the results obtained by four students for the % of iron in a sample. These are given in Table 5 and are shown graphically in Fig. 7. The true value is 20.50 mg of iron.

Table 5. Results obtained by four students for the percentage of iron in a sample.

	Student A	Student B	Student C	Student D
	20.50	20.75	20.20	20.20
	20.48	20.65	20.25	20.40
	20.52	20.40	20.15	20.00
	20.50	20.28	20.20	20.60
Mean =	**20.50**	**20.52**	**20.20**	**20.30**
Deviation =	**± 0.02**	**± 0.24**	**± 0.05**	**± 0.30**

The results obtained by student A have good precision because the individual measurements are close together (± 0.02) and also have good accuracy because the mean result is same as the true value. The results of student B have low or poor precision because the measurements are scattered (± 0.24) but have good accuracy because the mean value (20.52) is close to the true value. Student C has obtained results of good precision because the measurements are close together (± 0.05) but have poor accuracy because the mean value (20.20) is far from the true value. The results of student D have poor precision because the measurements are scattered and have poor accuracy

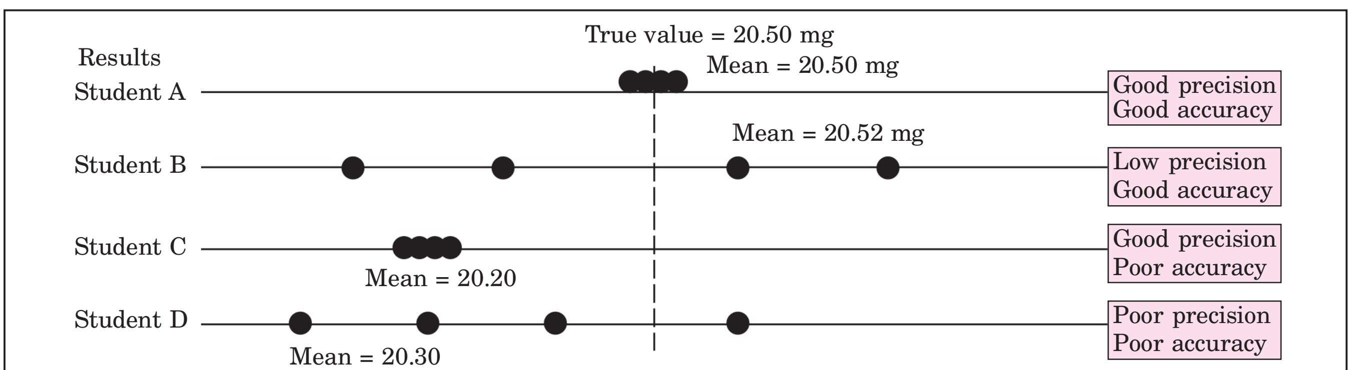


Fig. 7. Illustration of difference between accuracy and precision.

because the mean value (20.30) is far from the true value.

From the above observation, it can be concluded that accurate results are generally precise but precise results need not be accurate. In other words, ***good precision does not assure good accuracy.***

Exponential Number or Scientific Notation.

In Chemistry, we come across very large and very small numbers. It is very tedious to write down such numbers in the ordinary way. For example, it is not convenient to write Avogadro constant as 602, 213, 700,000,000,000,000,000. These numbers are usually expressed in a simple way known as *exponential form* or *scientific notation.* For example, the number 246.38 may be expressed as :

$$246.38 = 2.4638 \times 10 \times 10 \quad \text{or} \quad = 2.4638 \times 10^2$$

where 10^2 means 10×10 and 2 is the power or exponent to which 10 is raised.

In general, in scientific notation, a number may be expressed as :

$$N \times 10^n$$

where n is an exponent having positive or negative values and N is a single non-zero digit and lies between 1 to 10. N is called **digit term** and n is called **exponent**.

For example, 1487.2 may be written as:

$$1487.2 = 1.4872 \times 10 \times 10 \times 10 = 1.4872 \times 10^3$$

In a simple way, the decimal point is moved to the left until there is only one **non zero** digit before the decimal point. If the decimal point is moved x places, the exponent $n = x$. For example, in transforming 1487.2 to scientific notation, the decimal point is moved to left three places as :

$$1487.2$$

Here exponent, $n = 3$ and we can write

$$1487.2 = 1.4872 \times 10^3$$

Similarly, Avogadro number, may be expressed as :

$$602,213,700,000,000,000,000,000 = 6.022137 \times 10^{23}$$

Similarly, 0.00014872 may be written as:

$$0.00014872 = \frac{1.4872}{10 \times 10 \times 10 \times 10} = 1.4872 \times 10^{-4}$$

Here the exponent $n = -4$ means that the number 1.4872 has been divided by 10 four times. In a simple way, the decimal point is moved to the right until there is one non-zero digit before the decimal point. If the decimal point is moved y places, the exponent $n = -y$. For example, in transforming 0.00014872 to scientific notation, the decimal point is moved to right four places as :

$$0.0001\ 4872$$

so that exponent $n = -4$ and we can write

$$0.00014872 = 1.4872 \times 10^{-4}$$

Thus, in general, number is written in scientific notation as :

number with a single non-zero digit	number of places decimal point was moved
$N \times 10^n$	**+ moved left** **– moved right**

Arithmetics Using Scientific Notation, Addition and Subtraction

To add or subtract numbers in scientific notation, the exponent (n) must be the same in both numbers. If the exponent is not same in both numbers, it has to be made same before adding or subtracting. For example, suppose we want to add 4.236×10^4 and 3.582×10^3. We must first express both numbers so that they have the same exponent, n. So we can transform 3.582×10^3 to 0.3582×10^4 and then add

$$4.236 \times 10^4 + 0.3582 \times 10^4 = (4.236 + 0.3582) \times 10^4$$
$$= 4.5942 \times 10^4$$

Similarly, to subtract these numbers,

$$4.236 \times 10^4 - 3.582 \times 10^3$$
$$4.236 \times 10^4 - 0.3582 \times 10^4 = (4.236 - 0.3582) \times 10^4$$
$$= 3.8778 \times 10^4$$

KEY NOTE

Before adding or subtracting scientific numbers, it is important to make both numbers to the same power of 10.

Multiplication and Division

To multiply two numbers in scientific notation, we make use of the relation :

$$(10)^x \times (10)^y = 10^{(x+y)}$$

In other words, we add the exponents. For example,

$$(2.4 \times 10^3) \times (5.6 \times 10^5) = (2.4 \times 5.6) \times 10^{3+5}$$
$$= 13.44 \times 10^8$$
$$(3.025 \times 10^3) \times (6.217 \times 10^{-6}) = (3.025 \times 6.217) \times 10^{3+(-6)}$$
$$= 18.81 \times 10^{-3} = 1.881 \times 10^{-2}$$
$$(9.8 \times 10^{-2}) \times (2.5 \times 10^{-6}) = (9.8 \times 2.5) \times 10^{-2+(-6)}$$
$$= 24.50 \times 10^{-8} = 2.450 \times 10^{-7}$$

To divide two numbers in scientific notation, we make use of the relation :

$$\frac{10^x}{10^y} = 10^{x-y}$$

In other words, we subtract the power of 10 of the number in the denominator from the power of 10 of the number in the numerator. For example,

$$\frac{4.74\times10^{12}}{6.82\times10^{20}} = \left(\frac{4.74}{6.82}\right)\times10^{12-20}$$

$$= 0.695\times10^{-8} = 6.95\times10^{-9}$$

$$\frac{2.7\times10^{-3}}{5.5\times10^{4}} = \left(\frac{2.7}{5.5}\right)\times10^{-3-4}$$

$$= 0.4909\times10^{-7} = 4.909\times10^{-8}$$

Significant Figures

The uncertainty in the experimental or the calculated values is mainly due to the skill and accuracy of the observer and limitation of the measuring instrument as explained below:

(*i*) **Skill and accuracy of the observer.** Suppose three students A, B and C measure the volume of the liquid in a cylinder. They report the following values :

Result of A = 23.4 mL ; Result of B = 23.5 mL ;

Result of C = 23.6 mL

If the correct volume of the liquid is 23.5 mL, it means that the student B has measured the volume of the liquid *correctly* while the students A and C have made some error.

(*ii*) **Limitations of the measuring instrument.** Limitation of the measuring instrument is an important factor which leads to uncertainty in measurement. For example, suppose the mass of an object has been determined to be 9.265 g. If the accuracy of the analytical balance used is 0.001 g, this means that the actual mass of the object is 9.265 $\pm$ 0.001 g, i.e., it may be 9.264g or 9.266 g. Thus, in the reported mass, the first three digits (9, 2 and 6) are certain but the last digit(s) is uncertain and the uncertainty would be ±1 in the last digit. All measured quantities are reported in same way by writing the certain digits and the last uncertain digit. This is done in terms of *significant figures*.

The significant figures in a number are all the certain digits plus one uncertain digit.

For example, in the above value (9.265 g), there are four **significant figures**.

Thus, *the* ***number of significant figures conveys the information that except the last digit, all other digits are known with certainty.***

Rules for Determining the Number of Significant Figures

The following rules are followed to count the number of significant figures in a particular number :

(*i*) *All non-zero digits are significant.* For example,

3.132 has **four** significant figures

1.76 has **three** significant figures

5.4 has **two** significant figures

6.2316 has **five** significant figures.

The decimal place does not determine the number of significant figures.

(*ii*) *Zeros between two non zero digits* are significant. For example,

3.01 has **three** significant figures

2.005 has **four** significant figures

3.0023 has **five** significant figures

(*iii*) *The zeros preceding to the first non-zero number (i.e., to the left of the first non-zero number) are not significant.* Such zeros indicate the position of decimal point. For example,

0.324 has **three** significant figures

0.0052 has **two** significant figures

0.0003 has **one** significant figure

(*iv*) *All zeros placed at the end* or *to the right of a number are significant provided they are on the right side of the decimal point.* In fact, these represent the accuracy or the precision of the measuring scale. For example,

0.0200 has **three** significant figures

243.0 has **four** significant figures

243.00 has **five** significant figures

243.000 has **six** significant figures

KEY NOTE

- The results obtained by counting are **exact numbers** *i.e.,* numbers without any uncertainty. These exact numbers have infinite number of significant figures. For example, in 2 balls or 20 eggs, there are infinite number of significant figures because these are exact numbers and may be represented by writing infinite number of zeros after placing a decimal i.e. 2 = 2.000000 or 20 = 20.000000.
- Zeros at the end of a number or right of a number are significant provided they are on the right side of decimal point. For example, 0.300 has **three** significant figures. But zeros at the end of a number without decimal point are ambiguous. For example, 2500 may have two, three or four significant figures depending upon whether the uncertainty in measurements is 100, 10 or 1 respectively.

Exponential Numbers and Significant Figures

In exponential notations, the numerical portion (all digits) represents the number of significant figures. For example, a number, 0.000054 is expressed as 5.4×10^{-5} in terms of scientific notation. The number of significant figures in this number is **2**. Similarly, the number of significant figures in the Avogadro's number, 6.023×10^{23} is **four.**

Now, if we want to write five thousand with three significant figures, it can be written as

5.00×10^{3} (3 significant figures)

It cannot be written as 5000 because it has *four* and not three significant figures.

It must be **remembered** that the *decimal point does not count towards the number of significant figures.*

For example, consider the number 54023. It has five significant figures. We may write this number in different ways as:

54.023 cm
5.4023 decimetre
or 0.54023 metre

All these numbers have the same number of significant figures (*i.e.*, 5) regardless of the position of the decimal point. The different numbers simply represent different ways (or units) of expressing the measurement.

The above rules for calculating significant figures are illustrated below :

$\underbrace{0.000}_{\text{Not significant}}\overbrace{2040}^{\text{Significant}}$ **4 significant figures**

123.56 : 5 significant figures
5.407 : 4 significant figures (*zero is significant*)
900.0 : 4 significant figures (*all zeros are significant*)
0.0230 : 3 significant figures (*only last zero is significant*)

SOLVED EXAMPLES

Example 1.

State the number of significant figures in each of the following numbers :

(i) 207.35 (ii) 0.00368 (iii) 653
(iv) 3.653×10^4 (v) 0.378.

Solution:

(*i*) 207.35 has **five** significant figures.
(*ii*) 0.00368 has **three** significant figures. The three zeros in the beginning are not significant.
(*iii*) 653. It has **three** significant figures.
(*iv*) 3.653×10^4. It has **four** significant figures. The number has been expressed in scientific notation.
(*v*) 0.378 has **three** significant figures.

Example 2.

Express the following in the scientific notation :

(*i*) 0.0048 (*ii*) 234,000
(*iii*) 8008 (*iv*) 500.0
(*v*) 6.0012

N.C.E.R.T.

Solution:

(*i*) 0.0048 = 4.8×10^{-3}
(*ii*) 234,000 = 2.340×10^5
(*iii*) 8008 = 8.008×10^3
(*iv*) 500.0 = 5.000×10^2
(*v*) 6.0012 = 6.0012 or 6.0012×10^{-0}

Example 3.

Calculate the number of significant figures in the following values :

(*a*) *Planck's constant = 6.626×10^{-34} Js*
(*b*) *Avogadro number = 6.023×10^{23}*
(*c*) *Velocity of light = 3.0×10^{8} m s^{-1}*
(*d*) *Electronic charge = 1.602×10^{-19} C*

Solution:

(*a*) 6.626×10^{-34} J s = 4 significant figures
(*b*) 6.023×10^{23} = 4 significant figures
(*c*) 3.0×10^{8} m s^{-1} = 2 significant figures
(*d*) 1.602×10^{-19} C = 4 significant figures.

Example 4.

Calculate the number of significant figures in the following

(*i*) 0.0025 (*ii*) 208 (*iii*) 5005
(*iv*) 126,000 (*v*) 500.0 (*vi*) 2.0034

N.C.E.R.T.

Solution:

Number of significant figures :

(*i*) 0.0025 = 2
(*ii*) 208 = 3
(*iii*) 5005 = 4
(*iv*) 126000 = 3 (last three zeros are not significant)
(*v*) 500.0 = 4
(*vi*) 2.0034 = 5

Calculations Involving Significant Figures

The results of various measurements are to be added, subtracted, multiplied or divided, in most of the experiments to get desired results. But the different numbers do not have the same precision in measurement. In such cases, *the final result cannot be more precise than the least precise number involved in the calculations.* The number of the significant figures in such mathematical treatments are obtained by the following rules :

Rule 1. *When addition or subtraction is to be carried out, the final result should be reported upto the same number of decimal places as are present in the term having the least number of decimal places.*

(*a*) **Addition of numbers.** In the addition of numbers 7.23, 2.1 and 0.312, the number having the least decimal places is 2.1. Therefore, the final result after adding should also have digits only up to one decimal place.

7.23
2.1 ⟵ has only one decimal place
0.312
9.642 ⟵ answer should be reported upto one decimal place.

Correct answer = **9.6**

Some more examples of addition of numbers are :

89.232
+ 1.1
90.332

Since 1.1 has one decimal place, therefore, answer should be reported to one decimal place.

Correct answer = **90.3**

7.21 12.141 0.0028 19.3538 Correct answer = **19.35**	Since 7.21 has two decimal places, therefore, answer should be reported upto two decimal places.

(*b*) **Subtraction of numbers.** The subtraction of numbers is carried out in the same way as the addition. For example :

$$\begin{array}{r} 25.4630 \\ -24.21 \\ \hline 1.2530 \end{array} \longleftarrow \text{has two decimal places}$$

Correct answer = **1.25** (upto two decimal places)

Some other examples are :

$$\begin{array}{r} 18.4215 \\ -\ 6.0 \\ \hline 12.4215 \end{array} \qquad \begin{array}{r} 5.2748 \\ -5.2721 \\ \hline 0.0027 \end{array}$$

Since 6.0 has only one digit after decimal
Correct answer = **12.4**

Correct answer = **0.0027**

Rule 2. In multiplication or division, the final result should be reported up to the same number of significant figures as are present in the term with the least number of significant figures.

(*c*) **Multiplication of numbers.** If we multiply 5.1028 (having five significant figures) with 1.30 (having three significant figures), the value comes out to be 6.63364. But according to the rule, it must be reported with *three significant figures*. Therefore, the correct answer is 6.63.

5.1028 × 1.30 (3 significant figures)

= 6.63364

Correct answer = **6.63** (3 significant figures)

(*d*) **Division of numbers.** If we divide 5.2765 (having five significant figures) by 1.25 (having three significant figures), the result comes out to be 4.2212. But according to the rule, it must be reported with three significant figures. Thus, the correct answer is 4.22.

5.2765 ÷ 1.25

(3 significant figures)

= 4.2212

Correct answer = **4.22** (3 significant figures)

KEY NOTE

Calculations involving exact numbers

Exact numbers arise when we count items or **sometimes when we define a unit.** For example, when we say 10 pencils, we exactly mean 10, not 9.9 or 10.1. Also when we say that there are 12 inches in a foot, we mean exactly 12. **It may be noted that in calculations, the exact numbers will not have any effect on the number of significant figures.** The number of significant figures in a calculation result depends only on the numbers of significant figures in quantities having uncertainties. For example, we want to calculate the total mass of 8 pencils when each pencil has a mass of 3.0 g. The calculation is 3.0 × 8 = 24

Since 3.0 has two significant figures, the result should be 24 having two significant figures. The number 8 is exact and does not determine the number of significant figures.

It may be noted that **if a calculation involves a number of steps, the result should contain the same number of significant figures as that of the least precise number involved, other than the exact numbers.**

Retention of Significant Figures–Rounding off Figures

In solving problems, the final result often contains more digits than the number of significant figures. To retain the required number of significant figures, *rounding off* procedure is applied.

Rules for Rounding Off

(*i*) *If the digit coming after the desired number of significant figures happens to be more than 5, the preceding number is increased by 1.* For example, if we have to remove 6 in 1.386, we have to round it to 1.39.

(*ii*) *If the digit coming after the desired number of significant figures is less than 5, it is neglected as such i.e.,* the preceding number remains unchanged. For example, if we have to remove 4 in 6.234, we have to round it to 6.23.

(*iii*) *If the digit coming after the desired number of significant figures happens to be 5, then the preceding number is increased by 1 only in case it happens to be odd. In case of even number, the preceding number remains unchanged* For example, if 6.75 is to be rounded off by removing 5, we have to increase 7 to 8 giving 6.8 as the result. However, if 6.45 is to be rounded off, it is rounded off to 6.4.

Thus, to express the results to three significant figures

8.312 is rounded off to **8.31**
8.316 is rounded off to **8.32**
8.375 is rounded off to **8.38** [rule (*iii*)]
8.365 is rounded off to **8.36** [rule (*iii*)]

It must be **remembered** that if the problem involves more than one step, the 'rounding off' must be done only in the final answer. The intermediate steps of the calculations remain unchanged.

SOLVED EXAMPLES

Example 5.

Express the following numbers up to four significant figures :

(*i*) 5.607892 (*ii*) 32.392800
(*iii*) 0.007837 (*iv*) 1.78986×10^3
(*v*) 60000

Solution:

(*i*) 5.607892 = **5.608** (*ii*) 32.392800 = **32.39**
(*iii*) 0.007837 = **0.007837**
(*iv*) $1.78986 \times 10^3 = \mathbf{1.790 \times 10^3}$
(*v*) 6.000×10^4

Example 6.

Express the following up to three significant places :

(*a*) *the height of a man, 5 feet 9 inches in centimetres (1 inch = 2.54 cm)*
(*b*) *one millionth of one.*
(*c*) *four thousand*
(*d*) *decimal equivalent of 2/3*

Solution:

(*a*) 5 ft 9 inches = 69 inch = 69 × 2.54 = 175.26 cm. With three significant figures answer is **175 cm.**
(*b*) One millionth of one is $\frac{1}{10^6}$. It can be written with 3 significant figures as $\mathbf{1.00 \times 10^{-6}}$.
(*c*) four thousand = $\mathbf{4.00 \times 10^3}$
(*d*) $\frac{2}{3}$ = **0.667.**

Example 7.

Calculate to proper significant figures :

(*a*) 12.6×11.2 (*b*) 108/7.2

Solution:

(*a*) 12.6 × 11.2 = 141.12
Correct answer = **141** (upto 3 significant figures)
(*b*) $\frac{108}{7.2} = 15$
Correct answer = **15** (upto 2 significant figures as in 7.2 (leaving exact number).

Example 8.

How many significant figures should be present in the answer of the following calculations ?

(*i*) $\frac{0.02856 \times 298.15 \times 0.112}{0.5785}$
(*ii*) 5×5.364
(*iii*) 0.0125 + 0.7864 + 0.0215 **N.C.E.R.T.**

Solution:

(*i*) $\frac{0.02856 \times 298.15 \times 0.112}{0.5785}$ = 3 significant figures

The number of significant figures, in 0.112 (least number of significant figures) is 3, therefore, the result should have **3 significant figures.**

(*ii*) 5 × 5.364 = **4 significant figures**
5.364 has four significant figures (leaving the exact number 5)

(*iii*) 0.0125 + 0.7864 + 0.0215 = 4 significant figures
The result is to be reported upto four decimal places as :

$$\begin{array}{r} 0.0125 \\ 0.7864 \\ 0.0215 \\ \hline 0.8204 \end{array}$$

it has **4 significant figures.**

Example 9.

Express the results of the following calculations to the appropriate number of significant figures :

(*i*) $\frac{3.24 \times 0.08666}{5.006}$ (*ii*) $\frac{(1.36 \times 10^{-4})(0.5)}{2.6}$
(*iii*) 0.582 + 324.65 (*iv*) $2.64 \times 10^3 + 3.27 \times 10^2$
(*v*) $943 \times 0.00345 + 101.$

Solution:

(*i*) $\frac{3.24 \times 0.08666}{5.006} = 0.05608 = \mathbf{0.0561}$

In the calculation, the number of significant figures in the term 3.24 is 3, therefore, the result should have 3 significant figures. Therefore, the correct answer is **0.0561.** The number after 0 is 8 and therefore it is rounded off to 1.

(*ii*) $\frac{(1.36 \times 10^{-4})(0.5)}{2.6} = 0.2615 \times 10^{-4}$
$= \mathbf{0.3 \times 10^{-4}}$.

The answer should have one significant figure because 0.5 has one significant figure.

(*iii*)
$$\begin{array}{r} 0.582 \\ +\ 324.65 \\ \hline 325.232 \end{array}$$
(Second place of decimal)
= **325.23**

Since 324.65 contains digits up to second place of decimal, therefore, the answer must be reported up to second place.

(*iv*) $2.64 \times 10^3 + 3.27 \times 10^2$
or $2.64 \times 10^3 + 0.327 \times 10^3 = 2.967 \times 10^3 = \mathbf{2.97 \times 10^3}$

Since 2.64 has two digits after decimal place, the answer should be rounded off to two decimal places.

(*v*) The first two numbers are to be multiplied initially as follows :
943 × 0.00345 = 3.25335 = **3.25.**

According to the rules on multiplication of significant figures, the answer has to be reported up to three significant figures. It should be 3.25. Upon further addition, the final answer may be calculated as :

$$\begin{array}{r} 3.25 \\ +\ 101. \\ \hline 104.25 \end{array}$$
(zero place of decimal)
= **104**

∴ The answer should have zero place of decimal.

Example 10.

The mass of a piece of paper is 0.02 g and the mass of a solid substance and the piece of paper is 20.036 g. If the volume of the solid is 2.16 cm³, calculate the density of the substance up to proper number of significant digits.

Solution:

Mass of piece of paper = 0.02 g
Mass of solid substance and piece of paper = 20.036 g
∴ Mass of solid substance = 20.036 − 0.02 = 20.016
or = 20.02 (upto second decimal place).

Volume of solid $= 2.16\ cm^3$

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} = \frac{20.02}{2.16} = 9.268$$

Since 2.16 has three significant figures, the answer should also contain three significant figures.

∴ Density up to proper number of significant figures

$=$ **9.27 g/cm³.**

❑ Example 11.

Perform the following calculations and express the result to proper number of significant figures :

(i) $144.3\ m^2 + (2.54\ m \times 8.4\ m)$

(ii) $(4.05 \times 10^2\ mL) - (0.0225 \times 10^2\ mL)$

(iii) $(3.50 \times 10^2\ cm)\ (4.00 \times 10^6\ cm)$

Solution:

(*i*) $(144.3\ m^2) + (2.54\ m \times 8.4\ m)$

$2.54\ m \times 8.4\ m = 21.336\ m^2$ or $21\ m^2$ (upto 2 significant figures)

$$\begin{array}{r} 144.3\ m^2 \\ 21.0\ m^2 \\ \hline 165.3\ m^2 \end{array} \quad \text{or} \quad \mathbf{165\ m^2}$$

(*ii*) $(4.05 \times 10^2\ mL) - (0.0225 \times 10^2\ mL)$

$$\begin{array}{r} 4.05 \times 10^2\ mL \\ -\ 0.0225 \times 10^2\ mL \\ \hline 4.0275 \times 10^2\ mL \end{array} \quad \text{or} \quad \mathbf{4.03 \times 10^2\ mL}$$

(upto second decimal place as in 4.05)

(*iii*) $(3.50 \times 10^2\ cm) \times (4.00 \times 10^6\ cm)$

$= 14.0 \times 10^8\ cm^2$ (upto 3 significant figures)

Practice Problems

1. Calculate the number of significant figures in the following :

(*a*) 1.00×10^6 (*b*) 0.0050 (*c*) 1.234

(*d*) 0.0006 (*e*) 0.368

2. Express the following numbers to three significant figures :

(*i*) 6.023×10^{23} (*ii*) 6000 (*iii*) 32.362400

(*iv*) 5.6034 (*v*) decimal equivalent of 2/3

(*vi*) 1.6276×10^4

3. The density of ice is 0.921 g cm⁻³. Calculate the mass of a cubic block of ice which is 76 mm on each side.

4. Express the following numbers in exponential notations to three significant figures :

(*i*) 900035 (*ii*) 0.000002136

(*iii*) 406721 (*iv*) 0.000001

5. Perform the following calculations to proper number of significant figures :

(*a*) 108/7.2 (*b*) $(1.6 \times 10^2)^2$

(*c*) $\dfrac{(1.35\times10^{-6})(0.4)}{5.6}$ (*d*) $\dfrac{3.25 \times 0.08621}{4.002}$

(*e*) $(1.0042 - 0.0034)\ (1.23)$

6. Calculate

(*i*) area of a square whose side is 1.2 m

(*ii*) volume of a sphere whose radius is 1.6 cm

(*iii*) length of a rectangle having area $10.25\ m^2$ and breadth 2.5 m.

7. Perform the following calculations upto proper number of significant figures :

(*i*) $(1.20 \times 10^{-6}) + (6.00 \times 10^{-5}) = ?$

(*ii*) $(2.164 \times 10^5)^{\frac{1}{2}} = ?$

(*iii*) $(9.13 \times 10^{-2})\ (7.006 \times 10^{-3}) = ?$

(*iv*) $4.00 \times 10^{-2} + 3.26 \times 10^{-3} + 1 \times 10^{-6} = ?$

8. Calculate the number of significant figures up to which the following results will be expressed :

(*i*) $\dfrac{2.36 \times 0.07251}{2.130}$ (*ii*) $\dfrac{(28.2 - 21.2)\ (1.79 \times 10^6)}{1.62}$

9. Round up the following upto three significant figures :

(*i*) 34.216 (*ii*) 10.4107

(*iii*) 0.04597 (*iv*) 2808

10. If the speed of light is $3.0 \times 10^8\ ms^{-1}$, calculate the distance covered by the light in 2.00 ns.

Answers to Practice Problems

1. (*a*) Three (*b*) Two (*c*) Four (*d*) One (*e*) Three

2. (*i*) 6.02×10^{23} (*ii*) 6.00×10^3 (*iii*) 32.4 (*iv*) 5.60 (*v*) 0.667 (*vi*) 1.63×10^4

3. 4.0×10^2 g

4. (*i*) 9.00×10^5 (*ii*) 2.14×10^{-6} (*iii*) 4.07×10^5 (*iv*) 1.00×10^{-6}

5. (*a*) 15 (*b*) 2.6×10^4 (*c*) 1×10^{-7} (*d*) 0.0700 (*e*) 1.23

6. (*i*) $1.4\ m^2$ (*ii*) $17\ cm^3$ (*iii*) 4.1 m

7. (*i*) 6.12×10^{-5} (*ii*) 4.652×10^2 (*iii*) 6.40×10^{-4} (*iv*) 4.33×10^{-2}

8. (*i*) Three (*ii*) Two

9. (*i*) 34.2 (*ii*) 10.4 (*iii*) 0.0460 (*iv*) 2.81×10^3

10. 0.60m

Hints & Solutions on page 79

DIMENSIONAL ANALYSIS

Conversion of Units

It is frequently necessary to convert one set of units to another in calculations. This can be done by a method called **factor label method** or **unit factor method** or also called **dimensional analysis.** This is based on the relationship between different units that expresses the same quantity. For example, suppose we want to convert 25.6 metres to centimetres. We know that

$$1\ m = 100\ cm$$

$$\text{or} \qquad 1 = \frac{100\ cm}{1\ m}$$

100 cm/1 m is called *unit conversion factor* or *conversion factor* because *it converts a quantity expressed in one unit to a quantity expressed in another unit.* It is also called *unit factor* because the overall effect of multiplication by these factors is to multiply by 1.

Now, multiply the given quantity by the conversion factor, retaining the units of the physical quantity and that of the conversion factor in such a way that all units cancel out leaving behind only the required units i.e., numerator should have that part which is required in the desired result.

$$\text{Thus, } 25.6\ m = 25.6\ \cancel{m} \times \frac{100\ cm}{1\ \cancel{m}} = 2560\ cm.$$

- ***It may be remembered that we choose the conversion factor that has metres in the denominator.***

The advantage of this method is that if the equation is set up correctly, only then all the units will get cancelled except the required unit. If we do not get desired unit, then an error must have been made.

Let us consider another example. Suppose we want to convert 1.50 km^3 of water into volume in litres. We know that

$$1L = 1\ dm^3$$

So, we want to convert km^3 to dm^3 (litres). This can be done in two steps as :

(*i*) Convert km^3 to m^3

(*ii*) Convert m^3 to dm^3

For converting km^3 to m^3, the unit factor is

$$1\ km = 10^3\ m$$

so that $$1 = \left(\frac{10^3 m}{1\ km}\right)$$

$$\therefore \quad 1.50\ km^3 = 1.50\ \cancel{km^3} \times \left(\frac{10^3 m}{1\ \cancel{km}}\right)^3$$

$$= 1.50 \times 10^9\ m^3$$

Now $$1\ m = 10\ dm$$

or $$1 = \frac{10\ dm}{1\ m}$$

Applying unit factor,

$$\text{Volume of water} = 1.50 \times 10^9\ \cancel{m^3} \times \left(\frac{10\ dm}{1\ \cancel{m}}\right)^3$$

$$= 1.50 \times 10^{12}\ dm^3$$

Both the steps can be performed in a single step as

$$\text{Volume of water} = 1.50\ \cancel{km^3} \times \underbrace{\left(\frac{10^3\ \cancel{m}}{1\ \cancel{km}}\right)^3}_{\text{Converts } km^3 \text{ into } m^3} \times \underbrace{\left(\frac{10\ dm}{1\ \cancel{m}}\right)^3}_{\text{Converts } m^3 \text{ into } dm^3}$$

$$= 1.50 \times 10^{12}\ dm^3$$

Since cubic decimetre is equal to litres, the volume of water is

$$= 1.50 \times 10^{12}\ \text{litres}$$

It should also be noted that units can be handled just like numerical part i.e., it can be cancelled, multiplied, divided, squared, etc.

Let us convert litre atmosphere to joule (the SI unit of energy). Let us convert 1 L atm to joules. This can be done by multiplying with two unit factors as :

$$1\ L = 10^{-3}\ m^3 \quad \text{or} \quad 1 = \left(\frac{10^{-3}\ m^3}{L}\right)$$

$$1\ atm = 101.325 \times 10^3\ Pa$$

$$\text{or } 1 = \frac{101.325 \times 10^3\ Pa}{atm}$$

Applying these unit factors :

$$1\ L\ atm = \cancel{1\ L}\ \cancel{atm} \times \left(\frac{10^{-3}\ m^3}{\cancel{1L}}\right) \times \left(\frac{101.325 \times 10^3\ Pa}{\cancel{atm}}\right)$$

$$= 101.325\ Pa\ m^3$$

Now, $$Pa = \frac{N}{m^2}$$

$$\therefore \quad 101.325\ Pa\ m^3 = 101.325\left(\frac{N}{m^2}\right)m^3$$

$$= 101.325\ Nm$$

$$= 101.325\ J \qquad [\because Nm = J]$$

Thus, **1 L atm = 101.325 J**

Thus, we can easily apply several conversion factors in one step. For example, suppose we want to convert 2.50 miles into centimetres. The relationship between various units are :

1 mile = 1760 yards, 1 yard = 3 ft, 1 ft = 12 inches and 1 inch = 2.54 cm.

The unit factors are :

$$1 = \frac{1760\ yards}{1\ mile};\ 1 = \frac{3\ ft}{1\ yard};\ 1 = \frac{12\ inches}{1\ ft};$$

$$\text{and} \quad 1 = \frac{2.54\ cm}{1\ inch}$$

Applying these unit factors,

$$2.50\ \cancel{miles} \times \underbrace{\frac{1760\ \cancel{yards}}{1\ \cancel{mile}}}_{\text{Converts mile into yard}} \times \underbrace{\frac{3\ \cancel{ft}}{1\ \cancel{yard}}}_{\text{Converts yard into ft}} \times \underbrace{\frac{12\ \cancel{inches}}{1\ \cancel{ft}}}_{\text{Converts ft into inch}} \times \underbrace{\frac{2.54\ cm}{1\ \cancel{inch}}}_{\text{Converts inch into cm}}$$

$$= 402336\ cm$$

or $$= 4.02 \times 10^5\ cm$$

Table 6. Relation between various units.

Units of length	Units of mass	Units of volume
1 mile = 1760 yard	1 kg = 10^3 g	1 m^3 = 10^3 L
1 yard = 3 ft	1g = 1000 mg	1 dm^3 = 1 L
1 ft = 12 inch	1 lb = 453.59 g	1 cm^3 = 10^{-3} L
1 inch = 2.54 cm	1 oz = 28.35 g	1 ft^3 = 28.32 L
1 metre = 100 cm	1 metric ton = 1000 kg	1 quart = 0.9464 L
1 km = 1000 m	= 2205 lb	1 L = 1.056 quarts
1 mile = 1.609 km	1 g = 15.4 grains	
1 mile = 5280 ft	1 carat = 3.168 grains	

Other units

1 Å	= 10^{-10} m	1 dyne = 10^{-5} N
1 atm	= 760 mm or 760 torr	1 calorie = 4.184 J
	= 101,325 Pa or Nm^{-2}	1 erg = 10^{-7} J
	= 1.013×10^6 dyne cm^{-2}	1 *e*V = 1.6022×10^{-19} J
1 bar	= 10^5 Nm^{-2} = 10^5 Pa	
1 mm or 1 torr = 133.322 Pa or Nm^{-2}		

SOLVED EXAMPLES

Example 12.

The density of vanadium is 5.96 g cm⁻³. Convert the density to SI units of kg m⁻³.

Solution: Density = 5.96 g cm^{-3}

We know 1 kg = 1000 g

$$\therefore \quad \textit{Unit factor} = \frac{1\text{ kg}}{1000\text{ g}}$$

1 m = 100 cm

$$\therefore \quad \textit{Unit factor} = \frac{100\text{ cm}}{1\text{ m}}$$

Applying unit factor for g and cm^3, we get

$$\therefore 5.96\text{ g cm}^{-3} = \frac{5.96\,\cancel{g}}{\cancel{cm^3}} \times \frac{1\text{ kg}}{1000\,\cancel{g}} \times \left(\frac{100\,\cancel{cm}}{1\text{ m}}\right)^3$$

$$= 5.96 \times 10^3\text{ kg m}^{-3}$$

$$= \mathbf{5960\ kg/m^3}.$$

Example 13.

A jug contains 2 L of milk. Calculate the volume of milk in m^3. **N.C.E.R.T.**

Solution: We know

1 L = 1000 cm^3

$$\text{or} \quad \text{Unit factor} = \frac{1000\text{ cm}^3}{1\text{ L}}$$

1 m = 100 cm

$$\text{Unit factor} = \frac{1\text{ m}}{100\text{ cm}}$$

$$\text{Volume of milk} = 2\,\cancel{L} \times \left(\frac{1000\,\cancel{cm^3}}{1\,\cancel{L}}\right) \times \left(\frac{1\text{ m}}{100\text{ cm}}\right)^3$$

$$= \mathbf{2 \times 10^{-3}\ m^3}$$

Example 14.

Express each of the following in S I units :

(i) *93 million miles (this is the distance between the earth and the sun).*

(ii) *5 feet 2 inches (this is the average height of an Indian female).*

(iii) *100 miles per hour (this is the typical speed of Rajdhani Express).*

(iv) *0.74 Å (this is the bond length of hydrogen molecule).*

(v) *46°C (this is the peak summer temperature in Delhi).*

(vi) *150 pounds (this is the average weight of an Indian male).* **N.C.E.R.T.**

Solution:

(*i*) The S I unit of distance is metre (m)

1 mile = 1.60 kilometre = 1.60 × 1000 m

$$\textit{Unit factor} = \frac{1.60 \times 1000\text{ (m)}}{1\text{ mile}} = \frac{1.6 \times 10^3\text{ (m)}}{1\text{ mile}}$$

$$\therefore \quad 93\text{ million miles} = \frac{93 \times 10^6\,\cancel{(\text{miles})} \times 1.6 \times 10^3\text{ (m)}}{1\,\cancel{\text{mile}}}$$

$$= 93 \times 1.6 \times 10^9\text{ m}$$

$$= \mathbf{148.8 \times 10^9\ m = 1.49 \times 10^{11}\ m.}$$

(*ii*) 5 feet 2 inches = 62 inches

1 inch = 2.54×10^{-2} m

$$\textit{Unit factor} = \frac{2.54 \times 10^{-2}\text{ (m)}}{1\text{ inch}}$$

$$62\text{ inches} = \frac{62\,\cancel{(\text{inches})} \times 2.54 \times 10^{-2}\text{ m}}{1\,\cancel{\text{inch}}}$$

$$= 62 \times 2.54 \times 10^{-2}\text{ m}$$

$$= 157.48 \times 10^{-2}\text{ m} = \mathbf{1.57\ m.}$$

(*iii*) 1 mile = 1.60 km = 1.60×10^3 m

$$\textit{Unit factor} = \frac{1.60 \times 10^3\text{ m}}{1\text{ mile}}$$

1 hr = 60 × 60 s = 3.6×10^3 s

$$\textit{Unit factor} = \frac{3.6 \times 10^3\text{ s}}{1\text{ hr}}$$

$$\therefore \quad \text{Speed} = \frac{100\text{ miles}}{\text{hr}}$$

$$= \frac{100\,\cancel{\text{miles}}}{\cancel{\text{hr}}} \times \frac{1.60 \times 10^3\text{ m}}{1\,\cancel{\text{mile}}} \times \frac{1\,\cancel{\text{hr}}}{3.6 \times 10^3\text{ s}}$$

$$= \mathbf{44\ m\ s^{-1}}.$$

(*iv*) 1 Å = 10^{-10} m.

$$\textit{Unit factor} = \frac{10^{-10}\text{ m}}{1\text{ (Å)}}$$

$$\therefore \quad 0.74\text{ Å} = \frac{0.74\,\cancel{\text{Å}} \times 10^{-10}\text{ (m)}}{1\,\cancel{(\text{Å})}}$$

$$= 0.74 \times 10^{-10}\text{ m} \quad \text{or} = \mathbf{7.4 \times 10^{-11}\ m}$$

(*v*) 0°C = 273.15 K

46°C = 273.15 K + 46 K = **319.15 K**

(*vi*) 1 pound = 454×10^{-3} kg

$$\textit{Unit factor} = \frac{454 \times 10^{-3}\text{ (kg)}}{1\text{ (pound)}}$$

$$\therefore \quad 150\text{ pound} = \frac{150\,\cancel{(\text{pound})} \times 454 \times 10^{-3}\text{ (kg)}}{1\,\cancel{(\text{pound})}}$$

$$= 150 \times 454 \times 10^{-3}\text{ kg} = \mathbf{68.1\ kg.}$$

Example 15.

The mass of precious stones is expressed in terms of 'carat'. What is the mass of a ring in grams which contains 0.600 carat diamond and 8.500 g gold given that 1 carat = 3.168 grains and 1 g = 15.4 grains ?

Solution:

Weight of ring = 0.600 carat diamond + 8.500 g gold

Now 1 carat = 3.168 grains

$$\text{Unit factor} = \frac{3.168 \text{ grains}}{1 \text{ carat}}$$

1 g = 15.4 grains

$$\text{Unit factor} = \frac{1 \text{ g}}{15.4 \text{ grains}}$$

Weight of diamond = 0.600 carat

$$= 0.600 \cancel{\text{carat}} \times \frac{3.168 \cancel{\text{grains}}}{1 \cancel{\text{carat}}} \times \frac{1 \text{ gram}}{15.4 \cancel{\text{grains}}}$$

= 0.123 g

Total mass of ring = 8.500 + 0.123 **= 8.623 g**

Example 16.

A tennis ball was observed to travel at a speed of 96 miles per hour. Calculate the speed of the ball in metres per second.

Solution:

Speed of tennis ball = 96 miles per hour

Now 1 mile = 1.60 km = 1.60×10^3 m

$$\text{Unit factor} = \frac{1.60 \times 10^3 \text{ m}}{1 \text{ mile}}$$

$$1 \text{ hr} = 60 \times 60 \text{ s} = 3.6 \times 10^3 \text{ s}$$

$$\text{Unit factor} = \frac{3.6 \times 10^3 \text{ s}}{1 \text{ hr}}$$

$$\therefore \quad \text{Speed} = \frac{96 \text{ mile}}{\text{hr}}$$

$$= \frac{96 \cancel{\text{mile}}}{\cancel{\text{hr}}} \times \frac{1.60 \times 10^3 \text{m}}{1 \cancel{\text{mile}}} \times \frac{1 \cancel{\text{hr}}}{3.6 \times 10^3 \text{s}}$$

$$= \mathbf{42.7 \text{ m s}^{-1}}$$

Example 17.

(a) Convert the following in kilogram

(i) 0.91×10^{-27} g (mass of electron)

(ii) 1 fg (mass of human DNA molecule)

(b) Convert into metre

(i) 1.4 Gm (diameter of Sun)

(ii) 40 Em (thickness of Milky Way Galaxy)

Solution: (*a*) (*i*) 0.91×10^{-27} g (mass of electron)

$$1 \text{ g} = 10^{-3} \text{ kg}$$

$$\text{or} \quad \text{Unit factor, } 1 = \frac{10^{-3} \text{ kg}}{1 \text{ g}}$$

$$\therefore \quad 0.91 \times 10^{-27} \text{ g} = 0.91 \times 10^{-27} \cancel{\text{g}} \times \left(\frac{10^{-3} \text{ kg}}{1 \cancel{\text{g}}}\right)$$

$$= 9.1 \times 10^{-31} \text{ kg}$$

(*ii*) Mass of human DNA molecule = 1 fg

$$1 \text{ fg} = 10^{-15} \text{ g}$$

$$\text{or} \quad \text{Unit factor, } 1 = \frac{10^{-15} \text{g}}{1 \text{ fg}}$$

$$1 \text{ g} = 10^{-3} \text{ kg}$$

$$\text{or} \quad \text{Unit factor, } 1 = \frac{10^{-3} \text{ kg}}{1 \text{ g}}$$

$$\therefore \quad 1 \text{ fg} = 1 \text{ fg} \times \left(\frac{10^{-15} \cancel{\text{g}}}{1 \cancel{\text{fg}}}\right) \times \left(\frac{10^{-3} \text{kg}}{1 \cancel{\text{g}}}\right)$$

$$= 1 \times 10^{-18} \text{ kg}$$

(*b*) (*i*) 1.4 Gm (diameter of Sun)

We know $\quad 1 \text{ Gm} = 10^9 \text{ m}$

$$\text{or} \quad \text{Unit factor, } 1 = \frac{10^9 \text{m}}{1 \text{ Gm}}$$

$$\therefore \quad 1.4 \text{ Gm} = 1.4 \cancel{\text{Gm}} \times \left(\frac{10^9 \text{ m}}{1 \cancel{\text{Gm}}}\right) = \mathbf{1.4 \times 10^9 \text{ m}}$$

(*ii*) 40 Em (thickness of Milky Way Galaxy)

We know, $\quad 1 \text{ Em} = 10^{18} \text{ m}$

$$\text{or Unit factor, } 1 = \frac{10^{18} \text{m}}{1 \text{Em}}$$

$$\therefore \quad 40 \text{ Em} = 40 \cancel{\text{Em}} \times \left(\frac{10^{18} \text{m}}{1 \cancel{\text{Em}}}\right) = \mathbf{40 \times 10^{18} \text{ m}}$$

$$\text{or} \quad = \mathbf{4.0 \times 10^{19} \text{ m}}$$

Practice Problems

11. The wavelength of a yellow line in spectrum of sodium atom is 5896 Å. Express it in nm.

12. How many cubic centimetres are there in 1 m^3 ?

13. Convert

(*i*) 4.86 kg/L to grams per millilitre

(*ii*) 1.86 km to cm

(*iii*) 6.92×10^{-7} m to micrometres and Angstroms

(*iv*) 9.2×10^{-3} cm^3 to litres

14. What is the capacity of a tank 0.8 m long 10 cm wide and 50 mm deep ?

15. How many cubic centimetres are there in 100 L ?

16. Convert the following in kilogram :
(*i*) 500 Mg (mass of jumbo jet loaded)
(*ii*) 3.34×10^{-24} g (mass of hydrogen molecule)

17. Express the following in the designated units :
(*i*) 1.54 mm s^{-1} to pm μs^{-1}
(*ii*) 25 gL^{-1} to mg dL^{-1}
(*iii*) 25 L to m^3
(*iv*) 2.66 g cm^{-3} to µg μm^{-3}
(*v*) 4.2 Lh^{-2} to mL s^{-2}

18. Convert into metre :
(*i*) 7 nm (diameter of small virus)
(*ii*) 41 pm (distance of nearest star)

19. How many seconds are there in 2 days ?

N.C.E.R.T.

20. Convert the following into base units :
(*i*) 28.7 pm
(*ii*) 15.15 µs
(*iii*) 25365 mg

Answers to Practice Problems

11. 589.6 nm
12. 10^6 cm^3
13. (*i*) 4.86 g/mL (*ii*) 1.86×10^5 cm
(*iii*) 0.692 µm, 6920 Å (*iv*) 9.2×10^{-6} L
14. 4L
15. 10^5 cm^3
16. (*i*) 5.0×10^5 kg (*ii*) 3.34×10^{-27} kg
17. (*i*) 1.54×10^3 pm μs^{-1} (*ii*) 2.5×10^3 mg dL^{-1} (*iii*) 2.5×10^{-2} m^3 (*iv*) 2.66×10^{-6} µg μm^{-3} (*v*) 3.2×10^{-4} mL s^{-2}
18. (*i*) 7×10^{-9} m (*ii*) 41×10^{-12} m
19. 172800 s
20. (*i*) 2.87×10^{-11}m (*ii*) 1.515×10^{-5}s
(*iii*) 2.5365×10^{-2} kg

Hints & Solutions on page 79

LAWS OF CHEMICAL COMBINATION

In the seventeenth century, the scientists had been trying to find out methods for converting one substance into another. During their quantitative studies of chemical changes, they made certain generalisations. These generalisations are known as *laws of chemical combination.* These are :

1. Law of conservation of mass
2. Law of constant composition or definite proportions
3. Law of multiple proportions
4. Law of reciprocal proportions
5. Law of combining volumes (or Gay Lussac's law of combining volumes).

The first four laws of chemical combination deal with mass relationships while the fifth law deals with the volumes of the reacting gases.

1. Law of Conservation of Mass

This law deals with the relation between the *mass of reactants and the products during the chemical changes.* It was postulated by a French chemist Antoine Lavoisier in 1789.

Law of conservation of mass states that

during any physical or chemical change, the total mass of the products is equal to the total mass of the reactants.

In other words, ***matter can neither be created nor destroyed during any physical or chemical change.***

Therefore, this law is also known as **law of indestructibility of matter.**

Experimental verification of the law of conservation of mass.

The law can be verified with the help of *Landolt's* experiment. Landolt took the solutions of sodium chloride (NaCl) and silver nitrate ($AgNO_3$) separately in two limbs of a 'H' shaped tube (known as Landolt's tube). The tube was sealed and weighed. After weighing, the two solutions were mixed thoroughly by shaking the tube. As a result, the reaction occurred between silver nitrate and sodium chloride and a white ppt. of silver chloride is formed as:

$$AgNO_3(aq) + NaCl(aq) \longrightarrow \underset{\text{White ppt.}}{AgCl(s)} + NaNO_3(aq)$$

After the reaction, the tube was again weighed. It was observed that the weight remained practically unchanged. This verified the law of conservation of mass.

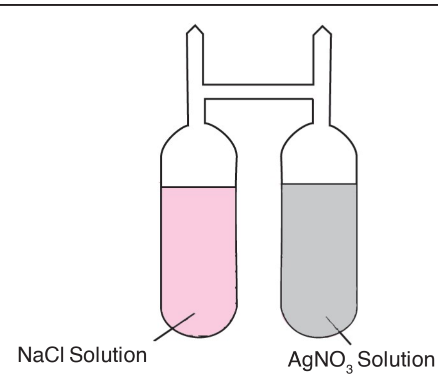


Fig. 8. Landolt's experiment.

LEARNING PLUS

Present Position of the Law in the Light of Recent Developments

It may be noted that the law of conservation of mass is not strictly valid for nuclear reactions.

According to Einstein theory of relativity, mass and energy are interconvertible. The mass (m) and energy (E) are related as, $E = mc^2$ where c is the velocity of light (3×10^8 m/sec). We know that chemical reactions are generally accompanied by liberation of energy. Since energy and mass are related to each other, this means that the energy must be coming from the reactants. As a result, there should be decrease in mass of the reactants. However, the mass changing into energy for ordinary chemical reactions (according to relation $m = E/c^2$) is extremely small because the value of c is very large and, therefore, there is no measurable change in mass during chemical processes. However, in case of **nuclear reactions** and **radioactive disintegrations,** the change in mass is quite significant because tremendous amount of energy is released during these reactions. Therefore, the law of conservation of mass does not hold good. In these reactions, some mass gets converted into energy. In such cases, mass and energy is totally conserved though mass and energy are not separately conserved. Thus, law of conservation of mass is modified and the modified law is known as **law of conservation of mass-energy** which states that ***mass and energy are interconvertible, but the total mass and energy of the system remains constant.***

2. Law of Constant Composition or Definite Proportions

The law of constant composition deals *with the composition of various elements present in a compound.* It was stated by a French chemist, *Joseph Proust,* in 1799. **Law of constant composition** states that

a pure chemical compound always contains same elements combined together in the same definite proportion by weight.

Proust worked with two samples of cupric carbonate; one of which was naturally occurring cupric carbonate and other was prepared in the laboratory. He found that the composition of the elements present in two samples of cupric carbonate was same as shown below :

	Percentage		
	Cu	C	O
Naturally occurring cupric carbonate	51.35	9.74	38.91
Cupric carbonate prepared in laboratory	51.35	9.74	38.91

The above results show that irrespective of the source, a given compound always contains same elements in the same proportion. This law is sometimes referred to as law of **definite composition** or **definite proportion**.

Similarly, carbon dioxide can be obtained by a number of methods, such as

(*i*) By burning coal or candle

$$C + O_2 \longrightarrow CO_2$$

(*ii*) By heating limestone ($CaCO_3$)

$$CaCO_3 \xrightarrow{\text{Heat}} CaO + CO_2$$

(*iii*) By the action of dilute hydrochloric acid on marble pieces.

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + CO_2 + H_2O$$

(*iv*) By heating sodium bicarbonate.

$$2NaHCO_3 \xrightarrow{\text{Heat}} Na_2CO_3 + CO_2 + H_2O$$

It has been observed that each sample of carbon dioxide contains carbon and oxygen elements in the ratio of 3 : 8 by weight.

Similarly, pure water can be obtained from many sources. Irrespective of the source, water always contains hydrogen and oxygen elements combined together in the ratio of 1 : 8 by weight.

Experimental verification of the law. The law can be easily verified in the laboratory. For example, copper oxide (CuO) can be prepared by the following methods :

(*i*) by heating copper powder in oxygen :

$$\underset{\text{Copper powder}}{2Cu} + O_2 \longrightarrow 2CuO$$

(*ii*) by heating copper carbonate :

$$\underset{\text{Copper carbonate}}{CuCO_3} \longrightarrow CuO + CO_2$$

(*iii*) by heating copper nitrate :

$$\underset{\text{Copper nitrate}}{2Cu(NO_3)_2} \longrightarrow 2CuO + 4NO_2 + O_2$$

The weighed quantities of three samples of copper oxide as obtained above were reduced to copper separately in a current of hydrogen :

$$CuO + H_2 \longrightarrow Cu + H_2O$$

The weight of copper left behind in each case is noted. From the weight of copper left behind, the weight of oxygen in the samples of copper oxides was calculated.

It was observed that in all the samples, the ratio of copper and oxygen by weight has been found to be the same *i.e.,* 4 : 1. This illustrates the law of constant composition.

Limitations of Law of Constant Composition

(*i*) The law of constant composition does not hold good when a compound is obtained by using different isotopes of the combining elements. For example, when CO_2 is formed from C-12 the ratio between C and O is 12 : 32. But when CO_2 is formed from C-14 isotope, the ratio of C : O is 14 : 32. Thus, *different isotopes of the same element give different mass ratio between combining atoms.*

(*ii*) The elements may combine in the same ratio but the compounds formed may be different.

For example, in the compounds C_2H_5OH and CH_3OCH_3 (both have same molecular formula C_2H_6O, known as isomers) the ratio of C : H : O is 24 : 6 : 16 or 12 : 3 : 8. Thus, *the inverse of the law is not correct*.

3. Law of Multiple Proportions

This law was proposed by Dalton in 1803. **Law of multiple proportions** states that

when two elements combine to form more than one compound, then the masses of one of the elements which combine with a fixed mass of the other element are in a simple whole number ratio.

The law may be **illustrated** by the following examples :

(*i*) Hydrogen combines with oxygen to form two compounds namely *water* and *hydrogen peroxide.*

In water, 2 parts by mass of hydrogen combine with 16 parts by mass of oxygen as :

$$\underset{2\text{ g}}{\text{Hydrogen}} + \underset{16\text{ g}}{\text{Oxygen}} \longrightarrow \underset{18\text{g}}{\text{Water}}$$

In hydrogen peroxide, 2 parts by mass of hydrogen combines with 32 parts by mass of oxygen as :

$$\underset{2\text{ g}}{\text{Hydrogen}} + \underset{32\text{ g}}{\text{Oxygen}} \longrightarrow \underset{34\text{g}}{\text{Hydrogen peroxide}}$$

Therefore, the masses of oxgen (16 g, 32 g) which combine with fixed mass of hydrogen (2 parts) bear a simple ratio *i.e.* 16 : 32 or 1 : 2.

(*ii*) Nitrogen and oxygen combine to form five oxides namely nitrous oxide (N_2O), nitric oxide (NO), nitrogen trioxide (N_2O_3), nitrogen tetroxide (N_2O_4) and nitrogen pentoxide (N_2O_5). The different weights of oxygen which combine with the fixed weight of nitrogen in all these oxides are calculated.

Oxide	Number of parts by weight of nitrogen	Number of parts by weight of oxygen	Fixed weight of nitrogen (14 parts)	Number of parts by weight of oxygen combining with 14 parts by weight of nitrogen
N_2O	28	16	14	8
NO	14	16	14	16
N_2O_3	28	48	14	24
N_2O_4	28	64	14	32
N_2O_5	28	80	14	40

The ratio between the different weights of oxygen in different compounds which combine with the same weight of nitrogen (14 parts) is :

$$8 : 16 : 24 : 32 : 40$$
$$\text{or} \quad 1 : 2 : 3 : 4 : 5$$

This is a simple whole number and hence, supports the law.

Experimental Verification of Law of Multiple Proportions

Copper forms two oxides cuprous oxide (Cu_2O) and cupric oxide (CuO). 1.00 gm of each oxide of copper (CuO and Cu_2O respectively) is heated in a current of hydrogen. Both the oxides react with hydrogen producing metallic copper. From the weight of copper obtained, the respective weights of oxygen in the two compounds are obtained. Then, the different weights of oxygen which combine with the same weight of copper in the two compounds are calculated. These weights are found to bear a simple whole number ratio. Thus, the law has been verified.

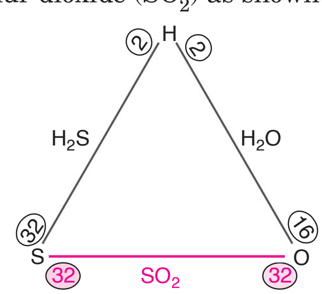
4. Law of Reciprocal Proportions or Law of Equivalent Proportions

This law was put forward by Richter in 1792. It states that

when two different elements combine separately with a fixed mass of a third element, the ratio in which they do so will be the same or some simple multiple of the ratio in which they combine with each other.

In other words, the mass ratio of two elements A and B which combine with the fixed mass of C separately, is either the same or some simple whole number multiple of the mass ratio in which A and B combine together. This law may be illustrated by the following examples :

Consider three elements sulphur, oxygen and hydrogen. Both sulphur and oxygen separately combine to form hydrogen sulphide (H_2S) and water (H_2O) respectively. They also combine with each other to form sulphur dioxide (SO_2) as shown:



According to the law, the ratio of weights of S and O which combine with the same weight of H will either be same or a simple multiple of the ratio in which S and O combine with each other. This may be verified as :

In hydrogen sulphide (H_2S), 2 parts by weight of hydrogen combine with 32 parts by weight of sulphur.

In water molecule (H_2O), 2 parts by weight of hydrogen combine with 16 parts by weight of oxygen.

$\therefore$ The ratio of the weights of sulphur and oxygen which, combine separately with the fixed weight (= 2 parts) of hydrogen is,

32 : 16 or 2 : 1 ...(*i*)

Now, let us calculate the ratio of sulphur and oxygen in SO_2.

In *sulphur dioxide* (SO_2)

Sulphur and oxygen combine in the ratio of 32 : 32 or 1 : 1 ...(*ii*)

The two ratios (*i*) and (*ii*) are related to each other as

$$\frac{2}{1} : \frac{1}{1} \text{ or } 2 : 1$$

which are simple multiple of each other.

However, sulphur and oxygen also react to form sulphur trioxide (SO_3) which is also in accordance with the law.

In *sulphur trioxide* (SO_3)

Sulphur and oxygen combine in the weight ratio of 32 : 48 or 2 : 3 ...(*iii*)

The two ratios (*i*) and (*iii*) are also related to each other as

$$\frac{2}{1} : \frac{2}{3} \text{ or } 3 : 1$$

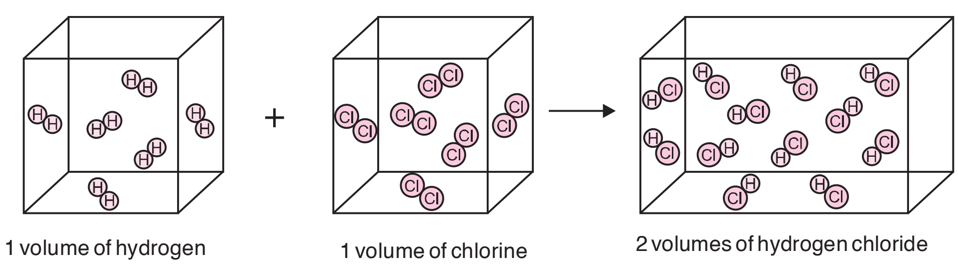
which are simple multiple of each other.

5. Gay Lussac's Law of Combining Volumes

Gay Lussac performed a number of experiments on reactions involving gases and found that some regularity exists between the volumes of the gaseous reactants and products. In 1808, he put forward a generalisation known as **Gay Lussac's law of combining volumes.** It states that

when gases react together or produced in a chemical reaction, they do so in a simple ratio by volume to one another and to the volumes of the products (if these are also gases) provided all gases are at the same temperature and pressure.

This law may be illustrated by the following examples :



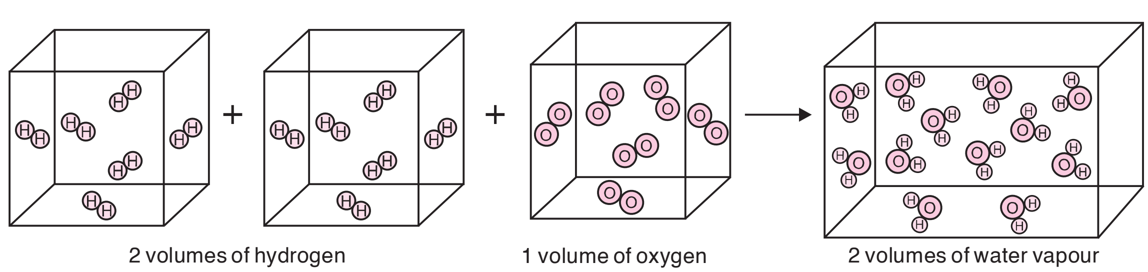
It has been experimentally found that 1 volume of hydrogen reacts with 1 volume of chlorine to give 2 volumes of hydrogen chloride as :

Thus, the volume ratio of hydrogen : chlorine : hydrogen chloride is 1 : 1 : 2. This is a simple whole number ratio and is also in agreement with their molar ratios when they are involved in the reaction.

$$\underset{\text{1 volume}}{H_2(g)} + \underset{\text{1 volume}}{Cl_2(g)} \longrightarrow \underset{\text{2 volumes}}{2HCl(g)}$$

For example, 100 mL of hydrogen combine with 100 mL of chlorine to give 200 mL of hydrogen chloride.

Similarly, we observe that 2 volumes of hydrogen combine with 1 volume of oxygen to give 2 volumes of water vapour as



Thus, the volumes of hydrogen and oxygen which combine to form water is in the ratio : 2 : 1 : 2.

$$\underset{2 \text{ volumes}}{2H_2(g)} + \underset{1 \text{ volume}}{O_2(g)} \longrightarrow \underset{2 \text{ volumes.}}{2H_2O(g)}$$

Similarly, it has been found that one volume of nitrogen combines with three volumes of hydrogen to give two volumes of ammonia gas.

$$\underset{1 \text{ volume}}{N_2(g)} + \underset{3 \text{ volumes}}{3H_2(g)} \longrightarrow \underset{2 \text{ volumes}}{2NH_3}$$

SOLVED EXAMPLES

Example 18.

If 6.3 g of $NaHCO_3$ are added to 15.0 g of CH_3COOH solution, the residue is found to weigh 18.0 g. What is the mass of CO_2 released in the reaction ?

Solution:

$$\underset{6.3 \text{ g}}{NaHCO_3} + \underset{15.0 \text{ g}}{CH_3COOH} \longrightarrow \underset{\substack{\text{Residue} \\ 18.0 \text{ g}}}{\underbrace{CH_3COONa + H_2O}} + \underset{x\,g}{CO_2}$$

Sum of the mass of reactants = Mass of $NaHCO_3$ + Mass of CH_3COOH
= 6.3 + 15.0 = 21.3 g

Sum of the mass of products = Mass of residue + Mass of CO_2
(18.0 + x) g
(where x is mass of CO_2 released)

According to law of conservation of mass

Mass of reactants = Mass of products

21.3 = 18.0 + x

or x = 21.3 – 18.0 = 3.3 g

∴ Mass of CO_2 released = **3.3 g**

Example 19.

Carbon and oxygen are known to form two compounds. The carbon content in one of these compounds is 42.9% while in the other, it is 27.3%. Show that the data are in the agreement with the law of multiple proportions. **N.C.E.R.T.**

Solution: *In the first compound,*

Mass of carbon = 42.9 g

Mass of oxygen = 100 – 42.9 = 57.1 g

In the second compound,

Mass of carbon = 27.3 g

Mass of oxygen = 100 – 27.3 = 72.7 g

In the first compound.

Mass of carbon combining with 57.1 g of oxygen = 42.9 g

Mass of carbon combining with 1.0 g of

$$\text{oxygen} = \frac{42.9 \text{ g}}{57.1} = 0.75 \text{ g}$$

In the second compound,

Mass of carbon combining with 72.7 g of oxygen = 27.3 g

Mass of carbon combining with 1.0 g of

$$\text{oxygen} = \frac{27.3}{72.7} = 0.375 \text{ g}$$

The ratio of mass of carbon combining with fixed mass of oxygen (*i.e.*, 1 g) is

0.75 : 0.375 or **2 : 1.**

This is a simple ratio and therefore, illustrates the **law of multiple proportions**.

Example 20.

2.0 g of a metal burnt in oxygen gave 3.2 g of its oxide. 1.42 g of the same metal heated in steam gave 2.27 g of its oxide. Which law is shown by this data?

Solution: In the first compound,

3.2 g of metal oxide contained 2.0 g of metal

100 g of metal oxide contained metal

$$= \frac{2.0}{3.2} \times 100 = 62.5 \text{ g}$$

∴ % Metal in first compound = 62.5%

In the second compound,

2.27 g of metal oxide contained metal = 1.42 g

100 g of metal oxide contained metal

$$= \frac{1.42}{2.27} \times 100 = 62.55 \text{ g}$$

∴ % Metal in second compound = 62.55%

Thus, the percentage of metal in metal oxide obtained from two experiments is nearly same. Hence, the above data illustrate the law of constant composition.

Example 21.

Phosphorus and chlorine form two compounds. The first compound contains 22.54% by mass of phosphorus and 77.46% by mass of chlorine. In the second compound the percentages are 14.88 for phosphorus and 85.12 for chlorine. Show that these data are consistent with the law of multiple proportions.

Solution: *In the first compound,*

Percentage of phosphorus = 22.54

Percentage of chlorine = 77.46

Thus, 22.54 g of phosphorus combines with 77.46 g of chlorine.

In the second compound,

Percentage of phosphorus = 14.88

Percentage of chlorine = 85.12

Thus, 14.88 g of phosphorus combine with 85.12 g of chlorine.

Let us fix the mass of phosphorus as 1 g and find the different masses of chlorine which combine with 1 g of phosphorus in two compounds.

In the first compound,

Mass of chlorine which combines with 22.54 g of phosphorus = 77.46 g

The mass of chlorine which combines with 1 g of phosphorus = $\dfrac{77.46}{22.54}$ = 3.44 g

In the second compound,

Mass of chlorine which combines with 14.88 g of phosphorus = 85.12 g

Mass of chlorine which combines with 1 g of phosphorus

$$= \frac{85.12}{14.88} = 5.72$$

The ratio of the masses of chlorine which combines with the fixed mass of phosphorus (1 g) in the two compounds is

3.44 : 5.72

1 : 1.66 or 3 : 5 (approximately)

This is a simple whole number ratio. Therefore, the data is in agreement with the **law of multiple proportions.**

❑ Example 22.

Three oxides of lead on analysis were found to contain lead as under :

(i) 3.45 g of yellow oxide contains 3.21 g of lead.

(ii) 1.195 g of brown oxide contains 1.035 g of lead.

(iii) 1.77 g of red oxide contains 1.61 g of lead.

Show that these data illustrate law of multiple proportions.

Solution: The amounts of lead and oxygen in three oxides are :

(*i*) Yellow oxide : Mass of lead = 3.21 g
Mass of oxygen = 3.45 – 3.21
= 0.24 g

(*ii*) Brown oxide : Mass of lead = 1.035 g
Mass of oxygen = 1.195 – 1.035
= 0.160 g

(*iii*) Red oxide : Mass of lead = 1.61 g
Mass of oxygen = 1.77 – 1.61
= 0.16 g

Let us fix the mass of lead as 1 g and calculate the different weights of oxygen which combine with 1 g of lead in these oxides.

(*i*) *Yellow oxide*

Mass of oxygen which combines with 3.21 g of lead = 0.24 g

Mass of oxygen which combines with 1 g of lead

$$= \frac{0.24}{3.21} = 0.075 \text{ g}$$

(*ii*) *Brown oxide*

Mass of oxygen which combines with 1.035 g of lead = 0.160 g

Mass of oxygen which combines with 1 g of lead

$$= \frac{0.160}{1.035} = 0.15 \text{ g}$$

(*iii*) *Red oxide*

Mass of oxygen which combines with 1.61 g of lead = 0.16 g

Mass of oxygen which combines with 1 g of lead

$$= \frac{0.16}{1.61} = 0.10 \text{ g}$$

The ratio of different masses of oxygen which combine with same mass of lead (1 g) in these oxides is :

0.075 : 0.15 : 0.10

3 : 6 : 4

This is a simple ratio.

Hence, the data illustrate the **law of multiple proportions.**

❑ Example 23.

Two oxides of a metal contain 27.6% and 30% of oxygen respectively. If the formula of the first compound is M_3O_4, find the formula of the second compound.

Solution:

First oxide	*Second oxide*
Oxygen = 27.6%	Oxygen = 30%
Metal = 72.4%	Metal = 70%

Formula of first oxide = M_3O_4

Suppose the atomic weight of metal = x

Percentage of metal in the compound M_3O_4

$$= \frac{3x}{3x + 64} \times 100$$

$$\therefore \quad \frac{3x}{3x+64} \times 100 = 72.4$$

or $300\,x = 217.2\,x + 4633.6$

or $82.8\,x = 4633.6$ or $x = 56$

Now in the second oxide, metal and oxygen are 70% and 30%. Therefore, their atomic ratio will be

M : O

$\dfrac{70}{56} : \dfrac{30}{16}$

1.25 : 1.875

or 1 : 1.5

or 2 : 3

Therefore, formula of the compound = $\mathbf{M_2O_3}$.

Example 24.

Hydrogen sulphide (H_2S) contains 94.11% sulphur, water (H_2O) contains 11.11% hydrogen and sulphur dioxide (SO_2) contain 50% oxygen. Show that the results are in agreement with law of reciprocal proportions.

Solution:

In water (H_2O)

Mass of hydrogen = 11.11 g

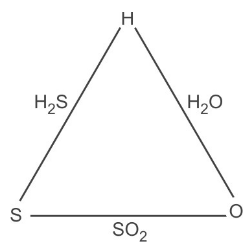
Mass of oxygen = 100 – 11.11 g

= 88.89 g

In sulphur dioxide (SO_2)

Mass of sulphur = 50 g

Mass of oxygen = 100 – 50 g = 50 g



Let us fix the mass of oxygen as 1 g.

Now in H_2O,

88.89 g of oxygen combine with hydrogen = 11.11

1 g of oxygen combines with hydrogen $= \frac{11.11}{88.89} \times 1$

= 0.125 g

In SO_2,

50 g of oxygen combine with sulphur = 50 g

1 g of oxygen combines with sulphur $= \frac{50}{50} \times 1 = 1$ g

Thus, the ratio of the masses of hydrogen and sulphur which combine with the fixed mass of oxygen (1 g) is

0.125 : 1 or 1 : 8 ...(*i*)

In hydrogen sulphide (H_2S)

Mass of sulphur = 94.11 g

Mass of hydrogen = 100 – 94.11 = 5.89 g

Therefore, the ratio by mass of hydrogen and sulphur in hydrogen sulphide is

5.89 : 94.11 or 1 : 16 ...(*ii*)

The ratios (*i*) and (*ii*) are

$$\frac{1}{8} : \frac{1}{16} \text{ or } 2 : 1$$

which is simple whole number.

Hence, the law of reciprocal proportions is verified.

Example 25.

Water contains 88.90% oxygen and 11.10% of hydrogen, ammonia contains 82.35% of nitrogen and 17.65% of hydrogen and dinitrogen trioxide contains 63.15% oxygen and 36.85% of nitrogen. Show that these data illustrate law of reciprocal proportions.

Solution:

In water (H_2O)

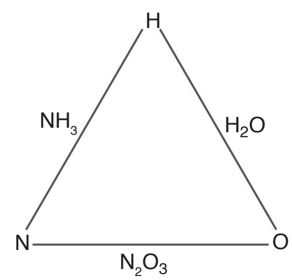
Mass of hydrogen = 11.10 g

Mass of oxygen = 88.90 g

In ammonia (NH_3)

Mass of hydrogen = 17.65 g

Mass of nitrogen = 82.35 g



Let us fix the mass of hydrogen as 1 g.

Now in H_2O,

11.10 g of hydrogen combine with oxygen = 88.90 g

1 g of hydrogen combines with oxygen $= \frac{88.90}{11.9}$

= 8.01 g

In NH_3,

17.65 g of hydrogen combine with nitrogen = 82.35 g

1 g of hydrogen combines with nitrogen $= \frac{82.35}{17.65}$

= 4.67 g

Thus, the ratio of nitrogen and oxygen which combine with the fixed mass of hydrogen (1 g) is

4.67 : 8.01 or 1 : 1.172 ...(*i*)

In dinitrogen trioxide

Mass of nitrogen = 36.85 g

Mass of oxygen = 63.15 g

Therefore, the ratio by mass of nitrogen and oxygen in dinitrogen trioxide is

36.85 : 63.15 or 1 : 1.172 ...(*ii*)

Thus, the two ratios are the same. Hence, it illustrates the law of reciprocal proportions.

Practice Problems

21. When 4.2 g of $NaHCO_3$ is added to a solution of CH_3COOH (acetic acid) weighing 10.0 g, it is observed that 2.2 g of CO_2 is released to the atmosphere. The residue left is found to weigh 12.0 g. Show that these observations are in agreement with the law of conservation of mass.

Hint : Mass of reactants = 14.2 g : Mass of products = 14.2 g

22. Hydrogen peroxide and water contain 5.93% and 11.2% of hydrogen respectively. Show that the data illustrates law of multiple proportions.

Hint : The ratio of weights of oxygen combining with 1 g of hydrogen is 2 : 1.

23. Elements A and B combine to form two different compounds as :

$$0.3 \text{ g of A} + 0.4 \text{ g of B} \longrightarrow 0.7 \text{ g of compound X}$$

$$18 \text{ g of A} + 48 \text{ g of B} \longrightarrow 66 \text{ g of compound Y.}$$

Show that the data illustrates the law of multiple proportions.

24. Carbon combines with hydrogen to form compounds having the following compositions :

Compound	Carbon (%)	Hydrogen (%)
A	75	25
B	85.7	14.3
C	92.3	7.7

Show that the data illustrate the law of multiple proportions.

Hint : Ratio between weights of H which combine with fixed weight of C is 4 : 2 : l.

25. 1.375 g CuO was reduced by hydrogen and 1.098 g Cu was obtained. In another experiment, 1.178 g of Cu was dissolved in nitric acid and the resulting copper nitrate was converted into CuO by ignition. The weight of CuO formed was 1.476 g. Show that these results prove the law of constant composition.

26. Illustrate the law of definite proportions from the following data :

(*i*) 0.32 g of sulphur on burning in air produced 224 ml of SO_2 at N.T.P.

(*ii*) A metal sulphite reacts with a mineral acid to give SO_2 gas which contains 50% of sulphur and 50% of oxygen.

27. Copper sulphate crystals contain 25.45% Cu and 36.07% H_2O. If the law of constant proportions is true then calculate the weight of Cu required to obtain 40 g of crystalline copper sulphate.

Ans. 10.18 g

28. Copper sulphide contains 66.5% Cu, copper oxide contains 79.9% Cu and sulphur trioxide contains 40% S. Show that the data illustrates the law of reciprocal proportions.

29. Two oxides of a certain metal were separately heated in a current of hydrogen until constant weights were obtained. The water produced in each case was carefully collected and weighed. It was observed that 1 g of each oxide gave 0.1254 g and 0.2263 g of water respectively. Show that the data illustrate the law of multiple proportion.

Hints & Solutions on page 79

DALTON ATOMIC THEORY

To provide theoretical justification to the laws of chemical combination, John Dalton postulated a simple theory of matter. The **basic postulates** of the theory are given below :

(*i*) *Matter is made up of extremely small indivisible and indestructible ultimate particles called atoms.*

(*ii*) *Atoms of the same element are identical in all respects i.e. in shape, size, mass and chemical properties.*

(*iii*) *Atoms of different elements are different in all respects and have different masses and chemical properties.*

(*iv*) *Atom is the smallest unit that takes part in chemical combinations.*

(*v*) *Atoms of two or more elements combine in a fixed ratio to form compound atoms (now a days called molecules).*

(*vi*) *Atoms can neither be created nor destroyed during any physical or chemical change.*

(*vii*) *Chemical reactions involve only combination, separation or rearrangement of atoms.*

Limitations of Dalton's Atomic Theory

Dalton' s atomic theory gave a powerful initiative to scientists in the beginning of the 19th century. The main achievement of the theory was that one could derive the laws of chemical combination from it. It held the field for a century. But as a result of discoveries in the beginning of 20th century by Sir J. J. Thomson, Rutherford, Neils Bohr and others, the atomic theory was reviewed and modified. The main drawbacks of Dalton's atomic theory are :

(*i*) It could explain the laws of chemical combination by mass but failed to explain the law of gaseous volumes.

(*ii*) It could not explain why atoms of different elements have different masses, sizes, valencies, etc.

(*iii*) It could not explain how and why atoms of different elements combine with each other to form compound atoms or molecules.

(*iv*) It failed to explain the nature of forces that bind together atoms in a molecule.

(*v*) It gave no satisfactory explanation between the ultimate particle of an element and that of a compound.

Modern Atomic Theory

The main modifications made in the Dalton's atomic theory as a result of new discoveries about atom are :

1. Atom is no longer considered to be indivisible. It has been found that an atom has a complex structure and is made up of subatomic particles such as electrons, protons and neutrons.

2. Atoms of same elements may not be similar in all respects. Atoms of same element may possess different relative masses. For example, there are two different types of atoms of chlorine with atomic masses 35 a.m.u. and 37 a.m.u. respectively. Such atoms of the same element which possess different atomic masses are called **isotopes.**

3. Atoms of different elements may have similar one or more properties. There are certain atoms of different elements which possess same relative masses. For example, atomic mass of calcium and argon is same (40 a.m.u.) but their chemical properties are entirely different. Such atoms of different elements which possess same mass are called **isobars.**

4. Atom is the smallest unit which takes part in chemical reactions. Though an atom is composed of subatomic particles yet it is the smallest particle which takes part in chemical reactions.

5. The ratio in which the different atoms combine may be fixed and integral but may not always be simple. For example, in sugar molecule ($C_{12}H_{22}O_{11}$), the ratio of C, H and O atoms is 12 : 22 : 11, which is not simple.

6. Atom of one element may be changed into atoms of other element. For example, atoms of nitrogen can be changed into oxygen atoms by interaction with α-rays. Similarly, uranium ($^{235}_{92}U$) can be converted into plutonium ($^{239}_{94}Pu$) through interaction with neutrons. This process is called **transmutation.** Thus, atom is no longer indestructible.

7. The mass of atom can be changed into energy. According to Einstein's equation, $E = mc^2$ (E = energy, m = mass and c = velocity of light), mass and energy are interconvertible. Thus, **atom is no longer indestructible**. However, during chemical reactions, atom remains unchanged.

AVOGADRO'S HYPOTHESIS

While trying to correlate Gay Lussac's law of gaseous volumes and Dalton's atomic theory, Berzelius, the Swedish scientist, put forward a generalisation, in 1811 known as *Berzelius hypothesis.* According to **Berzelius hypothesis :**

equal volumes of all gases contain equal number of atoms under similar conditions of temperature and pressure.

However, when the Berzelius hypothesis was applied to some chemical reactions it was observed that even a fraction of atoms was involved in some cases. The idea of fraction of atoms is against the concept of Dalton's atomic theory :

For example, consider the formation of hydrochloric acid gas (HCl). It has been observed experimentally that one volume of hydrogen combines with one volume of chlorine to produce two volumes of hydrochloric acid gas. This may be expressed as :

Hydrogen	+ Chlorine	⟶ Hydrochloric acid gas
1 vol	1 vol	2 vols

According to Berzelius hypothesis, equal volumes will contain equal number of atoms. Let one volume contains n atoms :

Hydrogen	+ Chlorine	⟶ Hydrochloric acid gas
or n atoms	n atoms	$2n$ compound atoms
or 1 atom	1 atom	2 compound atoms
½ atom	½ atom	1 compound atom

This means that one compound atom of hydrochloric acid gas is formed by the combination of ½ atom of hydrogen and ½ atom of chlorine. This means that fraction of atoms take part in chemical combination and therefore, atoms may undergo division during chemical reactions. But this idea of fraction of atoms is against the concept of Dalton's atomic theory according to which atom is the smallest unit of an element which takes part in chemical reactions.

To solve this problem of conflict between the Dalton's atomic theory and Berzelius hypothesis, **Avogadro** suggested that matter consists of two kinds of ultimate particles. These are *atoms* and *molecules.* According to him,

the smallest particle of an element which may or may not have independent existence and takes part in a chemical reactions *is called* ***an atom.***

The smallest particle of a substance (element or compound) capable of independent existence *is called* ***a molecule.***

Thus, the smallest particles of gases which can have independent existence are molecules and not the atoms and therefore, volume of gases should be related to number of molecules and not number of atoms. He could explain the results by considering the molecules to be **polyatomic.** He put forward a hypothesis known as **Avogadro's law.** It may be stated as :

under similar conditions of temperature and pressure, equal volumes of all gases contain equal number of molecules.

Mathematically, we can say

$$V \propto N$$

where N is the number of molecules.

For example, if we enclose equal volumes of three gases hydrogen (H_2), oxygen (O_2) and chlorine (Cl_2) in different flasks of the same capacity under similar conditions of temperature and pressure, we find that all the flasks have the same number of molecules. However, these molecules may differ in size and mass.

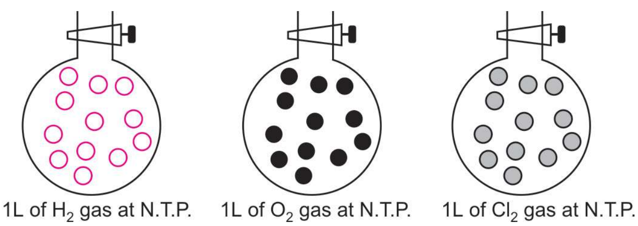


Fig. 10. Illustration of Avogadro hypothesis.

Gay-Lussac and Avogadro's laws can be illustrated as follows :

	$2H_2(g)$ +	$O_2(g) \longrightarrow$	$2H_2O\ (g)$
Gay-Lussac's law	2 vol	1 vol	2 vol
Avogadro's law	$2n$ molecules	n molecules	$2n$ molecules
	2 molecules	1 molecule	2 molecules

	$N_2(g)$ +	$3H_2(g) \longrightarrow$	$2NH_3\ (g)$
Gay-Lussac's law	1 vol	3 vol	2 vol
Avogadro's law	n molecules	$3n$ molecules	$2n$ molecules
	1 molecule	3 molecules	2 molecules

The Avogadro hypothesis can be successfully applied to different chemical reactions.

Formation of Hydrochloric acid gas

Consider the formation of hydrochloric acid gas. It is observed that 1 volume of hydrogen combines with 1 volume of chlorine to give 2 volumes of hydrogen chloride (all volumes under same conditions).

Hydrogen +	Chlorine $\longrightarrow$	Hydrogen chloride
1 vol	1 vol	2 vols

Applying Avogadro's hypothesis assuming that 1 volume contains n molecules, it follows that

Hydrogen +	Chlorine $\longrightarrow$	Hydrogen chloride
n molecules	n molecules	$2n$ molecules

Dividing throughout by $2n$, we get

$\frac{1}{2}$molecule	$\frac{1}{2}$molecule	1 molecule

This means that 1 molecule of hydrogen chloride contain ½ molecule of hydrogen and ½ molecule of chlorine.

Now ½ molecule of hydrogen can exist because one molecule of hydrogen contains 2 atoms of hydrogen and ½ molecule of hydrogen means one atom of hydrogen. Similarly, ½ molecule of chlorine contains an atom of chlorine because chlorine is also a diatomic molecule. Thus, one molecule of hydrogen chloride is formed from one atom of hydrogen and one atom of chlorine. *This is in agreement with Dalton's atomic theory.*

Applications of Avogadro's Law

1. Deduction of Atomicity of Elementary Gases

The number of atoms present in its one molecule of the substance is called its **atomicity.** Avogadro's law helps in determining the atomicity of elementary gases such as hydrogen, oxygen, chlorine, etc. For example,

Atomicity of oxygen. The atomicity of oxygen can be calculated by considering the reaction between hydrogen and oxygen to form water vapour. It has been experimentally found that 2 volumes of hydrogen react with 1 volume of oxygen to form 2 volumes of water vapour.

Hydrogen +	Oxygen $\longrightarrow$	Water vaopur
2 volumes	1 volume	2 volumes

Applying Avogardro's hypothesis

2 n molecules	n molecules	$2n$ molecules
2 molecules	1 molecule	2 molecules
1 molecule	$\frac{1}{2}$molecule	1 molecule

Thus, one molecule of water contains one molecule of hydrogen (2 atoms) and ½ molecule of oxygen. The molecular mass of water has been found to be 18 a.m.u. Since one molecule of hydrogen contains two atoms of hydrogen, therefore, weight of oxygen is one molecule of water is 18 – 2 = 16 a.m.u. This corresponds to the weight of one atom of oxygen. Thus, one atom is present in half molecule of oxygen. Therefore,

$$\frac{1}{2} \text{ molecule of oxygen} = 1 \text{ atom}$$

or 1 molecule of oxygen = 2 atoms

Thus, **atomicity of oxygen is 2.**

2. Determination of relationship between vapour density and molar mass of a gas

The vapour density of a gas is *the ratio between the mass of a certain volume of the gas to the mass of the same volume of hydrogen gas under similar* *conditions of temperature and pressure.* Thus,

Vapour density (V.D.) of a gas

$$= \frac{\text{Mass of certain volume of gas}}{\text{Mass of same volume of hydrogen}}$$

(similar conditions)

According to Avogadro's hypothesis, equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules. Let the given volume of the gas and hydrogen contain n molecules at S.T.P. conditions.

Vapour density

$$= \frac{\text{Mass of } n \text{ molecules of gas}}{\text{Mass of } n \text{ molecules of hydrogen}}$$

$$= \frac{\text{Mass of 1 molecule of gas}}{\text{Mass of 1 molecule of hydrogen}}$$

Since 1 molecule of hydrogen contains 2 atoms of hydrogen,

$$\text{Vapour density} = \frac{\text{Mass of 1 molecule of gas}}{\text{Mass of 2 atoms of hydrogen}}$$

But the ratio of the mass of one molecule of gas to the mass of an atom of hydrogen is called molar mass (discussed later).

$$\therefore \quad \text{Vapour density} = \frac{\text{Molar mass}}{2}$$

or **Molar mass = 2 × Vapour density**

Vapour density is also called *relative density* of the gas.

3. Determination of relationship between mass and volume of gas.

As discussed above,

$$\text{Molar mass} = 2 \times \text{Vapour density}$$

$$= 2 \times \frac{\text{Mass of certain volume of gas at S.T.P.}}{\text{Mass of same volume of hydrogen at S.T.P.}}$$

$$= 2 \times \frac{\text{Mass of 1L of gas at S.T.P.}}{\text{Mass of 1L of hydrogen at S.T.P.}}$$

But mass of 1L of hydrogen gas is 0.089 g

$$\therefore \text{ Molar mass} = 2 \times \frac{\text{Mass of 1L of gas at S.T.P.}}{0.089}$$

$$= \frac{2}{0.089} \times \text{Mass of 1L of gas at S.T.P.}$$

$$= 22.4 \times \text{Mass of 1L of gas at S.T.P.}$$

$$= \text{Mass of 22.4L of gas at S.T.P.}$$

Thus, 22.4 L of any gas at S.T.P. weigh equal to molar mass of gas expressed in grams. This is also called **Gram Molecular Volume (G.M.V.).**

add on Conceptual Questions 1

CONCEPTUAL QA

Q.1. How many significant figures are there in each of the following numbers :

(*i*) 1.00×10^6 (*ii*) 0.00010 (*iii*) π

Ans. (*i*) Three (*ii*) Two (*iii*) an infinite number.

Q.2. Convert 22.4 L in cubic metres.

Ans.

$$22.4 \text{ L} \times \left(\frac{10^3 \text{cm}^3}{\text{L}}\right) = 22.4 \times 10^3 \text{cm}^3$$

$$= 22.4 \times 10^3 \text{cm}^3 \times \left(\frac{10^{-6} \text{m}^3}{\text{cm}^3}\right)$$

$$= 22.4 \times 10^{-3} \text{ m}^3$$

Q.3. What physical quantities are represented by the following units and what are their common names ?

(*i*) kg m^2 s^{-2} (*ii*) kg m s^{-2} (*iii*) dm^3

Ans. (*i*) Energy, Joule (*ii*) Force, Newton (*iii*) Volume, Litre.

Q.4. The longest visible rays, at the end of the visible spectrum are 7.8×10^{-7} m in length. Express this length in (*i*) micrometers and (*ii*) nanometers.

Ans. (*i*)

$$1 \text{ m} = 10^6 \ \mu\text{m}$$

$$\text{Unit factor} = \frac{10^6 \mu\text{m}}{1\text{m}}$$

$$\therefore \quad 7.8 \times 10^{-7} \text{ m} = 7.8 \times 10^{-7} \text{m} \times \frac{10^6 \mu\text{m}}{1\text{m}} = 0.78 \ \mu\text{m}$$

(*ii*) $1\text{m} = 10^9 \text{ nm}$

$$\text{Unit factor} = \frac{10^9 \text{ nm}}{1\text{m}}$$

$$\therefore \quad 7.8 \times 10^{-7} \text{ m} = 7.8 \times 10^{-7} \text{ m} \times \frac{10^9 \text{ nm}}{1\text{m}} = 780 \text{ nm.}$$

Q.5. Which of the following mixture are homogeneous ?

(*a*) tap water (*b*) air (*c*) soil (*d*) smoke (*e*) cloud

Ans. Tap water, air, cloud.

Q.6. Is the molar volume of CO_2 same or different from CO ?

Ans. Same.

Q.7. At what temperature have the Celsius and Fahrenheit reading the same numerical value ?

Ans.
$$°C = \frac{5}{9}(F - 32)$$

$$x = \frac{5}{9}(x - 32)$$

$$4x = -160 \qquad \therefore x = -40$$

$\therefore -40°C, \quad -40°F.$

Q.8. What is the difference between 0.006 and 6.00 × 10^{-3} g ?

Ans. 0.006 g contains one significant digit while 6.00 × 10^{-3} g contains 3 significant digits.

Q.9. Calssify the following substances into elements, compounds and mixtures :

(*i*) Milk (*ii*) 22 carat gold (*iii*) Iodized table salt (*iv*) Diamond (*v*) Smoke (*vi*) Steel (*vii*) Brass (*viii*) Dry ice (*ix*) Mercury (*x*) Air (*xi*) Aerated drinks (*xii*) Glucose (*xiii*) Petrol (*xiv*) Glass (*xv*) Wood

Ans. Elements : (*iv*), (*ix*)

Compounds : (*viii*), (*xii*)

Mixtures : (*i*), (*ii*), (*iii*), (*v*), (*vi*), (*vii*), (*x*), (*xiii*), (*xiv*), (*xv*) ***N.C.E.R.T.***

Q.10. Given that density of water is 1 g mL^{-1}. What is the density in SI units?

Ans. $1 \text{ g mL}^{-1} = 1 \text{ g cm}^{-3} = 1 \cancel{g} \times \left(\frac{1 \text{ kg}}{1000 \cancel{g}}\right) \times \frac{1}{\cancel{cm^3}} \times \left(\frac{100 \cancel{cm}}{1\text{ m}}\right)^3 = \mathbf{1000 \text{ kg m}^{-3}}.$

Q.11. Is the law of constant composition true for all types of compounds? Explain why or why not?

Ans. Law of constant composition is not true for all types of compounds. It is true only for the compounds obtained from one isotope. For example, carbon exists in two common isotopes, ^{12}C and ^{14}C. When CO_2 is formed from ^{12}C, the ratio of masses is 12 : 32 or 3 : 8, but when it is formed from ^{14}C, the ratio will be 14 : 32 or 7 : 16, which is not same as in first case.

Q.12. Which postulate of the Dalton's atomic theory was modified after the discovery of isotopes ?

Ans. According to postulates of the Dalton's atomic theory, atoms of same element are identical in all respects *i.e.*, in shape, size, mass and chemical properties. However, after the discovery of isotopes, it was modified as : atoms of same element may not be similar in all respects. For example, there are two types of atoms of chlorine with atomic masses 35 amu and 37 amu respectively (called isotopes).

ATOMS AND MOLECULES

Atom

Atom is the ultimate particle of an element. To understand this, consider a piece of an element, say gold. If we break this into smaller and smaller pieces, it will become so small that we will not be able to see it without a microscope. If we keep on breaking this particle, we will ultimately reach a stage when this particle cannot be further broken. This ultimate particle of an element is called an **atom.** Thus,

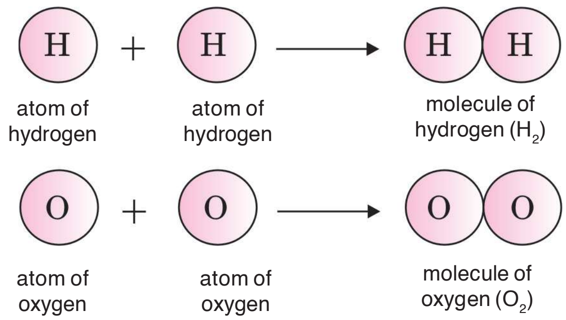
the smallest particle of an element which may or may not have independent existence is called an **atom**.

The atoms of certain elements, such as hydrogen, oxygen, nitrogen, etc. are not capable of independent existence whereas atoms of helium, neon, argon, etc., are capable of independent existence.

All elements are composed of atoms. There are about 118 elements known and therefore, there are 118 different types of atoms.

Molecule

Just as the ultimate particle of an element is **atom,** the ultimate particle of a chemical compound is called a **molecule.** When two or more atoms of different elements combine, the molecule of a comound is formed. For example,



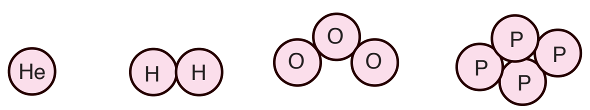
The molecules are made up of atoms of the same or different elements and are capable of independent existence. Thus,

***the smallest particle of a substance (element or compound) which is capable of independent existence** is called a* **molecule.**

Thus, a molecule contains two or more atoms. The properties of a substance are due to the properties of its molecules. Molecules can be classified into two types :

(*i*) **Homoatomic molecules.**

These molecules are made up of the atoms of the same element. Most of the elementary gases consist of homoatomic molecules. For example, hydrogen gas consists of two atoms of hydrogen (H_2). Similarly, chlorine gas consists of two atoms of chlorine (Cl_2). Similarly, nitrogen and oxygen gases are N_2 and O_2, respectively. These molecules may be further classified as *monoatomic, diatomic, triatomic,* or *polyatomic* molecules according as they contain one, two, three or more than three atoms respectively. The number of atoms in a molecule of an element is called **atomicity.** For example, He, Ne (monoatomic), O_2, N_2, H_2 (diatomic), O_3 (triatomic), P_4 (tetraatomic), S_8 (polyatomic), etc.

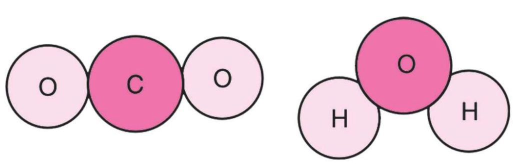


Helium (monoatomic) | Hydrogen (diatomic) | Ozone (triatomic) | Phosphorus P_4 (Tetraatomic)

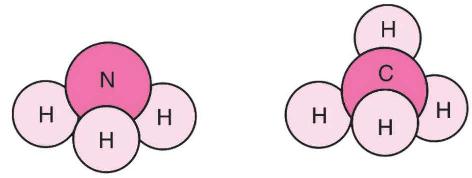
(*ii*) **Heteroatomic molecules.**

These molecules are made up of atoms of different elements. They are also classified as di, tri, tetra...etc. depending upon the number of atoms present. For example H_2O, HCl, CO_2, NH_3, CH_4, PCl_5 are heteroatomic molecules.

The molecules of some common compounds are shown below :



Carbon dioxide molecule CO_2 | Water molecule H_2O



Ammonia molecule NH_3 | Methane molecule CH_4

ATOMIC AND MOLECULAR MASSES

Atomic Mass

An atom of an element is so small particle that its mass cannot be determined even with the help of most sensitive balance. For example, by an indirect method, the absolute mass of hydrogen atom has been found to be 1.6736×10^{-24} g. Such extremely small numbers are very inconvenient for calculations. The difficulty was overcome by expressing *atomic masses, as* ***relative masses****. i.e., with respect to the mass of an atom of a standard substance.*

Earlier the chemists selected **hydrogen** atom as the standard substance because it is the lightest of all known atoms. The atomic mass of hydrogen was taken as one. Thus, the **atomic mass** of an element *is the number of times an atom of that element is heavier than an atom of hydrogen*. The relative atomic masses were referred to as the **atomic weights.** For example, it has been found that an atom of oxygen is 16 times heavier than an atom of hydrogen. Thus, the atomic mass of oxygen relative to hydrogen is 16. Later on, oxygen was fixed as the standard because it is more reactive than hydrogen and forms a large number of compounds.

In 1961, the International Union of Chemists selected a new unit for expressing the atomic masses. They accepted the isotope of carbon (^{12}C) with mass number 12 as the standard for comparing the atomic and molecular masses of elements and compounds. Carbon-12 is the stablest isotope of carbon and can be represented as ^{12}C.

Thus, **atomic mass of an element** is defined as

the average relative mass of an atom of an element as compared to the mass of an atom of carbon (^{12}C) taken as 12.

In other words, atomic mass is a number which expresses *as to how many times an atom of the element is heavier than 1/12th of the mass of carbon atom* (^{12}C).

Therefore,

$$\text{Atomic mass} = \frac{\text{Mass of an atom}}{\frac{1}{12}\text{ mass of a carbon atom } (^{12}C)}$$

This scale of relative masses of atoms is called **atomic mass unit scale** and is abbreviated as a.m.u. However, the new symbol used is 'u' (known as unified mass) in place of a.m.u. In this system, **^{12}C is assigned a mass of exactly 12 atomic mass unit** and masses of all other atoms are given relative to this standard. Hence, **one atomic mass unit** is defined as

the quantity of mass equal to 1/12th of the mass of an atom of carbon (^{12}C).

Mass of 1amu has been calculated to be **1 amu = 1.66056 × 10^{-24} g**

Mass of an atom of hydrogen = 1.6736 × 10^{-24} g

Thus, in terms of amu, the mass of hydrogen atom

$$= \frac{1.6736 \times 10^{-24}\ g}{1.66056 \times 10^{-24}\ g} = 1.008\ amu$$

Thus, the atomic mass of hydrogen is 1.008 u. Similarly, the mass of oxygen – 16 (^{16}O) atom is 15.995 u or 16 u; of nitrogen is 14 u and of sulphur is 32 u. This means that an atom of hydrogen is 1.008 times heavier than 1/12th of the mass of carbon atom (^{12}C); an atom of oxygen is 16 times heavier than $^1/_{12}$th of the mass of carbon atom (^{12}C) and so on.

The atomic masses of some common elements are given in Table 7. It may be noted that only the approximate values of atomic masses which are normally used in the chemical calculations, are given. For example, atomic mass of hydrogen is 1.008 g but it is taken as 1.

Nowadays atomic masses of the elements have been determined accurately using an instrument called **mass spectrometer.**

Table 7. Atomic mass of a few common elements referred to ^{12}C = 12.0

Element	Symbol	Atomic mass
Aluminium	Al	27.0
Antimony	Sb	121.8
Arsenic	As	74.9
Argon	Ar	18
Beryllium	Be	9.01
Bismuth	Bi	209
Boron	B	10.8
Bromine	Br	79.9
Calcium	Ca	40.1
Carbon	C	12.0
Chlorine	Cl	35.5
Copper	Cu	63.5
Fluorine	F	19.0
Helium	He	4.0
Hydrogen	H	1.0
Iodine	I	126.9
Irone	Fe	55.8
Lead	Pb	207.2
Lithium	Li	6.94
Manganese	Mn	54.9
Magnesium	Mg	24.3
Mercury	Hg	200.6
Neon	Ne	20.2
Nickel	Ni	58.7
Nitrogen	N	14.0
Oxygen	O	16.0
Phosphorus	P	31.0
Potassium	K	39.0
Silicon	Si	28.1
Silver	Ag	107.9
Sodium	Na	23.0
Sulphur	S	32.1
Tin	Sn	118.7
Zinc	Zn	65.4

Concept of Average Atomic Mass

The word average has been used in the above definition and is very significant. It has been found that the majority of the elements occur in nature as mixtures of several isotopes. **Isotopes** *are the different atoms of the same element possessing different atomic masses but same atomic number.* The atomic mass of such an element is the average of the atomic masses of the naturally occurring isotopes of the element taking into account their relative abundance. This is called **average atomic mass.**

For example, carbon occurs as three isotopes with relative abundances and masses as :

Isotope	Relative abundance (%)	Atomic mass (amu)
^{12}C	98.892	12
^{13}C	1.108	13.00335
^{14}C	2 × 10^{-10}	14.00317

Therefore, average atomic mass of carbon

$$= \frac{98.892 \times 12 + 1.108 \times 13.00335 + 2 \times 10^{-10} \times 14.00317}{100}$$

= **12.001 u**

Similarly, chlorine occurs in nature in the form of two isotopes with atomic mass 35 and 37 in the ratio of 3 : 1 respectively. Therefore,

$$\text{Average atomic mass of chlorine} = \frac{35 \times 3 + 37 \times 1}{3 + 1} = 35.5\ u$$

Thus, we say that on an average, an atom of chlorine is 35.5 times heavier than 1/12th of the mass of a carbon atom (C^{12}).

Similarly, neon occurs as three isotopes and their relative abundance are as follows :

Isotope	Fractional abundance
^{20}Ne	0.9051
^{21}Ne	0.0027
^{22}Ne	0.0922

Average atomic mass of neon

= 20 × 0.9051 + 21 × 0.0027 + 22 × 0.0922

= 20.179 u

Since the atomic mass is the ratio of the masses, it has no units. However, we express it as a.m.u., which only signifies that it is taken on atomic mass unit scale in which 1/12th of carbon atom is fixed as 1 u.

SOLVED EXAMPLES

Example 26.

Calculate the atomic mass of naturally occurring argon from the following data :

Isotope	Isotopic mass (g mol^{-1})	abundance
^{36}Ar	35.96755	0.337%
^{38}Ar	37.96272	0.063%
^{40}Ar	39.9624	99.600%

N.C.E.R.T.

Solution: Average atomic mass of argon

$$= \frac{35.96755 \times 0.337 + 37.96272 \times 0.063 + 39.9624 \times 99.60}{100}$$

$= \mathbf{39.948 g\ mol^{-1}}$

Example 27.

Boron occurs in nature in the form of two isotopes having atomic mass 10 and 11. What are the percentage abundances of two isotopes in a sample of boron having averge atomic mass 10.8 ?

Solution: Let the % abundance of ^{10}B isotope = x

Then % abudance of ^{11}B isotope = $100 - x$

The average atomic mass = $\frac{x \times 10 + (100 - x) \times 11}{100}$

But, average atomic mass = 10.8

$\therefore \quad \frac{x \times 10 + (100 - x) \times 11}{100} = 10.8$

or $\quad 10x + 1100 - 11x = 10.8 \times 100$

$-x = -1100 + 1080$

or $\quad x = 20$

Thus, percentage abundance; $\mathbf{^{10}B = 20,\ ^{11}B = 80}$

Gram Atomic Mass or Gram Atom

The atomic mass of an element expressed in grams is called **gram atomic mass.** *This amount of the element* is also called **one gram atom.** For example, the atomic mass of hydrogen is 1.008 a.m.u. and thus, gram atomic mass or a gram atom of hydrogen is 1.008 g.

Similarly, the atomic mass of oxygen = 16 a.m.u.

Therefore, gram atomic mass of oxygen = 16 g

However, the term *gram atom should not be confused with mass of one atom of the element*. For example, the mass of one atom of oxygen is only 2.66×10^{-23} g whereas, the gram atom of it (which is also the atomic mass of oxygen in grams) is 16 g. The mass of one atom of an element is known as **actual mass** of the atom.

Molecular Mass

Like atoms, the actual masses of molecules are also very small and cannot be measured by actual weighing. Like atomic masses, molecular masses are expressed relative to the stable isotope of carbon (^{12}C) having mass number 12. Thus,

Molecular mass of a substance may be defined as ***the average relative mass of its molecule as compared to the mass of an atom of carbon (^{12}C) taken as 12.***

In other words, molecular mass expresses as to how many times a molecule of the substance is heavier than 1/12th of the mass of an atom of carbon. For example, a molecule of carbon dioxide is 44 atom. Therefore, the molecular mass of carbon dioxide (CO_2) is 44 a.m.u.

Molecular mass from atomic masses. The molecular mass may be calculated as the sum of the atomic masses of all the atoms in a molecule. For example,

Molecular mass of methane (CH_4)

= At. mass of C + 4 × At. mass of H

= 12.001 + 4 × 1.008

= **16.033 u**

For simplicity, we may take rounded values of atomic masses of elements such as C = 12, H = 1, O = 16, ... so on.

Molecular mass of carbon dioxide (CO_2)

= At. mass of C + 2 × At. mass of O

= 12 + 2 × 16 = **44 u.**

Similarly,

Molecular mass of H_2SO_4 = 2 × Atomic mass of H + Atomic mass of S + 4 × Atomic mass of O

= 2 × 1 + 32 + 4 × 16

= **98 u**

Molecular mass of sucrose ($C_{12}H_{22}O_{11}$)

= 12 × Atomic mass of C + 22 × Atomic mass of H + 11 × Atomic mass of O

= 12 × 12 + 22 × 1 + 11 × 16

= **342 u**

Like atomic mass, molecular mass is expressed as a.m.u. which only signifies that it is taken on atomic mass unit scale.

Gram Molecular Mass or Gram Molecule

The molecular mass of a substance expressed in grams is called **gram molecular mass.** This amount of a substance is also known as **one gram molecule.** For example, the molecular mass of oxygen is 32 and, therefore, its gram molecular mass or gram molecule is 32 g.

It may be noted that the gram molecular mass may simply be taken as the *molecular mass expressed in grams.* However, this should not be confused with the mass of one molecule of the substance in grams. For example, the mass of one molecule of oxygen is only 5.32×10^{-23} g, whereas the gram molecular mass, which is the mass in grams equal to molecular mass, is 32 g. *The mass of one molecule of a substance, is known as its* **actual mass** of the molecule.

LEARNING PLUS

Formula mass

Some substances such as sodium chloride, potassium chloride (ionic compounds) do not have discrete molecules as their constituent units. In such compounds, positive (sodium) and negative (chloride) ions are arranged in a three dimensional structure. In this structure, it has been observed that each Na^+ ion is octahedrally surrounded by six Cl^- ions and each Cl^- ion is surrounded by six Na^+ ions. The structure of NaCl is shown below in Fig. 11.

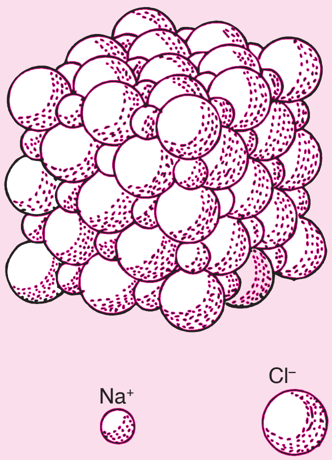


Fig. 11. Structure of NaCl.

In these compounds, we call formula unit (rather than molecular formula) which gives the simple ratio between various ions present. So, for these compounds we use **formula mass** instead of molecular mass. Formula mass is the sum of atomic masses of all atoms in a formula unit of the compound. For example,

Formula mass of sodium chloride

= At. mass of Na + At. mass of Cl

= 23.0 + 35.5

= 58.5 u.

It may be **remembered** that

Gram atomic mass

= Atomic mass expressed in grams

= gram atom

Gram molecular mass

= Molecular mass expressed in grams

= gram molecule

e.g., Gram atomic mass of oxygen or gram atom of oxygen = 16 g

Gram molecular mass of oxygen (O_2) or gram molecule = 32 g

SOLVED EXAMPLES

❑ Example 28.

Calculate

(a) mass of 1.5 gram atoms of calcium (at. mass = 40)

(b) gram atoms in 12.8 g of oxygen (at. mass = 16).

Solution:

(*a*) 1 gram atom of calcium = 40 g

1.5 gram atoms of calcium = 40 × 1.5 = **60 g**

(*b*) 1 gram atom of oxygen = 16 g

16 g of oxygen = 1 gram atom

$$12.8 \text{ g of oxygen} = \frac{1}{16} \times 12.8$$

= **0.8 gram atom**

❑ Example 29.

(a) Calculate the gram molecular mass of sugar having molecular formula $C_{12}H_{22}O_{11}$

(b) Calculate

(i) the mass of 0.5 gram molecule of sugar and

(ii) Gram molecule of sugar in 547.2 gram.

Solution: (*a*) Molecular mass of sugar ($C_{12}H_{22}O_{11}$)

= 12 × at. mass of C + 22 × at. mass of H + 11 × at. mass of O

= 12 × 12 + 22 × 1 + 11 × 16 = 342

∴ Gram molecular mass of sugar = 324 g

(*b*) (*i*) 1 gram molecule of sugar = 342 g

∴ 0.5 gram molecule of sugar = 342 × 0.5

= **171 g**

(*ii*) 342 g of sugar = 1 gram molecule

$$547.2 \text{ g of sugar} = \frac{1}{342} \times 547.2$$

= **1.6 gram molecule**

❑ Example 30.

Calculate the molecular mass of glucose ($C_6H_{12}O_6$) molecule. Given at. masses :

H = 1.008 amu, C = 12.011 amu, O = 16.0 amu.

N.C.E.R.T.

Solution: Molecular mass of glucose ($C_6H_{12}O_6$)

= 6 × At. mass of C + 12 × At. mass of H + 6 × At. mass O

= 6 × (12.011) + 12 × (1.008) + 6 × (16.00)

= **180.162 amu**

❑ Example 31.

Calculate

(a) mass of 2.6 gram molecule of SO_2.

(b) number of gram molecules of water in a beaker containing 576 g of water.

Solution: (*i*)Molecular mass of SO_2 = Atomic mass of S + 2 × Atomic mass of O

= 32 + 2 × 16 = 64

1 gram molecule of SO_2 = 64 g

2.6 gram molecules of SO_2 = 64 × 2.6 = 166.4 g

(*ii*) Molecular mass of H_2O = 2 × 1 + 16 = 18

∴ 18 g of water = 1 gram molecule

No. of gram molecules in 576 g = $\frac{1}{18} \times 576 =$ **32**

Practice Problems

30. The relative abundance of various isotopes of silicon is as Si (28) = 92.25%, Si (29) = 4.65% and Si (30) = 3.10%. Calculate the average atomic mass of silicon.

31. Calculate the mass of (*a*) 1.6 gram atoms of oxygen (*b*) 5.6 gram atoms of sulphur (*c*) 2.4 gram atoms of iodine.

(Atomic masses : O = 16, S = 32, I = 127)

32. Calculate the mass of

(*i*)2.5 gram molecules of H_2S, (*ii*) 3.6 gram molecules of glucose ($C_6H_{12}O_6$)

33. Calculate the number of

(*i*) gram atoms in 669.6 g of iron (at. mass = 55.8)

(*ii*) gram molecules in 73.6 g of C_2H_5OH.

34. Which of the following has maximum mass

(*a*) 2.6 gram atoms of sulphur (*b*) 2.6 gram molecules of sucrose ($C_{12}H_{22}O_{11}$) (*c*) 2.6 g of iodine

Answers to Practice Problems

30. 28.11 u.

31. (*a*) 25.6 g (*b*) 179.2 g (*c*) 304.8 g

32. (*i*) 85 g (*ii*) 648 g

33. (*i*) 12 (*ii*) 1.6

34. (*b*)

Hints & Solutions on page 79

MOLE CONCEPT

Quite commonly, we use different units for counting such as dozen for 12 articles, score for 20 articles and gross for 144 articles, irrespective of their nature. For example, one dozen books means 12 books whereas, one dozen apples means 12 apples. In a similar way, chemists use the unit mole for counting atoms, molecules, ions, etc. A mole is a collection of 6.022×10^{23} particles. Thus,

a mole represents 6.022×10^{23} particles

The number 6.022×10^{23} is called **Avogadro number or Avogadro constant**. It is denoted by N_A in honour of nineteenth century Italian scientist, Amedeo Avogadro. In other words, *a mole is an Avogadro number of particles.* For example,

1 mole of hydrogen atoms = 6.022×10^{23} hydrogen atoms.

1 mole of hydrogen molecules = 6.022×10^{23} hydrogen molecules.

1 mole of sodium ions = 6.022×10^{23} sodium ions.

1 mole of electrons = 6.022×10^{23} electrons.

It may be noted that the exact value of Avogadro number, N_A is 6.022137×10^{23}. However, in most of the calculations the value 6.022×10^{23} entities/mol for N_A is generally used.

A curious question comes to our mind. Why did chemists choose the number 6.022×10^{23} for a mole ? This seems to be very odd choice ? They would have selected simple number such as 10^3 (one thousand), 10^6 (one million), 10^{10}, 10^{20} or 10^{50}. However, this choice is based on the fact that chemists preferred the definition in *terms of a quantity that can be measured easily.* We have learnt that atoms and molecules are so small that we cannot see them even with a powerful microscope. Therefore, it is very difficult to count the actual number of atoms or molecules in a given sample of substance. Thus, weighing is easier than direct counting when the number of particles to be counted is very large. A mole was, therefore, originally defined in terms of mass. On this basis, chemists selected a standard number, which is equal to the number of atoms present in exactly 12.0 g of carbon (^{12}C).

Thus, a **mole** is defined as

the amount of substance that contains as many particles or entities (atoms, molecules or ions), as there are atoms in exactly 12 g (or 0.012 kg) of carbon-12 isotope.

In order to determine the number of atoms in 12 g of C-12, the mass of a carbon-12 was determined by a mass spectrometer. It was found to be equal to 1.992648×10^{-23} g. Knowing that 1 mol of carbon weights 12 g, the number of atoms in 1 mol of C-12 is

$$= \frac{12 \text{ g/mol of C} - 12}{1.992648 \times 10^{-23} \text{ g/atom of C}}$$

$$= 6.022137 \times 10^{23} \text{ atoms/mol}$$

Thus, we can say that a mole has 6.022137×10^{23} entities (atoms, molecules or ions, etc.). For simplicity, its value may be given equal to 6.022×10^{23} or Avogadro constant.

WATCH OUT !

It may be noted that while using the term mole, it is essential to specify the kind of particles involved. In case of certain substances, it is not clear whether we are talking of 1 mol atom or 1 mol molecule. If we simply say 1 mol of hydrogen, it means we are talking of naturally occurring form of hydrogen *i.e.,* H_2. So, 1 mol of hydrogen means 1 mole of hydrogen molecules. Consequently, to avoid confusion, we must be careful to specify the kind of particles when the term mole is used. For example, one mole of oxygen atoms contains 6.022×10^{23} atoms but one mole of oxygen molecules (O_2) contains 6.022×10^{23} molecules. Therefore, a mole of oxygen molecules is equal to two moles of oxygen atoms. However, if it is not mentioned, then we should consider that it is the natural form of that substance. These days, the molecular forms of hydrogen, nitrogen and oxygen are called dinitrogen (N_2), dihydrogen (H_2) and dioxygen (O_2) respectively to avoid confusion.

SOLVED EXAMPLES

Example 32.

How many molecules and atoms of sulphur are present in 0.1 mole of S_8 molecules ?

Solution:

$$1 \text{ mol of } S_8 \text{ molecule} = 6.022 \times 10^{23} \text{ molecules}$$

$$\therefore \quad 0.1 \text{ mol of } S_8 \text{ molecules} = 6.022 \times 10^{23} \times 0.1 = \mathbf{6.022 \times 10^{22} \text{ molecules}}$$

One molecule of S_8 contains 8 atoms of sulphur

$$\therefore \quad 6.022 \times 10^{22} \text{ molecules of } S_8 \text{ contain} = 6.022 \times 10^{22} \times 8 = \mathbf{4.816 \times 10^{23} \text{ atoms}}$$

Example 33.

Calculate the number of moles of iodine in a sample containing 1.0×10^{22} molecules.

Solution: 6.022×10^{23} molecules of iodine = 1 mol

1.0×10^{22} molecules of iodine

$$= \frac{1}{6.022 \times 10^{23}} \times 1 \times 10^{22}$$

$$= \mathbf{0.0166 \text{ mol}}$$

MASS-MOLE CONVERSIONS

Mole and Gram Atomic Mass

We have learnt that a mole represents amount of substance that contains same number of particles as the number of atoms in 12 g of carbon. But 12 g of carbon also correspond to 1 g atom of carbon (as discussed earlier). Similarly, it has been found that the mass of one mole atoms (6.022×10^{23} atoms) of hydrogen is equal to 1 g (or 1 gram atom of hydrogen). This can also be easily understood as given below.

We know that an atom of carbon is twelve times heavier than an atom of hydrogen. This means that the mass of an atom of hydrogen is 1/12th the mass of a carbon atom. Thus, if 6.022×10^{23} atoms of carbon weigh 12 g, the same number of hydrogen atoms will weigh $12 \times 1/12 = 1$g.

Similarly, the mass of 6.022×10^{23} atoms of oxygen, sodium and argon are 16 g (1 gram atom of oxygen), 23 g (1 gram atom of sodium) and 40 g (1 gram atom of argon) respectively. Thus, it may be concluded that

mass of 6.022×10^{23} atoms (or one mole atoms) of any element in grams is equal to its gram atomic mass or one gram atom.

In other words,

One mole of atoms ≡ 6.022×10^{23} atoms
≡ Gram atomic mass of the element

Mole and Gram Molecular Mass

The mass of 6.022×10^{23} molecules of a substance is equal to its gram molecular mass or gram molecule. For example, the mass of 1 mole molecules (6.022×10^{23} molecules) of water is equal to 18 g (1 gram mole molecules of oxygen is 32 g (1 gram mole of O_2) and that of 6.022×10^{23} molecules of carbon dioxide is 44 g (1 gram mole of CO_2). Thus,

mass of 6.022×10^{23} molecules (or 1 mole molecules) of any substance in grams is equal to its gram molecular mass or one gram molecule.

In other words,

One mole of molecules ≡ 6.022×10^{23} molecules
≡ Gram molecular mass

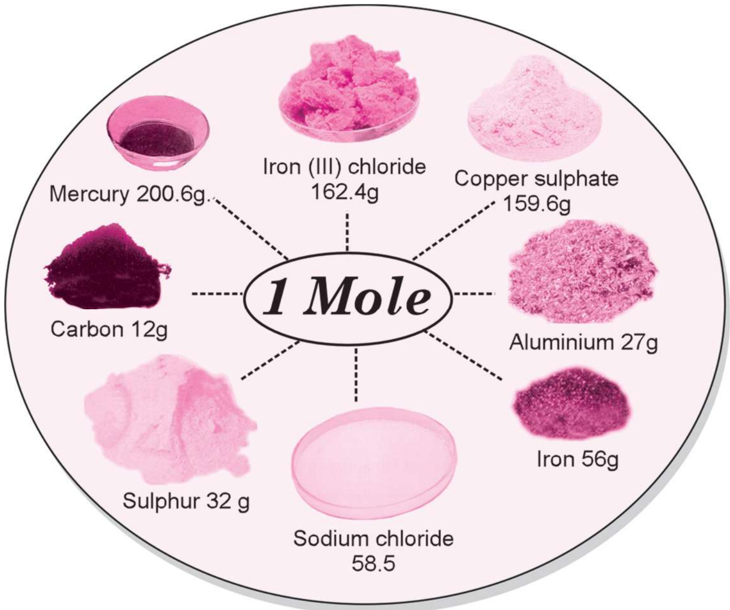
Mole Concept as applied to Ionic Compounds

It may be noted that the mole concept is also applicable to ionic compounds which do not contain molecules. These are made up of positively and negatively charged ions. In such cases, the formula of an ionic compound does not represent one molecule, but it gives the ratio between the constituent ions. The term mole is also applied for ionic compounds and one mole of formula units means 6.022×10^{23} formula units. The mass of one mole formula units in grams is equal to formula mass expressed in grams or gram formula mass of the compound. Thus,

mass of 6.022×10^{23} formula units (or one mole formula units) of any ionic substance in grams is equal to its gram formula mass.

For example, a mole of NaCl weighs equal to 58.5 g (one gram formula mass) and contains 6.022×10^{23} formula units of NaCl or 6.022×10^{23} Na^+ ions and 6.022×10^{23} Cl^- ions. Similarly, 1 mole of calcium chloride ($CaCl_2$) has mass equal to its gram formula mass *i.e.,* 111 g and contains 6.022×10^{23} Ca^{2+} ions and $2 \times 6.022 \times 10^{23}$ Cl^- ions.

1 mole of some substances (elements and compounds) are shown below :



The mass of 1 mol of a substance is called its **molar mass (M).** The units of molar mass are g mol^{-1} or kg mol^{-1}. Therefore, the molar mass is equal to atomic mass or molecular mass expressed in grams depending upon whether the substance contains atoms or molecules. We have also learnt that 1 mol of a substance contains 6.022×10^{23} entities or particles. Thus, mass of 1 mol of any substance will be its molar mass and it contains 6.022×10^{23} particles of that substance. For example, the molar mass of sodium *atoms* is the mass of 1 mol of sodium atoms, which is 23 g mol^{-1} and it contains 6.022×10^{23} sodium atoms. Similarly, the molar mass of water *molecules* is the mass of 1 mol of water molecules *i.e.,* 18 g mol^{-1} and it contains 6.022×10^{23} water molecules. The molar mass of NaCl is the mass of 1 mol of NaCl formula units *i.e.,* 58.5 g mol^{-1} and it contains 6.022×10^{23} formula units of NaCl.

Mole in Terms of Volume

Mole is also related to the volume of the gaseous substance. Volume of one mole of any substance is called its **molar volume.** The molar volume of solids and liquids can be easily calculated if we know the molar mass and density at any given temperature and pressure because these do not change much with temperature and pressure. However, the molar volumes of gases change considerably with temperature and pressure. It has been observed that one mole of an ideal gas (*i.e.,* 6.022×10^{23} molecules) occupies 22.4 litres at N.T.P. (0°C and 1 atm pressure). For example,

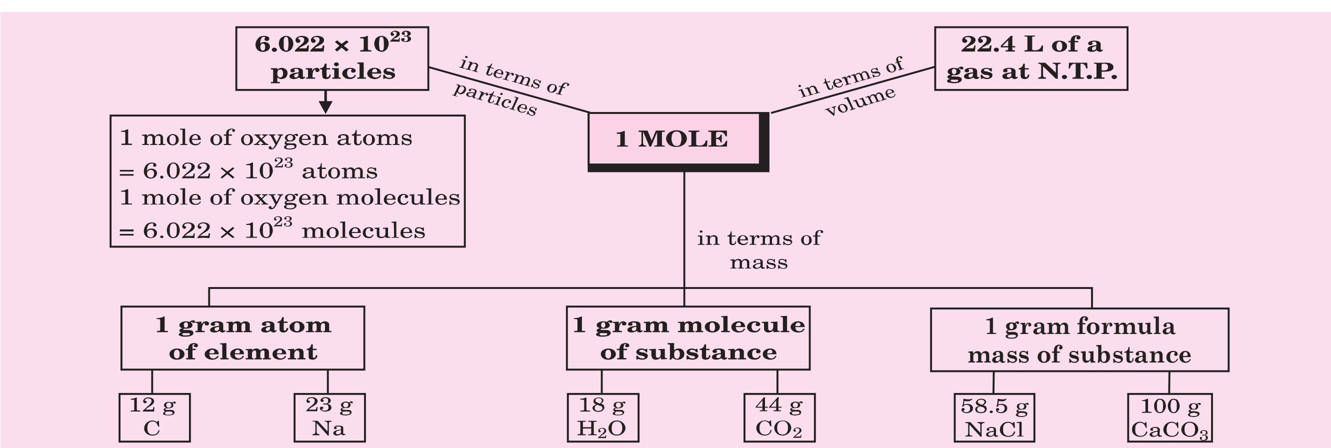
1 mole of hydrogen gas at N.T.P. = 22.4 litres

1 mole of CO_2 gas at N.T.P. = 22.4 litres

This provides us with a method to determine molecular masses (weights) of gases from experimentally determined density as

Molecular mass = Molar volume × Density

The summary of various relationships of mole may be illustrated as shown below :



These relations may also be **summarised** below :

One mole of atoms = 6.022×10^{23} atoms = Gram atomic mass of the element

Moles of an element = $\dfrac{\text{Mass of element}}{\text{Atomic mass}}$

Mass of one atom = $\dfrac{\text{Atomic mass}}{6.022 \times 10^{23}}$

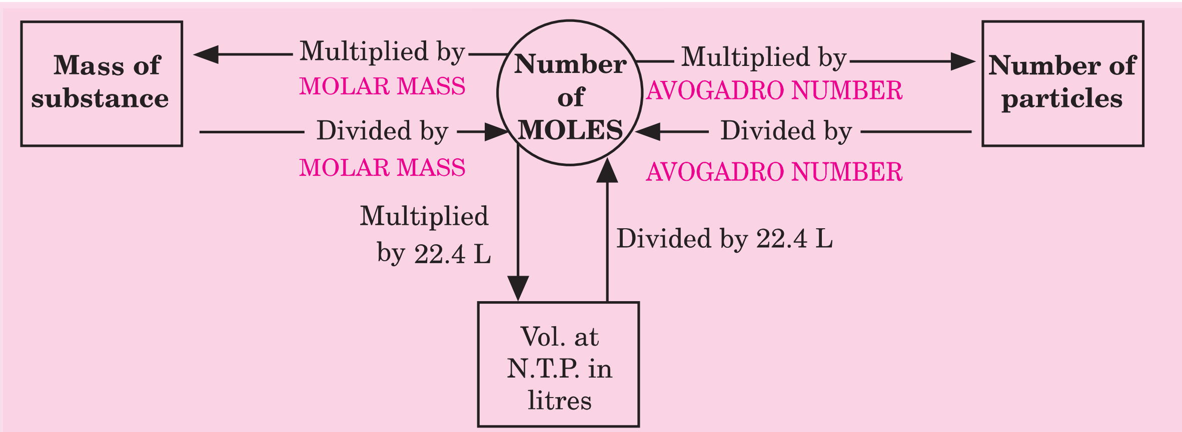
One mole of molecules = 6.022×10^{23} molecules = Gram molecular mass of the substance

Moles of a compound = $\dfrac{\text{Mass of compound}}{\text{Molecular mass}}$

Mass of one molecule = $\dfrac{\text{Molecular mass}}{6.022 \times 10^{23}}$

Volume occupied by 1 mole of a gas at N.T.P. = 22.4 L

The scheme for calculation of mole, mass and number of particles can be summarised below :



SOLVED EXAMPLES

Type I. Calculation of mass of atom(s) or molecule(s)

❑ Example 34.

(i) Calculate the mass of an atom of silver (atomic mass = 108).

(ii) 1 molecule of naphthalene ($C_{10}H_8$)

Solution:

(*i*) Mass of 6.022×10^{23} atoms of silver = 108 g

Mass of 1 atom of silver = $\dfrac{108}{6.022 \times 10^{23}}$ **= 1.793 × 10⁻²² g**

(*ii*) Molecular mass of naphthalene ($C_{10}H_8$)

= 10 × 12 + 8 × 1 = 128

Mass of 6.022×10^{23} molecules of napthalene = 128 g

Mass of 1 molecule of naphthalene = $\dfrac{128}{6.022 \times 10^{23}}$ = $\mathbf{2.12 \times 10^{-22}}$ **g**

❑ Example 35.

Calculate the mass of

(i) 1 atom of ^{14}C

(ii) 1 molecule of N_2

(iii) 1 molecule of water

(iv) 100 molecules of sucrose ($C_{12}H_{22}O_{11}$).

Solution:

(*i*) 6.022×10^{23} atoms of carbon ≡ Gram atomic mass of carbon (^{14}C) = 14 g

∴ Mass of 1 atom of ^{14}C = $\dfrac{14}{6.022 \times 10^{23}}$ = $\mathbf{2.32 \times 10^{-23}}$ **g**

(*ii*) 6.022×10^{23} molecules of N_2 = Gram molecular mass = 28 g

Mass of 1 molecule of N_2 = $\dfrac{28}{6.022 \times 10^{23}}$ = $\mathbf{4.65 \times 10^{-23}}$ **g**

(*iii*) 6.022×10^{23} molecules of H_2O = Gram molecular mass = 18 g

Mass of 1 molecule of H_2O = $\dfrac{18}{6.022 \times 10^{23}}$ = $\mathbf{2.99 \times 10^{-23}}$ **g**

(*iv*) 6.022×10^{23} molecules of sucrose = Gram molecular mass = 342 g

Mass of 100 molecules of sucrose = $\dfrac{342}{6.022 \times 10^{23}} \times 100$

= $\mathbf{5.68 \times 10^{-20}}$ **g**

❑ Example 36.

Calculate the mass of 1 u (atomic mass unit) in grams.

Solution: 1 u = $\dfrac{1}{12}$ th of the mass of carbon atom

Let us calculate the mass of a carbon atom.

6.022×10^{23} atoms of carbon = 12 g

$$\text{Mass of one carbon atom} = \frac{12}{6.022 \times 10^{23}} = 1.99 \times 10^{-23}\ g$$

$$\therefore \quad 1\ u = \frac{1}{12} \times 1.99 \times 10^{-23}\ g = \mathbf{1.66 \times 10^{-24}\ g}$$

Type II. Calculation of number of molecules present in a given volume of a gas

Example 37.

Calculate the number of molecules and number of atoms present in 11.2 litres of oxygen (O_2) at N.T.P.

Solution: We know that one mole of O_2 at N.T.P. occupies 22.4 litres.

$$11.2 \text{ litres of } O_2 \text{ at N.T.P.} = \frac{1}{22.4} \times 11.2 = 0.5 \text{ mole}$$

Now, 1 mole of O_2 contains $= 6.022 \times 10^{23}$ molecules

0.5 mole of O_2 contains $= 6.022 \times 10^{23} \times 0.5$

$= \mathbf{3.01 \times 10^{23}}$ **molecules**

1 molecule of oxygen $= 2 \times 3.01 \times 10^{23}$

$= \mathbf{6.02 \times 10^{23}}$ **atoms.**

Example 38.

The mass of 94.5 mL of a gas at S.T.P. is found to be 0.2231 g. Calculate its molecular mass.

Solution:

$$\text{Density of gas} = \frac{\text{Mass}}{\text{Volume}} = \frac{0.2231\ g}{94.5\ mL}$$

Molecular mass = Molar volume × Density

Molar volume = 22.4 L = 22.4 × 1000 mL

$$\therefore \text{ Molecular mass} = \frac{22.4 \times 1000 \times 0.2231}{94.5} = \mathbf{52.9}$$

Type III. Converting mass to moles or to atoms/molecules

Example 39.

Calculate the number of moles in the following : (a) 7.85 g of iron (b) 4.68 mg of silicon (c) 65.6 μg of carbon **N.C.E.R.T.**

Solution: (*a*) 7.85 g of iron

Atomic mass of Fe = 55.8 u

$$\text{Mole of iron} = \frac{\text{Mass of iron}}{\text{Atomic mass}} = \frac{7.85}{55.8} = \mathbf{0.141}\ \text{mol}$$

(*b*) 4.68 mg of silicon $= 4.68 \times 10^{-3}$ g of silicon

Atomic mass of silicon = 28.1

$$\text{Moles of silicon} = \frac{\text{Mass of silicon}}{\text{Atomic mass}} = \frac{4.68 \times 10^{-3}}{28.1} = \mathbf{1.66 \times 10^{-4}\ mol}$$

(*c*) 65.6 μg of carbon $= 65.6 \times 10^{-6}$ g of carbon

Atomic mass of carbon = 12

$$\text{Molecules of carbon} = \frac{\text{Mass of carbon}}{\text{Atomic mass}} = \frac{65.6 \times 10^{-6}}{12} = \mathbf{5.47 \times 10^{-6}\ mol}$$

Example 40.

Calculate the number of molecules in a drop of water weighing 0.05 g (H = 1, O = 16).

Solution: Gram molecular mass of H_2O = 18.0 g

18.0 g of H_2O contain $= 6.022 \times 10^{23}$ molecules

$$\therefore \quad 0.05 \text{ g of } H_2O \text{ contain} = \frac{6.022 \times 10^{23} \times 0.05}{18.0} = \mathbf{1.672 \times 10^{21}}\ \textbf{molecules}$$

Example 41.

Calculate the number of atoms in each of the following :

(i) 52 mol of Ar (ii) 52 u of He (iii) 52 g of He. **N.C.E.R.T.**

Solution:

(*i*) 1 mol of Ar contains $= 6.022 \times 10^{23}$ atoms

52 mol of Ar will contain $= 6.022 \times 10^{23} \times 52$

$= \mathbf{3.13 \times 10^{25}}$ **atoms**

(*ii*) 4 u of He = 1 atom

$$52 \text{ u of He} = \frac{1}{4} \times 52 = \mathbf{13\ atoms}$$

(*iii*) 4 g of He contain $= 6.022 \times 10^{23}$ atoms

$$52 \text{ g of He will contain} = \frac{6.022 \times 10^{23} \times 52}{4} = \mathbf{7.83 \times 10^{24}}\ \textbf{atoms}$$

Example 42.

Calculate the number of moles in the following masses :

(i) 1.46 metric tones of Al (1 metric ton = 10^3 kg)

(ii) 7.9 mg of Ca **N.C.E.R.T.**

Solution:

(*i*) 1.6 metric tons of Al $= 1.46 \times 10^3 \times 10^3$ g of Al $= 1.46 \times 10^6$ g

Atomic mass of Al = 27

$$\text{Moles of Al} = \frac{\text{Mass of Al}}{\text{Atomic mass}} = \frac{1.46 \times 10^6}{27} = 5.41 \times 10^4 \text{ mol}$$

(*ii*) 7.9 mg of Ca $= 7.9 \times 10^{-3}$ g of Ca

Atomic mass of Ca = 40.1

$$\text{Moles of Ca} = \frac{\text{Mass of Ca}}{\text{Atomic mass}} = \frac{7.9 \times 10^{-3}}{40.1} = \mathbf{1.97 \times 10^{-4}\ mol}$$

Example 43.

Suppose the chemists had chosen 10^{20} as the number of particles in a mole. What would be the molecular mass of oxygen gas ?

Solution: According to present concept of mole,

6.022×10^{23} molecules of oxygen weigh = 32 g

$$\therefore \quad \text{Mass of 1 molecule of oxygen} = \frac{32}{6.022 \times 10^{23}} = 5.32 \times 10^{-23}\ g$$

If the mole contained 10^{20} particles

The mass of one mole *i.e.,* 10^{20} particles

$= 5.32 \times 10^{-23} \times 10^{20} = 5.32 \times 10^{-3}$ g

$\therefore$ Molecular mass of oxygen = **5.32×10^{-3} g**

❑ Example 44.

Calculate the number of atoms of each type in 5.3 g of Na_2CO_3

Solution:

Gram molecular mass of $Na_2CO_3 = 2 \times 32 + 12 + 3 \times 16$
$= 106.0$ g

106.0 g of Na_2CO_3 = 1 mol

5.3 g of Na_2CO_3 = $\frac{1 \times 5.3}{106.0} = 0.05$ mol

Now, 1 mol of Na_2CO_3 contains $= 2 \times 6.022 \times 10^{23}$ sodium atoms

$\therefore$ 0.05 mol of Na_2CO_3 contains $= 0.05 \times 2 \times 6.022 \times 10^{23}$ sodium atoms = **6.022×10^{22} Na atoms**

1 mol of Na_2CO_3 contains $= 6.022 \times 10^{23}$ carbon atoms

0.05 mol of Na_2CO_3 contains $= 0.05 \times 6.022 \times 10^{23}$ carbon atom = **3.01×10^{22} carbon atoms**

1 mol of Na_2CO_3 contains $= 6.022 \times 10^{23}$ oxygen atoms

$\therefore$ 0.05 mol of Na_2CO_3 contains $= 0.05 \times 6.022 \times 10^{23}$ oxygen atoms

$\therefore$ 0.05 mol of Na_2CO_3 contains $= 0.05 \times 3 \times 6.022 \times 10^{23}$ oxygen atoms **$= 9.03 \times 10^{22}$ oxygen atoms**

❑ Example 45.

Calculate the number of molecules present in

(a) 1 kg oxygen,

(b) 1 dm^3 of hydrogen at S.T.P.

Solution: (*a*) 1 mole of oxygen or 32 g oxygen contains 6.022×10^{23} molecules of oxygen. Thus,

32 g of oxygen contain molecules $= 6.022 \times 10^{23}$

1000 g of oxygen contain molecules $= \frac{6.022 \times 10^{23}}{32} \times 1000$

$= 1.88 \times 10^{25}$

(*b*) We know that 1 mole of H_2 at S.T.P. occupies 22.4 dm^3 and contains 6.022×10^{23} molecules.

22.4 dm^3 of hydrogen contain molecules $= 6.022 \times 10^{23}$

1 dm^3 of hydrogen contains molecules $= \frac{6.022 \times 10^{23}}{22.4}$

$= 2.69 \times 10^{22}$

❑ Example 46.

Chlorophyll, the green colouring matter of plants contains 2.68% of magnesium by weight. Calculate the number of magnesium atoms in 2.00 g of chlorophyll (at. mass of Mg = 24). **N.C.E.R.T.**

Solution: Mass of chlorophyll = 2.0 g

Percentage of Mg = 2.68 g

Mass of Mg in 2.0 g of chlorophyll $= \frac{2.68 \times 2.0}{100}$

$= 0.054$ g

6.022×10^{23} atoms of magnesium = 24 g

24 g of Mg contains $= 6.022 \times 10^{23}$ atoms

0.054 g of Mg contains $= \frac{6.022 \times 10^{23}}{24} \times 0.054$

$= 1.3 \times 10^{21}$ atoms

Type IV. Calculation of sizes of individual atoms/ molecules.

❑ Example 47.

Calculate

(a) the actual volume of one molecule of water.

(b) the radius of a water molecule assuming to be spherical (density of water = 1 g ml^{-1}).

Solution:

(*a*) Density of water = 1 g ml^{-1}

$\therefore$ 1 g of H_2O = 1 ml

1 mole of water = 18 g of H_2O = 18 ml of water

$\therefore$ 6.022×10^{23} molecules of water occupy volume = 18ml

1 molecule of water occupy volume $= \frac{18}{6.022 \times 10^{23}}$

$= 2.99 \times 10^{-23}$ mol

(*b*) Since water molecule is assumed to be spherical, its volume is equal to $\frac{4}{3}\pi r^3$, where r is the radius of water molecule.

$$\therefore \quad \frac{4}{3}\pi r^3 = 2.99 \times 10^{-23}$$

$$\text{or} \quad r^3 = \frac{2.99 \times 10^{-23} \times 3}{4 \times 3.143} \quad (\pi = 3.143)$$

$$\text{or} \quad r = \left(\frac{2.99 \times 10^{-23} \times 3}{4 \times 3.143}\right)^{1/3}$$

$$= (7.13 \times 10^{-24})^{1/3} = 1.925 \times 10^{-8} \text{ cm}$$

$$= \mathbf{1.925\ Å}$$

Type V. Miscellaneous calculations on mole concept

❑ Example 48.

KBr (potassium bromide) contains 32.9% by weight of potassium. If 6.40 g of bromine reacts with 3.60 g of potassium, calculate the number of moles of potassium which combine with bromine to from KBr.

Solution:

Suppose mass of KBr = 100 g

Mass of potassium = 32.9 g

106.0 g of Na_2CO_3 = 1 mol

Mass of bromine = 67.1 g

67.1 g of bromine combine with potassium = 32.9 g

$\therefore$ 6.4 g of bromine will combine with potassium

$= \frac{32.9}{67.1} \times 6.4 = 3.14$ g

Atomic mass of potassium = 39 u

$\therefore$ Moles of potassium which combine with bromine to from KBr $= \frac{3.14}{39}$ **= 0.0805 mol**

❑ Example 49.

The cost of table salt (NaCl) and sugar ($C_{12}H_{22}O_{11}$) are ₹ 12 per kg and ₹ 36 per kg respectively. Calculate their cost per mole.

Solution: (*a*) *Cost of table salt per mole :*

Molecular mass of table salt (NaCl) = 23 + 35.5 = 58.5

Now, 1000 g of NaCl cost = ₹ 12

$\therefore$ 58.5 g of NaCl will cost = $\frac{12}{1000} \times 58.5$ = ₹ 0.702

= **70 paise (approx.)**

(*b*) *Cost of sugar per mole :*

Molecular mass of sugar ($C_{12}H_{22}O_{11}$) = 12 × 12 + 22 × 1 + 11 × 16 = 342

Now, 1000 g of sugar cost = ₹ 36

$\therefore$ 342 g of sugar will cost = $\frac{36}{1000} \times 342$ = ₹ 12.312

= ₹ **12 (approx)**

Example 50.

Silver is a very precious metal and is used in Jewellery. One million atoms of silver weigh 1.79 × 10^{-16} g. Calculate the atomic mass of silver.

Solution: Atomic mass of an element is the mass of 6.022×10^{23} atoms.

Now, 10^6 atoms of silver weigh = 1.79×10^{-16} g

6.022×10^{23} atoms of silver weigh

$$= \frac{1.79 \times 10^{-16} \times 6.022 \times 10^{23}}{10^6} = 107.8 \text{ g}$$

Atomic mass of silver = **107.8**

Example 51.

Calculate the weight of carbon monoxide having same number of oxygen atoms as are present in 88 g of carbon dioxide.

Solution: Molecular mass of CO_2 = 12 + 2 × 16 = 14

1 mole of CO_2 = 44 g

44 g of CO_2 contain = 6.022×10^{23} molecules

88 g of CO_2 contain = $\frac{6.022 \times 10^{23} \times 88}{44}$

= 12.044×10^{23}

Each molecule of CO_2 contains 2 oxygen atoms

No. of oxygen atoms in 12.046×10^{23} molecules of CO_2

= 2 × 12.046×10^{23} = 2.4088×10^{24}

Now, we are to calculate the mass of CO containing 2.4092×10^{24} atoms of oxygen

Mass of CO containing 6.022×10^{23} oxygen atoms

= 28 g

Mass of CO containing 2.4088×10^{24} oxygen atoms

$$= \frac{28}{6.022 \times 10^{23}} \times 2.4088 \times 10^{24} = \mathbf{112\ g}$$

Example 52.

A certain public water supply contained 0.10 parts per billion of chloroform ($CHCl_3$). How many molecules of $CHCl_3$ would be contained in a 0.05 ml drop of this water ?

Solution: Volume = 0.05 ml

Amount of $CHCl_3$ in the drop = $\frac{0.05 \times 0.1}{10^9} = 5.0 \times 10^{-12}$

Molecular mass of $CHCl_3$ = 12 + 1 + 3 × 35.5

= 119.5

Moles of $CHCl_3$ present = $\frac{5.0 \times 10^{-12}}{119.5}$

Molecules of $CHCl_3$ present in the drop

$$= \frac{5.0 \times 10^{-12}}{119.5} \times 6.022 \times 10^{23} = \mathbf{2.5 \times 10^{10}}.$$

Example 53.

Oxygen and nitrogen are present in a mixture in the ratio of 1 : 4 by weight. Calculate the ratio between their molecules.

Solution: Let wt. of given mixture = Wg

Amount of O_2 in mixture = $\frac{W \times 1}{5} = \frac{W}{5}$ g

Amount of N_2 in mixture = $\frac{W \times 4}{5} = \frac{4W}{5}$ g

Moles of O_2 = $\frac{W}{5 \times 32}$

Molecules of O_2 = $\frac{W}{5 \times 32} \times 6.022 \times 10^{23}$

Moles of N_2 = $\frac{4W}{5 \times 28}$

Molecules of N_2 = $\frac{4W}{5 \times 28} \times 6.022 \times 10^{23}$

Ratio of molecules of O_2 and N_2

$$\frac{W}{5 \times 32} \times 6.022 \times 10^{23} : \frac{4W}{5 \times 28} \times 6.022 \times 10^{23}$$

$$\frac{1}{32} : \frac{1}{7}$$

7 : 32.

Practice Problems

35. Calculate the mass of

(*i*) one atom of calcium

(*ii*) one molecule of sulphur dioxide (SO_2).

36. Calculate the number of atoms in

(*i*) 0.5 mole atoms of carbon (C^{12})

(*ii*) 3.2 g of sulphur

(*iii*) 18.0 g of glucose ($C_6H_{12}O_6$)

(*iv*) 0.20 mole molecules of oxygen.

37. Calculate the mass of sodium which contains same number of atoms as are present in 15 g of calcium (at. mass of Ca = 40, Na = 23).

38. What volume is occupied at N.T.P. by

(*i*) 1.4 g of nitrogen

(*ii*) 6.023×10^{21} molecules of oxygen

(*iii*) 0.2 mole of ammonia ?

39. How many years it would take to spend Avogadro number of rupees at the rate of 10 lakh rupees per second ?

40. One atom of an element 'X' weighs 6.644×10^{-23} g. Calculate the number of gram atoms in 80 kg of it.

41. Calculate the number of molecules and number of atoms present in 5.60 L of ozone (O_3) at N.T.P.

42. Calculate the number of gold atoms in 300 mg of a gold ring of 20 carat gold (atomic mass of gold = 197, pure gold is 24 carat).

43. What mass in kilogram of K_2O contains the same number of moles of K atoms as are present in 1.0 kg of KCl ?

44. How many molecules of water of hydration are present in 630 mg of oxalic acid ($H_2C_2O_4 . 2H_2O$) ?

45. How many molecules of CO_2 are present in one litre of air containing 0.03% by volume of CO_2 at STP ?

46. A dot '.' containing carbon has 1 microgram weight. Calculate number of carbon atoms used to make the dot.

47. How many litres of liquid CCl_4 (d = 1.5 g/cc) must be measured out to contain 1×10^{25} Cl atoms ?

48. Calculate the difference in the number of carbon atoms in 1.0 g of C-14 isotope and 1.0 g of C-12 isotope.

49. Calculate the mass of oxygen in grams present in 0.1 mole of $Na_2CO_3.10H_2O$.

Answers to Practice Problems

35. (*i*) 6.64×10^{-23} g (*ii*) 1.06×10^{-22} g

36. (*i*) 3.011×10^{23} atoms (*ii*) 6.022×10^{22} atoms (*iii*) 1.445×10^{24} atoms (*iv*) 2.409×10^{23} atoms

37. 8.624 g

38. (*i*) 1.12 L (*ii*) 0.224 L (*iii*) 4.48 L

39. 1.91×10^{10} years **40.** 1999.5 gram atoms

41. 1.506×10^{23} molecules, 4.518×10^{23} atoms.

42. 7.64×10^{20} **43.** 0.631 kg

44. 6.022×10^{21} molecules **45.** 8.07×10^{18} molecules

46. 5.019×10^{16} atoms **47.** 0.426 litres

48. 7.2×10^{21} **49.** 20.8 g.

Hints & Solutions on page 79

add on Conceptual Questions 2

CONCEPTUAL QA

Q. 1. Calculate the total number of electrons present in 1.6 g of methane.

Ans. 16 g of methane contains = 6.022×10^{23} molecules

$$1.6 \text{ g of methane contains} = \frac{6.022 \times 10^{23} \times 1.6}{16} = 6.022 \times 10^{22} \text{ molecules}$$

One molecule of CH_4 contains = 6 + 4 = 10 electrons

$$\text{No. of electrons in 1.6 g of methane} = 6.022 \times 10^{22} \times 10 = 6.022 \times 10^{23}.$$

Q.2. How many molecules of aspirin (molar mass = 180 amu) are present in 50 mg tablet ?

Ans. Molecules of aspirin = $\frac{50 \times 10^{-3}}{180} \times 6.022 \times 10^{23} = 1.673 \times 10^{20}$.

Q.3. Lithium exists in nature in the form of two isotopes, Li-6 and Li-7 with atomic masses 6.015 μ and 7.016 μ and the percentages 8.24 and 91.76 respectively. Calculate the average atomic mass of Li.

Ans. Average atomic mass of Li = $\frac{6.015 \times 8.24 + 7.016 \times 91.76}{100} = \mathbf{6.934\ \mu}$

Q. 4. What is the ratio of molecules between 1 mole of H_2O and 1 mole of sucrose ($C_{12}H_{22}O_{11}$) ?

Ans. 1 : 1

Q. 5. If atomic weight of C and S are 12 and 32 respectively, then an atom of S is .... time heavier than an atom of C.

Ans. 8/3.

Q. 6. What is the mass of a mole of water containing 50% of heavy water (D_2O) ?

Ans. 19 g

Q. 7. What is the mass of a molecule of carbon-14 dioxide ($^{14}CO_2$) ?

Ans. $= 7.64 \times 10^{-23}$ g.

Q. 8. Why do atomic masses of most of the elements in atomic mass units involve fraction ?

Ans. This is because atomic mass of an element is the average relative mass of its various isotopes. While taking average, the result comes out to be fraction.

Q.9. What is the approximate molecular mass of dry air containing 78% N_2 and 22% O_2 ?

Ans. Molecular mass = $28 \times 0.78 + 32 \times 0.22 = 28.88$.

Q.10. ^{35}Cl and ^{37}Cl are the naturally occurring isotopes of chlorine. What percentage distribution accounts for the atomic mass 35.45 ?

Ans. Let fraction of ^{35}C $= x$, $\therefore$ Fraction of ^{37}C $= 1 - x$

$\therefore$ Atomic mass $= x\,(35) + (1.0 - x)\,37 = 35.45$

$\therefore$ $x = 0.775$

$\therefore$ ^{35}Cl $= 77.5\%$ and ^{37}Cl $= 22.5\%$

Q. 11. Two bulbs A and B of equal capacity contain 10g of oxygen (O_2) and ozone (O_3) respectively. Which bulb will have

(*i*) larger number of molecules?

(*ii*) larger number of oxygen atoms?

Ans. $$10\text{g of } O_2 = \frac{10}{32}\text{ mol} = \frac{10}{32} \times 6.022 \times 10^{23}\text{ molecules} = 1.88 \times 10^{23}\text{ molecules}$$

$$\text{or} \quad = 2 \times 1.88 \times 10^{23}\text{ O-atoms} = 3.76 \times 10^{23}\text{ O-atoms}$$

$$10\text{g of } O_3 = \frac{10}{48}\text{ mol} = \frac{10}{48} \times 6.022 \times 10^{23}\text{ molecules} = 1.254 \times 10^{23}\text{ molecules}$$

$$\text{or} \quad = 3 \times 1.254 \times 10^{23}\text{ O-atoms} = 3.76 \times 10^{23}\text{ O-atoms}$$

(*i*) Bulb A contain larger number of molecules

(*ii*) Both bulbs contain the same number of O-atoms.

Q. 12. In three moles of ethane (C_2H_6) calculate the following :

(*i*) Number of moles of carbon

(*ii*) Number of moles of hydrogen atoms

(*iii*) Number of molecules of ethane

Ans. (*i*) 1 mole of C_2H_6 contains 2 moles of carbon

$\therefore$ Number of moles of carbon in 3 moles of $C_2H_6 = 6$

(*ii*) 1 mole of C_2H_6 contains 6 moles atoms of hydrogen

$\therefore$ Number of moles of hydrogen atoms in 3 mole of $C_2H_6 = 3 \times 6 = 18$

(*iii*) 1 mole of $C_2H_6 = 6.022 \times 10^{23}$ molecules

$\therefore$ Number of molecules in 3 moles of $C_2H_6 = 3 \times 6.022 \times 10^{23}$

$= \mathbf{1.807 \times 10^{24}}$ **molecules**

Q. 13. Which one of the following will have largest number of atoms ?

(*i*) 1 g Au (*s*) (*ii*) 1 g Na (*s*) (*iii*) 1 g Li (*s*) (*iv*) 1 g Cl_2 (*g*)

(at. mass : Au = 197, Na = 23, Li = 7, Cl = 35.5) ***(NCERT)***

Ans. No. of atoms can be calculated as :

$$(i)\ 1\text{ g Au} = \frac{1}{197} \times 6.022 \times 10^{23} = \frac{6.022}{197} \times 10^{23}$$

$$(ii)\ 1\text{ g Na} = \frac{6.022}{23} \times 10^{23}$$

CONCEPTUAL QA

(*iii*) 1 g Li $= \dfrac{6.022 \times 10^{23}}{7}$

(*iv*) 1 g Cl_2 $= \dfrac{2 \times 6.022 \times 10^{23}}{71} = \dfrac{6.022 \times 10^{23}}{35.5}$

It is clear that 1 g Li contains largest number of atoms.

Q.14. Calculate the average atomic mass of chlorine using the following data.

	% Natural abundance	Molar mass
^{35}Cl	**75.77**	**34.9689**
^{37}Cl	**24.23**	**36.9659**

Ans. Average atomic mass $= \dfrac{34.9689 \times 75.77 + 36.9659 \times 24.23}{75.77 + 24.23}$

$= 26.49 + 8.96$

$=$ **35.45**

Advanced Level PROBLEMS
Accelerate Your Potential (for JEE Advance)

Problem 1. An alloy of metals X and Y weighs 12 g and contains atoms X and Y in the ratio of 2 : 5. The percentage by mass of X in the sample is 20. If atomic mass of X is 40, what is the atomic mass of metal Y ?

Solution Mass of metal X in alloy $= \dfrac{12 \times 20}{100} = 2.40$ g

Mass of metal Y in alloy $= 12 - 2.4 = 9.6$ g

Number of atoms of X $= \dfrac{6.022 \times 10^{23} \times 2.4}{40}$

$= 3.61 \times 10^{22}$

Ratio of atoms of X and Y = 2 : 5

No. of atoms of Y $= \dfrac{3.61 \times 10^{22} \times 5}{2}$

$= 9.025 \times 10^{22}$ atoms

Now, 9.025×10^{22} atoms of Y are present in 9.6 g

6.022×10^{23} atoms of Y are present in

$\dfrac{9.6}{9.025 \times 10^{22}} \times 6.022 \times 10^{23}$

$= 64.0$ g

$\therefore$ Atomic mass of Y = **64 a.m.u.**

Problem 2. Calculate (*i*) the total number of neutrons and (*ii*) total mass of neutrons in 7 mg of C^{14} (assume mass of neutron = mass of hydrogen atom).

Solution 1 mol of C^{14} = 14 g

or 14 g of C^{14} = 1 mol

7×10^{-3} g of C^{14} $= \dfrac{1}{14} \times 7 \times 10^{-3} = 0.5 \times 10^{-3}$ mol

Now, 1 mol of C^{14} contains $= 6.022 \times 10^{23}$ atoms

0.5×10^{-3} mol of C^{14} contains

$= 6.022 \times 10^{23} \times 0.5 \times 10^{-3}$

$= 3.011 \times 10^{20}$ atoms

No. of electrons in C^{14} = 6

No. of protons in 1 atom of C^{14} = 6

$\therefore$ No. of neutrons in 1 atom of C^{14} = 14 – 6 = 8

(*i*) 3.011×10^{20} atoms of C^{14} contain

$= 8 \times 3.011 \times 10^{20}$

$=$ **2.409×10^{21} neutrons**

(*ii*) Mass of one neutron = Mass of H atom

$= \dfrac{1}{6.022 \times 10^{23}}$

Mass of 2.409×10^{21} neutrons

$= \dfrac{1}{6.022 \times 10^{23}} \times 2.409 \times 10^{21}$

$= 4 \times 10^{-3}$ g = **4 mg.**

Problem 3. Calculate the number of molecules of carbon dioxide present in 300 mL of gas at 273K and 2.5 atm pressure.

Solution First of all, we have to calculate the volume of gas at STP by applying gas equation.

Given conditions	At STP
$V_1 = 300$ mL	$V_2 = ?$
$T_1 = 273$ K	$T_2 = 273$ K
$p_1 = 2.5$ atm	$p_2 = 1$ atm

Applying gas equation, $\dfrac{p_1V_1}{T_1} = \dfrac{p_2V_2}{T_2}$

$\dfrac{2.5 \times 300}{273} = \dfrac{1 \times V_2}{273}$

$\therefore$ $V_2 = 2.5 \times 300 = 750$ mL

We know that 1 mol of a gas at STP occupies 22400 mL and contains 6.022×10^{23} molecules

$\therefore$ 750 mL of CO_2 at STP will contain

$$= \frac{6.022 \times 10^{23}}{22400} \times 750$$

$$= \mathbf{2.01 \times 10^{22}} \textbf{ molecules}$$

Problem 4. If atomic mass of carbon was taken as 100 u, then what would be the value of Avogadro's number?

Solution Avogadro number would be the number of atoms in 100 g of carbon.

12 g of carbon contain 6.022×10^{23} atoms

then 100 g of carbon would contain atoms

$$= \frac{6.022 \times 10^{23} \times 100}{12}$$

$$= \mathbf{5.01 \times 10^{24}} \textbf{ atoms}$$

Problem 5. A 0.005 cm thick coating of copper is deposited on a plate of 0.5 m² total area. Calculate the number of copper atoms deposited on the plate (density of copper = 7.2 g cm⁻³, atomic mass = 63.5).

Solution Area of plate = $0.5 \text{ m}^2 = 0.5 \times 10^4 \text{ cm}^2$

Thickness of coating = 0.005 cm

Volume of copper deposited = $0.5 \times 10^4 \times 0.005$

= 25 cm^3

Mass of copper deposited = $25 \times 7.2 = 180$ g

Now, 63.5 g of copper contain atoms = 6.022×10^{23}

$$180 \text{ g of copper contain atoms} = \frac{6.022 \times 10^{23}}{63.5} \times 180$$

$$= \mathbf{1.71 \times 10^{24}} \textbf{ atoms}$$

Problem 6. Calculate the number of molecules present in a spherical drop of water having a radius 1 mm if density of water is 1 g cm⁻³.

Solution Since water molecule is assumed to be spherical, its volume is equal to $\frac{4}{3}\pi r^3$ where r is the radius of water molecule.

Radius = 1 mm = 0.1 cm

$$\text{Volume} = \frac{4}{3}\pi(0.1)^3$$

$$= \frac{4}{3} \times 3.143 \times 0.001 = 0.00419 \text{ cm}^3$$

Mass of drop of water = Volume × Density

$= 0.00419 \text{ cm}^3 \times 1 \text{ g cm}^{-3} = 0.0419$ g

Now, 1 mole of water = 18 g of water contain 6.022×10^{23} molecules.

$\therefore$ 0.00419 g of water contain molecules

$$= \frac{6.022 \times 10^{23}}{18} \times 0.00419$$

$$= \mathbf{1.40 \times 10^{20}}$$

Problem 7. 1×10^{21} molecules are removed from 280 mg of carbon monoxide. Calculate the number of moles of carbon monoxide left.

Solution **Step I.** *To calculate the number of moles of carbon monoxide in 280 mg.*

$$\text{No. of moles of CO} = \frac{\text{Wt. of CO in grams}}{\text{Gram molecular mass of CO}}$$

$$= \frac{280}{1000} \times \frac{1}{28} = 0.01 \text{ mol.}$$

Step II. *To calculate the number of moles of carbon monoxide left.*

6.022×10^{23} molecules of CO = 1 mol.

1×10^{21} molecules of CO

$$= \frac{1 \times 10^{21}}{6.022 \times 10^{23}} = 0.0016 \text{ mol.}$$

No. of moles of CO left =

No. of moles of CO taken – No. of moles of CO removed

= 0.01 – 0.0016 = 0.0084

= $\mathbf{8.4 \times 10^{-3}}$ **mol.**

PERCENTAGE COMPOSITION AND MOLECULAR FORMULA

To determine the molecular formula of a compound, we begin by measuring the mass or mass percentage of each element present in the compound. The composition is generally expressed as the **mass percentage composition.**

Mass percentage

gives the mass of each element expressed as the percentage of the total mass.

In other words, it gives the number of grams of the element present in 100 g of the compound. It is obtained by dividing the mass of each element in 1 mole of the compound by molar mass of the compound and multiplying by 100. Thus,

Mass percentage of an element

$$= \frac{\text{Mass of that element in 1 mole compound}}{\text{Molar mass of compound}} \times 100$$

For example, let us calculate the mass percentage composition of glucose ($C_6H_{12}O_6$)

The formula of glucose = $C_6H_{12}O_6$

Molar mass of glucose = $6 \times 12 + 12 \times 1 + 6 \times 16 = 180$

The formula of glucose shows that there are 6 C atoms, 12 H atoms and 6 O atoms.

Mass of C = $6 \times 12 = 72$

$$\text{Mass percentage of C} = \frac{72}{180} \times 100 = \mathbf{40.0\%}$$

Mass of H = $12 \times 1 = 12$

$$\text{Mass percenatage of H} = \frac{12}{180} \times 100 = \mathbf{6.67\%}$$

Mass of O = 6 × 16 = 96

$$\text{Mass percentage of O} = \frac{96}{180} \times 100 = \mathbf{53.33\%}$$

Similarly, mass percentage of ammonia (NH_3) can be calculated as :

Molar mass of NH_3 = 14 + 3 = 17

The formula of NH_3 shows that there are 1 N atom and 3 H atoms

$$\text{Mass percentage of N} = \frac{\text{Mass of N in 1 mol of } NH_3}{\text{Molar mass of } NH_3} \times 100$$

$$= \frac{14}{17} \times 100 = 82.35\%$$

$$\text{Mass percentage of H} = \frac{\text{Mass of H in 1 mol of } NH_3}{\text{Molar mass of } NH_3} \times 100$$

$$= \frac{3}{17} \times 100 = 17.65\%$$

Let us consider one more example of ethanol having the molecular formula C_2H_5OH. Molar mass of ethanol = 2 × 12 + 6 × 1 × 16 = 46 g

Mass of C = 2 × 12 = 24

$$\text{Mass \% of C} = \frac{24}{46} \times 100 = 52.17\%$$

Mass of H = 6 × 1

$$\text{Mass \% of H} = \frac{6}{46} \times 100 = 13.04\%$$

Mass of O atom = 16.0

$$\text{Mass \% of O} = \frac{16}{46} \times 100 = 34.78\%$$

SOLVED EXAMPLES

Example 54.

Calculate the mass percentage composition of copper pyrites ($CuFeS_2$).

Solution: Molar mass of $CuFeS_2$

= 63.5 + 55.8 + 2 × 32 = 183.3

Mass of Cu = 63.5

$$\text{Mass percentage of Cu} = \frac{63.5}{183.3} \times 100 = 34.64\%$$

Mass of iron = 55.8

$$\text{Mass percentage of iron} = \frac{58.5}{183.3} \times 100 = 30.44\%$$

Mass of sulphur = 2 × 32 = 64

$$\text{Mass percentage of sulphur} = \frac{64}{183.3} \times 100$$

$$= \mathbf{34.91\%}$$

Example 55.

Calculate the percentage composition of the following compounds:

(i) Urea $CO(NH_2)_2$

(ii) Copper sulphate $CuSO_4.5H_2O$.

Solution: (*i*) Urea : $CO(NH_2)_2$

Molar mass of urea = 12 + 16 + 2 (14 + 2) = 60

Mass of carbon = 12

$$\text{Mass percentage of carbon} = \frac{12}{60} \times 100 = 20\%$$

Mass of oxygen = 16

$$\text{Mass percentage of oxygen} = \frac{16}{60} \times 100 = 26.67\%$$

Mass of nitrogen = 2 × 14 = 28

$$\text{Mass percentage of nitrogen} = \frac{28}{60} \times 100 = 46.66\%$$

Mass of H = 4 × 1 = 4

$$\text{Mass percentage of hydrogen} = \frac{4}{60} \times 100 = 6.67\%$$

(*ii*) Copper sulphate, $CuSO_4.5H_2O$

Molar mass of $CuSO_4$. $5H_2O$ = 63.5 + 32 + 4 × 16 + 5 × 18
= 249.5

Mass of Cu = 63.5

$$\text{Mass percentage of Cu} = \frac{63.5}{249.5} \times 100 = 25.45$$

Mass of S = 32

$$\text{Mass percentage of S} = \frac{32}{249.5} \times 100 = 12.82\%$$

Mass of O = 9 × 16 = 144

$$\text{Mass percentage of O} = \frac{144}{249.5} \times 100 = 57.77\%$$

Mass of H = 10 × 1 = 10

$$\text{Mass percentage of H} = \frac{10}{249.5} \times 100 = 4.01\%$$

Example 56.

Ferric sulphate is used in water and sewage treatment and in removal of suspended impurities. Its empirical formula is $Fe_2(SO_4)_3$. Calculate the mass percentage of iron, sulphur and oxygen in this compound.

Solution: Molar mass of $Fe_2(SO_4)_3$

= 2 × 56 + 3 (32 + 4 × 16)

= 400

Mass of iron = 2 × 56 = 112

$$\text{Mass percentage of iron} = \frac{2 \times 56}{400} \times 100 = 28\%$$

Mass of sulphur = 32 × 3 = 96

$$\text{Mass percentage of sulphur} = \frac{96}{400} \times 100 = 24\%$$

Mass of oxygen = 3 × 64 = 192

$$\text{Mass percentage of oxygen} = \frac{192}{400} \times 100 = 48\%$$

❑ **Example 57.**

Calculate the percentage of water of crystallisation in the sample of Mohr salt, $FeSO_4(NH_4)_2SO_4 \cdot 6H_2O$.

Solution: Molar mass of $FeSO_4 \cdot (NH_4)_2SO_4 \cdot 6H_2O$

$$= 56 + (32 + 64) + 2\,(14 + 4) + (32 + 64) + 6\,(2 \times 1 + 16)$$

$$= 56 + 96 + 36 + 96 + 108 = 392$$

Mass of water = 6 × 18 = 108

$$\text{Percentage of water} = \frac{108}{392} \times 100 = 27.55\%$$

∴ Water of crystallisation = **27.55%**.

Empirical and Molecular Formula

The chemical formula may be of two types :

(*i*) *Empirical formula* and

(*ii*) *Molecular formula.*

Empirical formula.

The formula which gives the simplest whole number ratio of the atoms of various elements present in one molecule of the compound is called **empirical formula.**

For example, empirical formula of hydrogen peroxide is HO. It represents that hydrogen and oxygen (*i.e.,* H : O) are present in the ratio of 1 : 1 in hydrogen peroxide. Similarly, empirical formula of benzene is CH which indicates that atomic ratio of C : H in benzene is 1 : 1.

Molecular formula

The formula which gives the actual number of atoms of various elements present in one molecule of the compound is called **molecular formula.**

For example, the molecular formula of hydrogen peroxide is H_2O_2, because one molecule of hydrogen peroxide contains two atoms of hydrogen and two atoms of oxygen. Similarly, molecular formula of benzene is C_6H_6 which tells that one molecule of benzene contains 6 atoms of carbon and 6 atoms of hydrogen.

Empirical and molecular formulae of some molecules are given below :

Compound	Empirical formula	Molecular formula
Hydrogen peroxide	HO	H_2O_2
Benzene	CH	C_6H_6
Glucose	CH_2O	$C_6H_{12}O_6$
Sucrose	$C_{12}H_{22}O_{11}$	$C_{12}H_{22}O_{11}$
Naphthalene	C_5H_4	$C_{10}H_8$

❑ **Example 58.**

Write the empirical formula of the compounds having the molecular formulae :

(i) C_6H_6 (ii) C_6H_{12} (iii) H_2O_2 (iv) Na_2CO_3 (v) B_2H_6 (vi) N_2O_4 (vii) H_3PO_4 (viii) Fe_2O_3 (ix) C_2H_2 (x) N_2O_5

Solution: Empirical formula gives the simple whole number ratio of atoms in one molecule of the compound.

(*i*)	C_6H_6	= CH	(*ii*)	C_6H_{12}	= CH_2
(*iii*)	H_2O_2	= HO	(*iv*)	Na_2CO_3	= Na_2CO_3
(*v*)	B_2H_6	= BH_3	(*vi*)	N_2O_4	= NO_2
(*vii*)	H_3PO_4	= H_3PO_4	(*viii*)	Fe_2O_3	= Fe_2O_3
(*ix*)	C_2H_2	= CH	(*x*)	N_2O_5	= N_2O_5

Relation between Empirical and Molecular Formulae

Molecular formula and empirical formula are related as

Molecular formula = *n* (Empirical formula)

where n is a simple whole number and may have values 1, 2, 3... It is equal to

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

For example, the molecular mass of benzene is 78. The empirical formula of benzene is CH and therefore, its empirical formula mass is 13.

Thus,
$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

$$= \frac{78}{13} = 6$$

Therefore, molecular formula of benzene = 6(CH) = C_6H_6.

It may be noted that in many cases, the value of n comes out to be one and, therefore, empirical formula and molecular formula are same in these cases. For example, empirical and molecular formulae are same for $CuSO_4$, NaCl, $K_2Cr_2O_7$, NH_3, etc.

Determination of the Empirical Formula of a Compound

The empirical formula of a compound can be determined from the *percentage composition of different elements and atomic masses of the elements.*

The **various steps** involved in determining the empirical formula are :

Step I. *Divide the percentage of each element by its atomic mass. This gives the moles of atoms of various elements in the molecule of the compound.*

$$\text{Moles of atoms} = \frac{\text{Percentage of an element}}{\text{Atomic mass of the element}}$$

Step II. *Divide the result obtained in the above step by the smallest value among them to get the simplest ratio of various atoms.*

Step III. *Make the values obtained above to the nearest whole number and multiply, if necessary, by a suitable integer to make the values whole numbers*. This gives the **simplest whole number ratio.**

Step IV. *Write the symbols of the various elements side by side and insert the numerical value at the right hand lower corner of each symbol*. The formula thus obtained represents the empirical formula of the compound.

Steps for Determination of the Molecular Formula of a Compound

Step I. *Determine the empirical formula as described above.*

Step II. *Calculate the empirical formula mass by adding the atomic masses of the atoms in the empirical formula.*

Step III. *Determine the molecular mass by a suitable method.*

Step IV. *Determine the value of* n *as*

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

Change n to the nearest whole number.

Step V. *Multiply empirical formula by* n *to get the molecular formula.*

Molecular formula = n × *Empirical formula.*

SOLVED EXAMPLES

Example 59.

The molecular mass of an organic compound is 78 and its percentage composition is 92.4% C and 7.6% H. Determine the molecular formula of the compound.

Solution: ***Calculation of empirical formula.***

Element	Percentage	Atomic mass	Moles of atoms	Mole ratio or atomic ratio	Simplest whole no. ratio
C	92.4	12.0	$\frac{92.4}{12.0} = 7.7$	$\frac{7.7}{7.6} = 1.01$	1
H	7.6	1	$\frac{7.6}{1} = 7.6$	$\frac{7.6}{7.6} = 1$	1

The simplest whole number ratio of C : H is 1 : 1

∴ The empirical formula of the compound is CH.

Calculation of molecular formula of the compound.

Empirical formula mass = 1 × 12 + 1 × 1 = 13

Molecular mass = 78

$$\therefore \quad n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{78}{13} = 6$$

Thus, molecular formula of the compound = 6 × (CH) = $\mathbf{C_6H_6}$.

Example 60.

An organic compound on analysis gave the following percentage composition : C = 57.8%, H = 3.6% and the rest is oxygen. The vapour density of the compound was found to be 83. Find out the molecular formula of the compound.

Solution: Calculation of empirical formula.

Element	Percentage	Atomic mass	Moles of atoms	Mole ratio or atomic ratio	Simplest whole no. ratio
C	57.8	12	$\frac{57.8}{12} = 4.82$	$\frac{4.82}{2.41} = 2$	4
H	3.6	1	$\frac{3.6}{1} = 3.60$	$\frac{3.60}{2.41} = 1.49$	3
O	100 – (57.8 + 3.6) = 38.6	16	$\frac{38.6}{16} = 2.41$	$\frac{2.41}{2.41} = 1$	2

$\therefore$ Empirical formula = $C_4H_3O_2$

Calculation of molecular formula

Empirical formula mass = $4 \times 12 + 3 \times 1 + 2 \times 16 = 83$

Molecular mass = $2 \times$ V.D. = $2 \times 83 = 166$

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

$$= \frac{166}{83} = 2$$

Molecular formula = n (Empirical formula)

$= 2\,(C_4H_3O_2) = \mathbf{C_8H_6O_4}$.

Example 61.

Four gram of copper chloride on analysis was found to contain 1.890 g of copper (Cu) and 2.110 of chlorine (Cl). What is the empirical formula of copper chloride ?

Solution:

Step I. ***Calculation of percentage of different elements.***

$$\text{Percentage of Cu in copper chloride} = \frac{\text{Wt. of copper}}{\text{Wt. of copper chloride}} \times 100$$

$$= \frac{1.890}{4.000} \times 100 = 47.3$$

$$\text{Percentage of Cl in copper chloride} = \frac{\text{Weight of chlorine}}{\text{Weight of copper chloride}} \times 100$$

$$= \frac{2.110}{4.000} \times 100 = 52.7$$

Step II. ***Calculation of empirical formula.***

Element	Percentage	Atomic mass	Moles of atoms	Mole ratio or atomic ratio	Simplest whole no. ratio
Cu	47.3	63.5	$\frac{47.3}{63.5} = 0.74$	$\frac{0.74}{0.74} = 1$	1
Cl	52.7	35.5	$\frac{52.7}{35.5} = 1.48$	$\frac{1.84}{0.74} = 2$	2

The simplest whole number ratio of various atoms is :

Cu : Cl as 1 : 2

$\therefore$ The empirical formula of the compound is $CuCl_2$.

Example 62.

A compound contains 4.07 % hydrogen, 24.27% carbon and 71.65% chlorine. Its molecular mass is 98.96. What are its empirical and molecular formulae ? **N.C.E.R.T.**

Solution: Step I. ***Calculation of empirical formula***

Elements	Percentage composition	At. mass	Moles of atoms	Mole ratio or atomic ratio
C	24.27	12	$\frac{24.27}{12} = 2.02$	$\frac{2.02}{2.02} = 1$
H	4.07	1	$\frac{4.07}{1} = 4.07$	$\frac{4.07}{2.02} = 2$
Cl	71.65	35.5	$\frac{71.65}{35.5} = 2.02$	$\frac{2.02}{2.02} = 1$

∴ Empirical formula of the compound = CH_2Cl

Step II. ***Calculation of molecular formula***

$$\text{Empirical formula mass} = \text{At. mass of C} + 2 \times \text{At. mass of H} + \text{At. mass of Cl}$$
$$= 12 + 2 \times 1 + 35.5 = 49.5$$
$$\text{Molecular mass} = 98.96$$
$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{98.96}{49.5} = 2.$$

∴ Molecular formula = $(CH_2Cl)_2 = \mathbf{C_2H_4Cl_2}$.

Example 63.

Determine the empirical formula of an oxide of iron which has 69.9% iron and 30.1% oxygen by mass. **N.C.E.R.T.**

Solution: Step I. Calculation of empirical formula

Elements	At. mass	Percentage composition	Moles of atoms	Mole ratio or atomic ratio	Simplest whole number ratio
Fe	56	69.9	$\frac{69.9}{56} = 1.25$	$\frac{1.25}{1.25} = 1$	2
O	16	30.1	$\frac{30.1}{16} = 1.88$	$\frac{1.88}{1.25} = 1.50$	3

Empirical formula : $\mathbf{Fe_2O_3}$.

Example 64.

(a) Butyric acid contains only C, H and O. A 4.24 mg sample of butyric acid is completely burnt. It gives 8.45 mg of carbon dioxide and 3.46 mg of water. What is the mass percentage of each element in butyric acid ?

(b) The molecular mass of butyric acid was determined by experiment to be 88u. What is the molecular formula ?

Solution: (*a*) Calculation of mass percentage of different elements

Percentage of carbon can be calculated as :

$$\underset{44\text{ mg}}{CO_2} \equiv \underset{12\text{ mg}}{C}$$
$$44 \text{ mg of } CO_2 \text{ contain C} = 12 \text{ mg}$$
$$8.45 \text{ mg of } CO_2 \text{ contain C} = \frac{12}{44} \times 8.45 \text{ mg}$$
$$\therefore \quad \text{Percentage of C} = \frac{\text{Weight of carbon}}{\text{Weight of compound}} \times 100$$

$$= \frac{12}{44} \times \frac{8.45}{4.24} \times 100 = 54.3\%$$

Percentage of hydrogen can be calculated as :

$$H_2O \equiv 2H$$
$$18 \text{ mg} \quad 2 \text{ mg}$$

$$18 \text{ mg of } H_2O \text{ contain H} = 2 \text{ mg}$$

$$3.46 \text{ mg of } H_2O \text{ contain H} = \frac{2}{18} \times 3.46 \text{ mg}$$

$$\therefore \quad \text{Percentage of H} = \frac{\text{Weight of hydrogen}}{\text{Weight of compound}} \times 100$$

$$= \frac{2}{18} \times \frac{3.46}{4.24} \times 100 = 9.0\%$$

$$\text{The sum of the percentages of C and H} = 54.3 + 9.0 = 63.3\%$$

$$\therefore \quad \text{Percentage of O} = 100 - 63.3 = 36.7\%$$

(*b*) Calculation of molecular formula

Element	Percentage	Atomic mass	Moles of atoms	Mole ratio or atomic ratio	Simplest whole no. ratio
C	54.3	12.0	$\frac{54.3}{12.0} = 4.52$	$\frac{4.52}{2.29} = 1.97$	2
H	9.0	1.008	$\frac{9.0}{1.008} = 8.93$	$\frac{8.93}{2.29} = 3.90$	4
O	36.7	16.0	$\frac{36.7}{16.0} = 2.29$	$\frac{2.29}{2.29} = 1.00$	1

The simplest whole number ratio of atom is :

$$C : H : O = 2 : 4 : 1$$

$\therefore$ The empirical formula is C_2H_4O

$$\text{Empirical formula mass} = 2 \times 12 + 4 \times 1 + 16$$
$$= 44 \text{ a.m.u.}$$
$$\text{Molecular mass} = 88 \text{ a.m.u.}$$

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

$$n = \frac{88}{44} = 2$$

$\therefore$ Molecular formula of butyric acid $2(C_2H_4O)$ = $\mathbf{C_4H_8O_2}$.

❑ **Example 65.**

A compound on analysis gave the following percentage composition.
Na = 14.31%, S = 9.97%, H = 6.22% and O = 69.5%.
Calculate the molecular formula of the compound on the assumption that all the hydrogens in the compound are present in combination with oxygen as water of crystallisation. The molecular mass of the compound is 322. [At. mass : Na = 23, S = 32, H = 1, O = 16].

Solution:

Step I. Calculation of empirical formula

Element	Atomic mass	Percentage composition	Moles of atoms	Mole ratio or atomic ratio
Na	23	14.31	$\frac{14.31}{23} = 0.62$	$\frac{0.62}{0.31} = 2$
S	32	9.97	$\frac{9.97}{32} = 0.31$	$\frac{0.31}{0.31} = 1$
H	1	6.22	$\frac{6.22}{1} = 6.22$	$\frac{6.22}{0.31} = 20$
O	16	69.50	$\frac{69.50}{16} = 4.34$	$\frac{4.34}{0.31} = 14$

Empirical formula of the compound is $Na_2SH_{20}O_{14}$.

Step II. Calculation of molecular formula

Empirical formula mass

$$= 2 \times \text{At. mass of Na} + \text{At. mass of S} + 20 \times \text{At. mass of H} + 14 \times \text{At. mass of O}$$

$$= 2 \times 23 + 1 \times 32 + 20 \times 1 + 14 \times 16 = 322$$

$$\text{Molecular mass} = 322$$

Now, $$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{322}{322} = 1$$

$$\text{Molecular formula} = (Na_2SH_{20}O_{14}) = Na_2SH_{20}O_{14}$$

Since all hydrogen atoms are present in combination with oxygen as water therefore, it means that all the 20 hydrogen atoms must be combined with 10 atoms of oxygen to form 10 molecules of water of crystallisation. The remaining 4 oxygen atoms are present with the compound. Thus,

Molecular formula of the compound = $\mathbf{Na_2SO_4.10H_2O}$

❑ **Example 66.**

A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g of carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at S.T.P.) of this welding gas is found to weigh 11.6 g. Calculate (i) empirical formula (ii) molar mass of the gas, and (iii) molecular formula.

Solution: $\underset{44\text{ g}}{CO_2} \equiv \underset{12\text{ g}}{C}$

$$\text{Mass of carbon} = \frac{12}{44} \times 3.38 = \mathbf{0.92\ g}$$

$$\underset{18\text{ g}}{H_2O} \equiv \underset{2\text{ g}}{H_2}$$

$$\text{Mass of hydrogen} = \frac{2}{18} \times 0.690 = \mathbf{0.077}$$

$$\text{Percentage of C} = \frac{0.922}{(0.922 + 0.077)} \times 100 = \mathbf{92.2\%}$$

$$\text{Precentage of H} = \frac{0.077}{(0.922 + 0.077)} \times 100 = \mathbf{7.7\%}$$

***(i)* Calculation of empirical formula**

Element	Percentage of element	Atomic mass	Moles of atoms	Mole ratio
C	92.2	12	$\frac{92.2}{12} = 7.7$	$\frac{7.7}{7.7} = 1$
H	7.77	1	$\frac{7.77}{1} = 7.7$	$\frac{7.7}{7.7} = 1$

Empirical formula = CH

***(ii)* Calculation of molar mass**

$$10.0 \text{ L of gas at S.T.P. weigh} = 11.6 \text{ g}$$

$$22.4 \text{ L of gas at S.T.P. weigh} = \frac{11.6}{10.0} \times 22.4 = \mathbf{26\ g\ mol^{-1}}.$$

***(iii)* Calculation of molecular formula**

Empirical formula mass = 12 + 1 = 13

$$\text{Molecular mass} = 26$$

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{26}{13} = \mathbf{2}$$

$\therefore$ Molecular formula = 2 (CH) = $\mathbf{C_2H_2}$

Example 67.

A crystalline salt on being rendered anhydrous loses 45.6% of its weight. The composition of anhydrous salt is : Al = 10.5%, K = 15.1%, S = 24.8% and O = 49.6%. Calculate the formula of anhydrous salt and crystalline salt (At. mass of Al = 27, K = 39, S = 32, O = 16).

Solution:

Step I. Calculation of empirical formula

Element	Percentage of element	Atomic mass	Moles of atoms	Mole ratio	Simplest whole no. ratio
K	15.1	39	$\frac{15.1}{39} = 0.38$	$\frac{0.38}{0.38} = 1$	1
Al	10.5	27	$\frac{10.5}{27} = 0.38$	$\frac{0.38}{0.38} = 1$	1
S	24.8	32	$\frac{24.8}{32} = 0.775$	$\frac{0.775}{0.38} = 2.04$	2
O	49.6	16	$\frac{49.6}{16} = 3.1$	$\frac{3.1}{0.38} = 8.1$	8

The simplest atomic ratio of K : Al : S : O is 1 : 1: 2 : 8. The empirical formula of the anhydrous salt is : $KAlS_2O_8$.

Step II. *To calculate the empirical formula mass of anhydrous salt.*

Empirical formula mass of anhydrous salt

= 39 + 27 + 64 + 128 = 258.

Step III. *To calculate the amount of water associated with 258 parts by weight of compound.*

Let the weight of hydrated salt be 100 g

$\therefore$ The weight of water lost = 45.6 g

$\therefore$ The weight of anhydrous salt = 100 – 45.6 g = 54.4 g

If the weight of anhydrous salt is 54.4 gm, the weight of water in the crystalline salt = 45.6 gm.

If the weight of anhydrous salt in 258 gm, the weight of water in the crystalline salt

$$= \frac{45.6}{54.4} \times 258 = 216.2 \text{ g}$$

Step IV. *To calculate the number of molecules of water of crystallisation.*

It is clear from the above calculations that 216.2 parts by weight of water are present in combination with 258 parts be weight of the salt.

$\therefore$ The number of molecules of water of crystallisation corresponding to 216.2 parts by weight of water

$$= \frac{216.2}{18} = 12$$

$\therefore$ The simplest formula of crystalline salt is $\mathbf{KAlS_2O_8.12H_2O}$.

Practice Problems

50. (*i*) Calculate the mass percentage of various elements present in magnesium sulphate, $MgSO_4$.
(*ii*) Calculate the percentage of cation in ammonium dichromate.

51. An organic compound containing carbon, hydrogen and oxygen gave the following percentage composition :
C = 40.68%, H = 5.08%
The vapour density of the compound is 59. Calculate the molecular formula of the compound.

52. (*i*) The elemental composition of butyric acid was found to be 54.2% C, 9.2% H and 36.6% O. Determine its empirical formula.
(*ii*) The molecular mass of butyric acid was determined by an experiment to be 88 u. What is its molecular formula ?

53. An oxide of nitrogen contains 30.43% of nitrogen. The molecular weight of the compound is equal to 92 a.m.u. Calculate the molecular formula of the compound.

54. Calculate the empirical and molecular formula of the compound having the following percentage composition :
Na = 36.5%, H = 0.8%, P = 24.6%, O = 38.1%
The molecular mass of the compound is 126 a.m.u. Also name the compound.

55. A crystalline salt when heated becomes anhydrous and loses 51.2% of its weight. The anhydrous salt on analysis gave the following percentage composition :
Mg = 20.0%, S = 26.66%, O = 53.33%
Calculate the molecular formula of the anhydrous salt and the crystalline salt. Molecular weight of the anhydrous salt is 120.

56. Calculate the empirical formula of a mineral which has the following percentage composition :
CuO = 44.82%, SiO_2 = 34.83% and water = 20.35%
(at. wt. of Cu = 63.5, Si = 28).

57. Determine the empirical formula of a compound having percentage composition as :
Iron = 20% ; sulphur = 11.5% ; oxygen = 23.1% and water molecules = 45.4%
(At. mass of Fe = 56, S = 32)
* **Hint.** Treat water molecules as one entity.

Answers to Practice Problems

50. (*i*) Mg = 20 %, S = 26.67 %, O = 53.33 %
(*ii*) 14.29%

51. $C_4H_6O_4$

52. (*i*) C_2H_4O (*ii*) $C_4H_8O_2$

53. N_2O_4

54. The empirical and molecular formula is, Na_2HPO_3, sodium hydrogen phosphite.

55. $MgSO_4 . 7H_2O$

56. $CuO. SiO_2. 2H_2O$

57. $FeSO_4 . 7H_2O$

Hints & Solutions on page 79

STOICHIOMETRY OF CHEMICAL REACTIONS

The term *stoichiometry means the quantitative relationships among the reactants and the products in a reaction.* Stoichiometry is derived from the Greek words *stoicheion* meaning element and *metron* meaning measure. It is based on the fact that *the stoichiometric coefficients in a chemical equation can be interpreted as the number of moles of each substance (reactants or products).* For example, consider the reaction between nitrogen and hydrogen to form ammonia as :

$$N_2(g) \quad + \quad 3H_2(g) \longrightarrow 2NH_3(g)$$

The stoichiometric coefficients in the equation indicate that 1 molecule of nitrogen combines with 3 molecules of hydrogen to form 2 molecules of ammonia.

$$\underset{\text{1 molecule}}{N_2(g)} \quad + \quad \underset{\text{3 molecules}}{3H_2(g)} \longrightarrow \underset{\text{2 molecules}}{2NH_3(g)}$$

Multiplying by 6.022×10^{23} the entire equation, we get

$$\underset{\text{molecules}}{1 \times 6.022 \times 10^{23}} \quad \underset{\text{molecules}}{3 \times 6.022 \times 10^{23}} \quad \underset{\text{molecules}}{2 \times 6.022 \times 10^{23}}$$

This means

1 mol 3 mol 2 mol

Taking molar masses into consideration

28.0 g $3 \times 2 = 6.0$ g $2 \times 17 = 34$ g

The interpretation of the coefficients as the number of moles is the basis of all stoichiometric calculations.

Thus, to carry out stoichiometric calculations, the first step is writing the chemical equation for the reaction and then balancing it. The balanced chemical equation gives the stoichiometric coefficients which give the proportion by moles.

Chemical Equations

A chemical equation is a statement of a chemical reaction in terms of the symbols and formulae of the species involved in the reaction. The **chemical equation** may be defined as :

a brief representation of a chemical change in terms of symbols and formulae of substances involved in it.

For example, the reaction of silver nitrate with sodium chloride to give silver chloride and sodium nitrate may be represented as

$$\underbrace{AgNO_3 + NaCl}_{\text{Reactants}} \longrightarrow \underbrace{AgCl + NaNO_3}_{\text{Products}}$$

The substances which react among themselves to bring about the chemical changes are known as **reactants** whereas, the substances which are produced as a result of the chemical change, are known as **products.**

Essentials of Chemical Equation

A chemical equation must fulfil the following conditions :

(*i*) *It must be consistent with the experimental facts. i.e.,* a chemical equation must represent a true chemical change. If a chemical reaction is not possible between certain substances it cannot be represented by a chemical equation.

(*ii*) *It should be balanced i.e.,* the total number of atoms an both sides of the equation must be equal.

(*iii*) *It should be molecular i.e.,* the elementary gases like hydrogen, oxygen, nitrogen, etc. must be represented in molecular form as H_2, O_2, N_2, etc.

Information Conveyed by a Chemical Equation

A chemical equation conveys both qualitative and quantitative informations.

Qualitatively, *a chemical equation tells the names of the various reactants and the products.*

Quantitatively, *a chemical equation represents :*

(*i*) the relative number of reactant and product species (atoms or molecules) taking part in the reaction.

(*ii*) *the relative number of moles of the reactants and products.*

(*iii*) *the relative masses of the reactants and products.*

(*iv*) *the relative volumes of gaseous reactants and products.*

Thus, the chemical equation for the reaction :

$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O\ (g)$

gives the following information :

(*i*) **One molecule** of $CH_4(g)$ reacts with **2 molecules** of $O_2(g)$ to give **one molecule** of $CO_2(g)$ and **2 molecules** of $H_2O(g)$.

(*ii*) **One mole** of $CH_4(g)$ reacts with **2 moles** of $O_2(g)$ to give **one mole** of $CO_2(g)$ and **2 moles** of $H_2O(g)$.

(*iii*) **16 *g*** of $CH_4(g)$ reacts with **2 × 32 *g*** of $O_2(g)$ to give **44 *g*** of CO_2 and **2 × 18 *g*** of $H_2O(g)$.

(*iv*) **22.4 L** of $CH_4(g)$ reacts with **44.8 L** of $O_2(g)$ to give **22.4 L** of CO_2 and **44.8 L** of $H_2O(g)$ at N.T.P.

Limitations of Chemical Equations and Methods to make them more Informative

The chemical equations, as such do not give us the following informations :

(*i*) The physical states of the reactants and products.
(*ii*) The concentration of the reactants and products.
(*iii*) The conditions such as temperature, pressure or catalyst, etc. which affect the reaction.
(*iv*) The heat changes accompanying the reaction *i.e.,* whether heat is evolved or absorbed during the reaction.
(*v*) The formation of a precipitate or the evolution of gas in the reaction.
(*vi*) The speed of the reaction.
(*vii*) The nature of the reaction *i.e.,* whether the chemical reaction is reversible or irreversible.

Removal of Drawbacks of Chemical Equations

The chemical equation can be made more informative by the following modifications;

(*a*) *The physical states of the reactants and products can be specified.* For this, we use *s* (for solids), *l* (for liquids) and *g* (for gases). Sometimes, the letter *aq* is given to represent that the given substance has been dissolved in excess of water. For example,

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

(*b*) In order to express the strength of acid or base used in the reaction, the words *conc.* for concentrated and *dil.* for dilute are used before the formula of the acid or base. For example,

$Zn\ (s) + \text{dil. } 2HCl\ (aq) \longrightarrow ZnCl_2(aq) + H_2(g)$

(*c*) *The conditions of the reaction such as temperature, pressure, catalyst, etc.* may be written on the arrow between the reactants and products. For example,

$N_2(g) + 3H_2(g) \xrightarrow[600\text{ atm.}]{Fe/Mo,\ 723\text{ K}} 2NH_3\ (g)$

The above equation indicates that the reaction between hydrogen and nitrogen has been carried out in the presence of Fe/Mo (catalyst) at 723 K and 600 atm. pressure.

(*d*) *Heat changes taking place during the reaction may be expressed in the equation.* For example,

$C(s) + O_2\ (g) \longrightarrow CO_2(g) + 394\text{ kJ}$ (Heat is evoled)
$N_2(g) + O_2(g) \longrightarrow 2NO\ (g) - 180.5\text{ kJ}$ (Heat is a absorbed)

(*e*) *The formation of a precipitate, if any, in a reaction can be expressed by writing the word ppt. or by an arrow pointing downward.* For example,

$AgNO_3(aq) + NaCl(aq) \longrightarrow \underset{\text{(ppt)}}{AgCl(s)\downarrow} + NaNO_3(aq)$

Similarly, the evolution of a gas in a chemical reaction can be indicated by an arrow pointing upward. For example,

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

(*f*) The distinction between slow and fast reactions can be made by writing word *slow* and *fast* on the arrow head.

(*g*) The reversible nature of the reaction maybe indicated by double headed arrow which indicates that the reaction is occurring in the forward as well as in the backward direction :

$N_2(g) + 3H_2\ (g) \underset{\text{Backward}}{\overset{\text{Forward}}{\rightleftharpoons}} 2NH_3\ (g)$

Balancing of Chemical Equation

A correct chemical equation must be in accordance with the law of conservation of mass according to which the total mass of the reactants must be equal to the total mass of the products. In other words, the number of atoms of each kind in the reactants must be equal to the number of atoms of same kind in the products. *Such a chemical equation which has an equal number of atoms of each element in the reactants and the products is called* **balanced chemical equation.** For example, the reaction between hydrogen and oxygen to form water may be written as

$$H_2 + O_2 \longrightarrow H_2O$$

But this equation has two atoms of O on the reactants side but only one atom of O on the products side. Therefore, it does not obey law of conservation of mass. By choosing suitable coefficients, the equation may be balanced as :

$$2H_2 + O_2 \longrightarrow 2H_2O$$

The balancing of a chemical equation means to equalise the number of atoms of each element on both sides of the equation. There are many methods to balance a chemical equation. Some of these methods are :

(*i*) Hit and Trial method or Trial and Error method.

(*ii*) Partial equation method

(*iii*) Oxidation number method

(*iv*) Ion-electron method.

The first two methods are discussed below while the other two methods will be taken up in Unit 8 (Redox reactions).

Hit and Trial Method or Trial and Error method

This method involves the following steps :

(*i*) *Write the symbols and formulae of the reactants and the products in the form of a skeleton equation.*

(*ii*) *If any elementary gas appears, change it to its atomic state.*

(*iii*) *Select the formula containing maximum number of atoms and start the process of balancing.*

(*iv*) *In case, the above method fails, then start balancing the atoms which appear minimum number of times.*

(*v*) *Atoms of elementary gases are balanced last of all.*

(*vi*) *Once all the atoms are balanced, change the equation into the molecular form.*

(*vii*) *Verify that the number of atoms of each element is balanced in the final equation.*

For example, let us consider the equation :

$$Fe + H_2O \longrightarrow Fe_3O_4 + H_2$$

(*i*) Changing the elementary substance hydrogen to atomic form,

$$Fe + H_2O \longrightarrow Fe_3O_4 + 2H$$

(*ii*) Fe_3O_4 has largest number of atoms. To balance this, multiply H_2O by 4 to balance oxygen atoms. In 4 molecules of H_2O, there are 8 atoms of H which are balanced by multiplying H on the R.H.S by 8.

$$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 8\,H$$

(*iii*) Converting to molecular form,

$$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$$

is a balanced equation.

Similarly, some other balanced equations are :

$$4Fe\,(s) + 3O_2\,(g) \longrightarrow 2Fe_2\,O_3\,(s)$$
$$P_4\,(s) + 5O_2\,(g) \longrightarrow P_4O_{10}\,(s)$$
$$2Mg\,(s) + O_2\,(g) \longrightarrow 2MgO\,(s)$$

SOLVED EXAMPLES

❑ **Example 68.**

Balance the equation

$$C_3H_8 + O_2 \longrightarrow CO_2 + H_2O$$

Solution: Converting the oxygen to elementary state

$$C_3H_8 + O \longrightarrow CO_2 + H_2O$$
unbalanced equation

Balancing carbon atoms

$$C_3H_8 + O \longrightarrow 3CO_2 + H_2O$$

Balancing hydrogen atoms

$$C_3H_8 + O \longrightarrow 3CO_2 + 4H_2O$$

Balancing oxygen atoms

$$C_3H_8 + 10\,O \longrightarrow 3CO_2 + 4H_2O$$

Changing to molecular form

$$C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$$

❑ **Example 69.**

Balance the following skeleton equation :

$$As_2O_3 + SnCl_2 + HCl \longrightarrow SnCl_4 + As + H_2O$$

Solution: The skeleton equation is :

$$As_2O_3 + SnCl_2 + HCl \longrightarrow SnCl_4 + As + H_2O$$

Balancing arsenic and oxygen atoms by multiplying As by 2 and H_2O by 3

$$As_2O_3 + SnCl_2 + HCl \longrightarrow SnCl_4 + 2As + 3H_2O$$

Multiply HCl by 6 to balance H atoms

$$As_2O_3 + SnCl_2 + 6HCl \longrightarrow SnCl_4 + 2As + 3H_2O$$

To balance chlorine atoms, multiply $SnCl_2$ and $SnCl_4$ both by 3.

$$As_2O_3 + 3SnCl_2 + 6HCl \longrightarrow 3SnCl_4 + 2As + 3H_2O$$

Partial Equation Method

This is a better method than the hit and trial method. It involves the following steps :

(*i*) *The given chemical reaction is supposed to occur* in steps and chemical equation called **partial equation,** is written for each step.

(*ii*) *Each partial equation is balanced separately by hit and trial method as discussed earlier.*

(*iii*) *The partial equations are multiplied by suitable numbers if necessary, so as to cancel out the species which are not involved in the original reactants and the final products.*

(*iv*) *Finally, the partial equations are added to get the final equation.*

Let us illustrate this method with an example of reaction between copper and nitric acid. The reaction is written as :

$$Zn + HNO_3 \longrightarrow Zn\,(NO_3)_2 + N_2O + H_2O$$

(*i*) The probable partial equations for the above reaction are :

$$Zn + HNO_3 \longrightarrow Zn\,(NO_3)_2 + 2H$$
$$HNO_3 + H \longrightarrow N_2O + H_2O$$

(*ii*) Balance the partial chemical equations separately by hit and trial method as

$$Zn + 2HNO_3 \longrightarrow Zn\,(NO_3)_2 + 2H$$
$$2HNO_3 + 8H \longrightarrow N_2O + 5H_2O$$

(*iii*) Multiply the first partial equation by 4 in order to cancel out H atoms in the two equations since they do not appear in the final equation. Finally add the two equations

$$[Zn + 2HNO_3 \longrightarrow Zn(NO_3)_2 + 2H] \times 4$$
$$2HNO_3 + 8H \longrightarrow N_2O + 5H_2O$$

$$4Zn + 10HNO_3 \longrightarrow 4Zn(NO_3)_2 + N_2O + 5H_2O$$

Practice Problems

58. Balance the following equations by hit and trial method :

(*a*) $KMnO_4 + HCl \longrightarrow KCl + MnCl_2 + H_2O + Cl_2$

(*b*) $H_2\,S + SO_2 \longrightarrow S + H_2O$

(*c*) $K_2Cr_2O_7 + H_2SO_4 \longrightarrow K_2SO_4 + Cr_2\,(SO_4)_3 + H_2O + O_2$

(*d*) $KMnO_4 + KOH \longrightarrow K_2MnO_4 + O_2 + H_2O$

(*e*) $Mg_3N_2 + H_2O \longrightarrow Mg(OH)_2 + NH_3$

(*f*) $Al_4C_3 + H_2O \longrightarrow Al(OH)_3 + CH_4$

(*g*) $FeS_2 + O_2 \longrightarrow Fe_2O_3 + SO_2$

(*h*) $KMnO_4 + H_2S + H_2SO_4 \longrightarrow KHSO_4 + MnSO_4 + S + H_2O$

(*i*) $C_3H_8\,(g) + O_2\,(g) \longrightarrow CO_2(g) + H_2O\,(l)$

59. Balance the following equations by partial equation method :

(*i*) $NaOH + Cl_2 \longrightarrow NaCl + NaClO_3 + H_2O$

(*ii*) $H_2S + HNO_3 \longrightarrow NO + H_2O + S$

(*iii*) $C + H_2SO_4 \longrightarrow CO_2 + SO_2 + H_2O$

(*iv*) $I_2 + HNO_3 \longrightarrow NO_2 + HIO_3 + H_2O$

(*v*) $P_4 + HNO_3 \longrightarrow H_3PO_4 + NO_2 + H_2O$

Answers to Practice Problems

58. (*a*) $2KMnO_4 + 16\,HCl \longrightarrow 2KCl + 2MnCl_2 + 8H_2O + 5Cl_2$

(*b*) $2H_2S + SO_2 \longrightarrow 2H_2O + 3S$

(*c*) $2K_2Cr_2O_7 + 8H_2SO_4 \longrightarrow 2K_2SO_4 + 2Cr_2(SO_4)_3 + 8H_2O + 3O_2$

(*d*) $4KMnO_4 + 4KOH \longrightarrow 4K_2MnO_4 + O_2 + 2H_2O$

(*e*) $Mg_3N_2 + 6H_2O \longrightarrow 3Mg(OH)_2 + 2NH_3$

(*f*) $Al_4C_3 + 12H_2O \longrightarrow 4Al(OH)_3 + 3CH_4$

(*g*) $4FeS_2 + 11O_2 \longrightarrow 2Fe_2O_3 + 8SO_2$

(*h*) $2KMnO_4 + 5H_2S + 4H_2SO_4 \longrightarrow 2KHSO_4 + 2MnSO_4 + 5S + 8H_2O$

(*i*) $C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O\,(l)$

59. (*i*) $6NaOH + 3Cl_2 \longrightarrow 5NaCl + NaClO_3 + 3H_2O$

(*ii*) $3H_2S + 2HNO_3 \longrightarrow 2NO + 4H_2O + 3S$

(*iii*) $C + 2H_2SO_4 \longrightarrow CO_2 + 2SO_2 + 2H_2O$

(*iv*) $I_2 + 10HNO_3 \longrightarrow 10NO_2 + 2HIO_3 + 4H_2O$

(*v*) $P_4 + 20HNO_3 \longrightarrow 4H_3PO_4 + 4H_2O + 20NO_2$

STOICHIOMETRIC CALCULATIONS

Problems Based on Chemical Equations (Stoichiometry)

Solving of stoichiometric problems is very important. It requires grasp and application of mole concept, balancing of chemical equations and care in the conversion of units.

The problems based upon chemical equations may be classified as :

(*i*) **Mole to mole relationships.** In these problems, the moles of one of the reactants/products is to be calculated if that of other reactants/products are given.

(*ii*) **Mass-mass relationships.** In these problems, the mass of one of the reactants/products is to be calculated if that of the other reactants/products are given.

(*iii*) **Mass-volume relationship.** In these problems, mass or volume of one of the reactants or products is calculated from the mass or volume of other substances.

(*iv*) **Volume-volume relationship.** In these problems, the volume of one of the reactants/products is given and that of the other is to be calculated.

The **main steps** for solving such problems are :

(i) Write down the balanced chemical equation.

(ii) Write down the moles or gram atomic or gram molecular masses of the substances whose quantitites are given or have to be calculated. In case, there are two or more atoms or molecules of a substance, multiply the mole or gram atomic mass or molecular mass by the number of atoms or molecules.

(iii) Write down the actual quantities of the substances given. For the substances whose weights/volumes have to be calculated, write the sign of interrogation (?)

(iv) Calculate the result by a unitary method.

This method can be illustrated by the following examples :

SOLVED EXAMPLES

Type I. Mole to mole relationship

Example 70.

How many moles of nitrogen are needed to produce 8.2 moles of ammonia by reaction with hydrogen ?

Solution: The balanced chemical equation is :

$$\underset{?}{N_2 + 3H_2} \longrightarrow \underset{8.2\ \text{mol}}{2NH_3}$$

2 mol of NH_3 are produced from = 1 mol of N_2

8.2 mol of NH_3 are produced from = $\frac{1}{2} \times 8.2$

= 4.1 mol of N_2

Example 71.

How many moles of iron can be made from Fe_2O_3 by the use of 16 mol of carbon monoxide in the following reaction :

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$$

Solution: The balanced equation is :

$$Fe_2O_3 + \underset{16\ \text{mol}}{3CO} \longrightarrow \underset{?}{2Fe} + 3CO_2$$

3 mol of CO are used to make = 2 mol Fe

16 mol of CO are used to make = $\frac{2}{3} \times 16$

= 10.67 mol

Type II. Mass to mass or Mole to mass relationship

Example 72.

Calculate the amount of water in grams produced by combustion of 16 g methane (CH_4). **N.C.E.R.T.**

Solution: The balanced chemical equation is

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

1 mol	2 mol
16 g	36 g
16 g	?

From the equation, it is clear that 16 g of methane produces 36 g of water.

Example 73.

How many moles of methane are required to produce 22g of $CO_2(g)$ after combustions ? **N.C.E.R.T.**

Solution: The balanced equation is

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

1 mol	1 mol
	44 g
?	22 g

44 g of CO_2 (g) are produced from = 1 mol of CH_4

22 g of CO_2 g will be produced by burning $CH_4 = \frac{1 \times 22}{44}$

= 0.5 mol.

Example 74.

Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid (HCl) according to the reaction :

$$MnO_2(s) + 4HCl(aq) \longrightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$$

How many grams of HCl react with 5.0 g of manganese dioxide ? **N.C.E.R.T.**

Solution: The balanced chemical equation is :

$$MnO_2(s) + 4HCl(aq) \longrightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$$

1 mol 4 mol

$55 + 2 \times 16$ $4 \times (1 + 35.5)$

$= 87$ g $= 146.0$ g

5 g ?

87 g of MnO_2 react with 146.0 g of HCl

5 g of MnO_2 will react with $= \frac{146.0}{86} \times 5$

$=$ **8.39 g**

❑ Example 75.

What mass of calcium oxide will be obtained by heating 3 mol of $CaCO_3$?

Solution: The balanced chemical equation is :

$$CaCO_3 \longrightarrow CaO + CO_2$$

1 mol 1 mol

$40 + 16$

$= 56$

3 mol ?

1 mol of $CaCO_3$ on heating gives CaO = 56 g

3 mol of $CaCO_3$ on heating gives CaO = 56×3

= **168 g**

❑ Example 76.

Oxygen is prepared by the catalytic decomposition of potassium chlorate ($KClO_3$). Decomposition of potassium chlorate gives potassium chloride (KCl) and oxygen (O_2). If 2.4 mol of oxygen is needed for an experiment, how many grams of potassium chlorate must be decomposed ?

Solution: The balanced equation is

$$2\,KClO_3\,(s) \longrightarrow 2KCl(s) + 3O_2(g)$$

2 mol 3 mol

$2 \times (30 + 35.5 + 3 \times 16)$

$= 245$ g

? 2.4 mol

3 mol of O_2 is produced by decomposition of $KClO_3$ = 245 g

2.4 mol of O_2 will be produced by the decomposition of

$KClO_3 = \frac{245 \times 2.4}{3}$ = **196.0 g**

❑ Example 77.

Calculate the weight of iron which will be converted into its oxide (Fe_3O_4) by the action of 14.4 g of steam on it.

Solution: The balanced chemical equation is:

$$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$$

3 mol 4 mol

3×56 4×18

$= 168$ g $= 72$ g

Now 72 g of steam react with = 168 g of Fe

14.4 g of steam will react with $= \frac{168}{72} \times 14.4$

= **33.6 g**

❑ Example 78.

How many grams of chlorine are required to completely react with 0.40 g of hydrogen (H_2) to yield hydrochloric acid (HCl) ? Also calculate the amount of HCl formed.

Solution: The balanced chemical equation is :

$$H_2 + Cl_2 \longrightarrow 2HCl$$

1 mol 1 mol

2 g $2 \times 35.5 = 71$ g

2 g of H_2 react with = 71 g Cl_2

0.40 g of H_2 would react with $= \frac{71}{2} \times 0.4 =$ **14.2 g Cl_2**

Now,

$$H_2 + Cl_2 \longrightarrow 2HCl$$

1 mol 2 mol

2 g 73 g

2 g of hydrogen give HCl = 73 g

0.40 g of hydrogen will give HCl $= \frac{73}{2} \times 0.40 =$ **14.6 g**

❑ Example 79.

What weight of zinc would be required to produce enough hydrogen to reduce completely 8.5 g of copper oxide to copper ?

Solution: (*i*) **Calculation of H_2 required to reduce CuO to Cu.**

$$CuO + H_2 \longrightarrow Cu + H_2O$$

1 mol 1 mol

$63.5 + 16 = 79.5$ g 2 g

One mole of CuO require 1 mole of H_2 for reduction to copper

or 79.5 g of CuO require H_2 for reduction to copper = 2 g

8.5 g of CuO require H_2 for reduction to copper $= \frac{2}{79.5} \times 8.5$

$= 0.214$ g

(*ii*) **Calculation of weight of zinc required to produce H_2**

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$$

1 mol 1 mol

65 g 2 g

? 0.214 g

2 g of H_2 is produced from zinc = 65 g

0.214 g of H_2 would be produced from zinc $= \frac{65}{2} \times 0.214$

= **6.955 g**

❑ Example 80.

Calculate the amount of lime $Ca(OH)_2$ required to remove the hardness of 60,000 litres of well water containing 16.2 g of calcium bicarbonate per hundred litre. (Atomic masses Ca = 40, C = 12, O = 16, H = 1)

Solution: (*i*) **Calculation of the weight of calcium bicarbonate present.**

Wt. of calcium bicarbonate present in 100 litres of well water = 16.2 g

Wt. of calcium bicarbonate present in 60,000 litres of well water $= \frac{16.2}{100} \times 60000$ g = 9720 g

(*ii*) *Calculation of the quantity of lime required.*

The equation involved is :

$$\underset{\substack{40+2+24+96 \\ =162 \\ 162}}{Ca(HCO_3)_2} + \underset{\substack{40+32+2 \\ =74 \\ ?}}{Ca(OH)_2} \longrightarrow 2CaCO_3 + 2H_2O$$

∴ One mole of calcium bicarbonate requires one mole of calcium hydroxide

or 162 g of $Ca(HCO_3)_2$ require = 74 g of $Ca(OH)_2$

∴ 9720 g of $Ca(HCO_3)_2$ will require $= \frac{74}{162} \times 9720$ g of $Ca(OH)_2$

= 4440 g of $Ca(OH)_2$

Hence the quantity of lime required = **4440 g = 4.44 kg**

❑ Example 81.

A 1.84 g mixture of calcium carbonate and magnesium carbonate upon heating gave 0.96 g residue. Calculate the percentage composition of the mixture (At. mass of Ca = 40, Mg = 24, C = 12 and O = 16).

Solution: Calcium carbonate and magnesium carbonate on heating give CaO and MgO as residues.

$$\underset{100\text{ g}}{CaCO_3} \xrightarrow{\text{heat}} \underset{56\text{ g}}{CaO} + CO_2$$

$$\underset{84\text{ g}}{MgCO_3} \xrightarrow{\text{heat}} \underset{40\text{ g}}{MgO} + CO_2$$

Let $CaCO_3$ present in the mixture = x g

Wt. of $MgCO_3$ in the mixture = $(1.84 - x)$ g

Now, 100 g of $CaCO_3$ give CaO = 56 g

x g of $CaCO_3$ give CaO $= \frac{56}{100} \times x$ g

Similarly, 84 g of $MgCO_3$ give MgO = 40 g

$(1.84 - x)$ g of $MgCO_3$ give MgO $= \frac{40}{84} \times (1.84 - x)$

Total residue obtained,

$$\frac{56x}{100} + \frac{40}{84}(1.84 - x) = 0.96$$

$$4704\,x + 7360 - 4000\,x = 8064$$

$$704\,x = 704$$

$$x = \frac{704}{704} = 1\text{ g}$$

Weight of $CaCO_3$ in the mixture = 1 g

Weight of $MgCO_3$ in the mixture = 1.84 − 1 = 0.84 g

% of $CaCO_3 = \frac{1}{1.84} \times 100$ = **54.35%**

% of $MgCO_3 = \frac{0.84}{1.84} \times 100$ = **45.65%**

Type III. Mass – Volume relationship

❑ Example 82.

An impure sample of sodium chloride which weighed 1.20 g gave on treatment with excess of $AgNO_3$ solution 2.40 g of silver chloride as a precipitate. Calculate the percentage purity of the sample.

Solution:

$$\underset{\substack{23+35.5 \\ =58.5 \\ ?}}{NaCl} + AgNO_3 \longrightarrow \underset{\substack{108+35.5 \\ =143.5 \\ 2.40\text{ g}}}{AgCl} + NaNO_3$$

143.5 g of AgCl is obtained from 58.5 g of pure NaCl

2.40 g of AgCl is obtained from pure NaCl

$$= \frac{58.5}{143.5} \times 2.40 = 0.978\text{ g}$$

Now, 1.20 g of NaCl contain pure NaCl = 0.978 g

100 g of NaCl contain pure NaCl $= \frac{0.978}{1.20} \times 100 = 81.5$ g

Percentage purity of sample = **81.5%**

❑ Example 83.

A 2.0 g of sample containing Na_2CO and $NaHCO_3$ loses 0.248 g when heated to 300°C, the temperature at which $NaHCO_3$ decomposes to Na_2CO_3, CO_2 and water. What is the percentage of Na_2CO_3 in the mixture ?

Solution: The balanced chemical equation is :

$$\underset{\substack{2(23+1+12+48) \\ =168}}{2NaHCO_3} \longrightarrow Na_2CO_3 + \underset{18\text{ g}}{H_2O} + \underset{44\text{ g}}{CO_2}$$

CO_2 and H_2O will escape as gases at 300°C. Therefore, loss in weight would correspond to the weight of H_2O and CO_2.

Total mass of CO_2 + water lost by heating 168 g of $NaHCO_3$ = (18 + 44) = 62 g

62 g of weight is lost by heating $NaHCO_3$ = 168 g

0.248 g of weight is lost by heating $NaHCO_3$

$$= \frac{168}{62} \times 0.248 = 0.672$$

2 g sample contain $NaHCO_3$ = 0.672 g

Weight of Na_2CO_3 in sample = 2 × 0.672 = 1.328 g

∴ Percentage of $Na_2CO_3 = \frac{1.328}{2} \times 100$ = **66.4%**

The equations involving volume are based upon the fact that 1 mol of all gases at N.T.P. occupy 22.4 L. The problems involving different conditions of temperature and pressure are described in next unit.

Example 84.

Calculate the amount of $KClO_3$ needed to supply sufficient oxygen for burning 112 L of CO gas at N.T.P.

Solution: Calclation of O_2 gas required to burn 112 L of CO.

$$\underset{\substack{1\text{ mol}\\22.4\text{ L at N.T.P.}}}{CO} + \underset{0.5\text{ mol}}{\frac{1}{2}O_2} \longrightarrow CO_2$$

22.4 L of CO at N.T.P. require O_2 = 0.5 mol

112 L of CO at N.T.P. require $O_2 = \frac{0.5}{22.4} \times 112$

$= 2.5$ mol

This O_2 is to be obtained by heating $KClO_3$.

$$\underset{2\text{ mol}}{2KClO_3} \longrightarrow 2\,KCl + \underset{3\text{ mol}}{3O_2}$$

2 (39 + 35.5 + 3 × 16) = 245 g

3 mol of O_2 is produced from $KClO_3$ = 245 g

2.5 mol of O_2 is produced from $KClO_3 = \frac{245 \times 25}{3}$

= **204.167 g**

Example 85.

What volume of air at N.T.P. containing 21% of oxygen by volume is required to completely burn 1000 g of sulphur containing 4% incombustible matter ?

Solution: The balanced chemical equation is :

$$\underset{\substack{1\text{ mol}\\32\text{ g}}}{S} + O_2 \longrightarrow \underset{\substack{1\text{ mol}\\22.4\text{ L}}}{SO_2}$$

Amount of sulphur in the sample

$= \frac{1000 \times 96}{100} = 960$ g

32 g of sulphur require oxygen at N.T.P. = 22.4 L

960 g of sulphur require oxygen at N.T.P.

$= \frac{22.4}{32} \times 960$

= 672 L

Volume of air required $= \frac{672 \times 100}{21}$

= **3200 L**

Type IV. Volume - Volume relationship

Example 86.

Calculate the volume of oxygen at N.T.P. that would be required to convert 5.2 L of carbon monoxide to carbon dioxide.

Solution: The balanced chemical equation is :

$$\underset{\substack{2\text{ mol}\\2\times 22.4\text{ L}}}{2CO} + \underset{\substack{1\text{ mol}\\22.4\text{ L}}}{O_2} \longrightarrow 2CO_2$$

Volume of oxygen required to convert 2 × 22.4 L of CO at N.T.P. = 22.4 L

Volume of oxygen required to convert 5.2 L of CO at N.T.P. $= \frac{22.4}{2 \times 22.4} \times 5.2$

= **2.6 L**

Example 87.

What volume of oxygen at S.T.P. is required to effect complete combustion of 200 cm³ of acetylene and what would be the volume of carbon dioxide formed ?

Solution: The balanced chemical equation is :

$$\underset{\substack{2\text{ mol}\\2\times 22400\text{ cm}^3}}{2C_2H_2} + \underset{\substack{5\text{ mol}\\5\times 22400\text{ cm}^3}}{5O_2} \longrightarrow \underset{\substack{4\text{ mol}\\4\times 22400\text{ cm}^3}}{4CO_2} + 2H_2O$$

(*i*) **Calculation of volume of O_2 at S.T.P. required to effect complete combustion of 200 cm³ of acetylene.**

2 × 22400 cm³ of acetylene require O_2 for complete combustion = 5 × 22400 cm³

200 cm³ of acetylene require O_2 for complete combustion $= \frac{5 \times 22400}{2 \times 22400} \times 200$

= **500 cm³ at S.T.P**

(*ii*) **Calculation of volume of carbon dioxide formed**

2 × 22400 cm³ of acetylene produce CO_2 = 4 × 22400

∴ 200 cm³ of acetylene produce CO_2

$= \frac{4 \times 22400}{2 \times 22400} \times 200$

= **400 cm³ at S.T.P.**

Practice Problems

60. How much iron can be obtained by the reduction of 1 kg of Fe_2O_3 ?

61. An hourly energy requirement of an astronaut can be satisfied by the energy released when 34 grams of sucrose ($C_{12}H_{22}O_{11}$) are burnt in his body. How many grams of oxygen would he need to be carried in space capsule to meet his requirement for one day ?

62. How much marble of 96.5% purity would be required to prepare 100 litres of carbon dioxide at S.T.P. when marble is acted upon by dil HCl ?

63. 5.6 litres of methane gas (CH_4) is ignited in oxygen gas. Calculate the number of moles of CO_2 formed.

64. Calculate the percentage yield of the reaction if 64 g $NaBH_4$ with iodine produced 15.0 g of BI_3.

$NaBH_4 + 4I_2 \longrightarrow BI_3 + NaI + 4HI$

(At. mass, Na = 23, B = 10.8, I = 127)

65. NO_2^- ion in KNO_2 is oxidised to NO_3^- ion by the action of $KMnO_4$ in H_2SO_4 solution according to the reaction :

$5KNO_2 + 2KMnO_4 + 3H_2SO_4 \longrightarrow 5KNO_3 + 2MnSO_4 + 5I_2 + 8H_2O$

How much $KMnO_4$ are needed to oxidise 11.4 g of KNO_2 ?

66. How many ml of $KMnO_4$ solution containing 158 g/L must be used to complete the conversion of 75 g of KI to I_2 in the acidic solution ?

$$2KMnO_4 + 10KI + 8H_2SO_4 \longrightarrow 6K_2SO_4 + 2MnSO_4 + 5I_2 + 8H_2O$$

Answers to Practice Problems

60. 0.7 kg
61. 916.2 g
62. 462.6 g
63. 0.25 mole
64. 2.27%
65. 8.47 g
66. 90.5 ml

Hints & Solutions on page 79

LIMITING REACTANT

In many cases, the substances in a mixture are not present in exactly the same amount as required by the balanced chemical equation. In such situations, one reactant is in excess over the other. The reactant which is present in lesser amount gets consumed after sometime and after that no further reaction takes place even though the other reactant is present. Thus,

the reactant which gets completely consumed in a reaction is called **limiting reactant.**

The concentration of the limiting reactant limits the amount of products formed. The other reactants present in quantities greater than those needed to react with the quantity of the limiting reagent present would be left unreacted. It is also called **excess reagent.** For example, in the reaction

$$2H_2 + O_2 \longrightarrow 2H_2O$$

if the reaction mixture contains 2 mol of H_2 and 2 mol of O_2, then only 1 mol of O_2 will be used up and 1 mol of O_2 will be left over. Therefore, in this case, *hydrogen is the limiting reactant. Oxygen is excess reactant.*

SOLVED EXAMPLES

Example 88.

In a reaction

$$A + B_2 \longrightarrow AB_2$$

identify the limiting reagent if any in the following reaction mixtures :

(i) 300 atoms of A + 200 molecules of B

(ii) 2 mol of A + 3 mol of B

(iii) 100 atoms of A + 100 molecules of B

(iv) 5 mol of A + 2.5 mol of B

(v) 2.5 mol of A + 5 mol of B. **N.C.E.R.T.**

Solution: According to the equation,
1 mole of A reacts with 1 mole of B_2 and 1 atom of A reacts with 1 molecule of B_2

(*i*) B is limiting reagent because 200 molecules of B_2 will react with 200 atoms of A and 100 atoms of A will be left in excess.

(*ii*) A is limiting reagent because 2 mole of A will react with 2 mol of B and 1 mol of B will be left in excess.

(*iii*) Both will react completely because it is stoichiometric mixture. No limiting reagent.

(*iv*) 2.5 mol of B will react with 2.5 mol of A and hence B is limiting reagent.

(*v*) 2.5 mol of A will react with 2.5 mol of B. Hence A is limiting reagent.

Example 89.

50.0 kg of N_2 (g) and 10.0 kg of H_2(g) are mixed to produce NH_3(g)

(a) Calculate the NH_3 (g) formed.

(b) Identify the limiting reagent in this reaction if any. **N.C.E.R.T.**

Solution:

$$\text{Moles of } N_2 = \frac{\text{Mass in } g}{\text{Molar mass}}$$

$$= \frac{50.0 \times 10^3}{28} = 1.786 \times 10^3 \text{ mol}$$

$$\text{Mole of } H_2 = \frac{\text{Mass in } g}{\text{Molar mass}}$$

$$= \frac{10 \times 10^3}{2} = 5.0 \times 10^3 \text{ mol}$$

$$\underset{1 \text{ mol}}{N_2 (g)} + \underset{3 \text{ mol}}{3H_2(g)} \rightleftharpoons 2NH_3 (g)$$

According to the equation 1 mol of N_2 (*g*) requires 3 mol of H_2(*g*) for the reaction.

1.786×10^3 mol of N_2(*g*) will require H_2(*g*)

$= 3 \times 1.786 \times 10^3$

$= 5.36 \times 10^3$ mol

But moles of H_2 present $= 5.0 \times 10^3$ mol

Therefore, N_2 is in excess and **dihydrogen is limiting reagent**

Now, 3 mol of H_2 (*g*) gives NH_3 (*g*) = 2 mol

5.0×10^3 mol of H_2(*g*) will give NH_3 (*g*)

$= \frac{2}{3} \times 5.0 \times 10^3$

$= 3.3 \times 10^3$ mol

Mass of NH_3 (*g*) formed = Moles × Molar mass

$= 3.3 \times 10^3 \times 17$

= 56.1 kg

Example 90.

3.0 g of H_2 react with 29.0 g of O_2 to yield H_2O.

(i) Which is the limiting reactant ?

(ii) Calculate the maximum amount of H_2O that can be formed.

(iii) Calculate the amount of one of the reactants which remains unreacted.

Solution: The balanced chemical equation is

$$\underset{2 \text{ mol}}{2H_2} + \underset{1 \text{ mol}}{O_2} \longrightarrow \underset{2 \text{ mol}}{2H_2O}$$

$$\text{Moles of } H_2 = \frac{\text{Wt. of } H_2}{\text{Molecular mass of } H_2} = \frac{3.0}{2} = 1.50 \text{ mol}$$

$$\text{Moles of } O_2 = \frac{\text{Wt. of } O_2}{\text{Molecular mass of } O_2} = \frac{29.0}{32.0} = 0.906 \text{ mol}$$

(*i*) **Calculation of limiting reactant**

According to the above equation, 2 mol of H_2 require O_2 = 1 mol

$\therefore$ 1.50 mol of H_2 require $O_2 = \frac{1.50}{2} = 0.75$ mol

But moles of O_2 actually present = 0.906 mol

Therefore, O_2 is in excess and H_2 is the **limiting reactant.**

(*ii*) **Calculation of maximum amount of H_2O formed.**

Now, 2 mol of H_2 form H_2O = 2 mol

$\therefore$ 1.50 mol of H_2 form H_2O = 1.50 mol

Thus, the maximum amount of H_2O formed = **1.50 mol**

(*iii*) **Calculation of the amount of one of the reactants which remains unreacted i.e., O_2**

Number of moles of O_2 added = 0.906 mol

Number of moles of O_2 used up = 0.75 mol

$\therefore$ Number of moles of O_2 unreacted = 0.906 − 0.75 = **0.156 mol.**

❑ Example 91.

If 20.0 g of $CaCO_3$ is treated with 20.0 g of HCl, how many grams of CO_2 can be produced according to the reaction :

$$CaCO_3\,(s) + 2HCl\,(aq) \longrightarrow CaCl_2\,(aq) + H_2O(l) + CO_2(g)$$

N.C.E.R.T.

Solution: In this problem, CO_2 can be calculated by using either of the reactants $CaCO_3$ or HCl. So, first of all we are to calculate the limiting reactant.

$$CaCO_3\,(s) \quad + \quad 2HCl\,(aq) \longrightarrow CaCl_2\,(aq) + H_2O(l) + CO_2(g)$$

1 mol	2 mol
40 + 12 + 3 × 16	2 × (1 + 35.5)
= 100 g	= 73 g

According to above equation,

100 g of $CaCO_3$ require 73 g of HCl

20 g of $CaCO_3$ require $= \frac{73}{100} \times 20 = 14.6$ g

But amount of HCl actually present = 20.0 g

Therefore, $CaCO_3$ is limiting reactant and HCl is excess reactant.

Now let us calculate the amount of CO_2 produced when entire quantity of limiting reactant reacts.

$$CaCO_3(s) + 2HCl\,(g) \longrightarrow CaCl_2\,(aq) + H_2O\,(l) + CO_2\,(g)$$

1 mol	1 mol
100 g	44 g
20 g	?

100 g of $CaCO_3$ produces CO_2 = 44 g

20 g of $CaCO_3$ will produce $CO_2 = \frac{44}{100} \times 20$ = **8.80 g**

❑ Example 92.

Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation :

$$N_2\,(g) \;+\; 3\,H_2(g) \longrightarrow 2NH_3\,(g)$$

(*i*) *Calculate the mass of ammonia produced if 2.00×10^3 g of dinitrogen reacts with 1.00×10^3 g of dihydrogen.*

(*ii*) *Will any of the two reactants remain unreacted ?*

(*iii*) *If yes, which one and what would be its mass ?*

Solution: The balanced chemical equation is

$$N_2\,(g) \;+ 3\,H_2\,(g) \rightleftharpoons 2NH_3\,(g)$$

$$\text{Moles of } N_2 = \frac{2.00 \times 10^3}{28} = 71.43 \text{ mol}$$

$$\text{Moles of } H_2 = \frac{1.00 \times 10^3}{2} = 500 \text{ mol}$$

According to above equation, 1 mol of N_2 require 3 mol of H_2.

$\therefore$ 71.43 mol of N_2 will require H_2 = 3 × 71.43 = 214.29 mol

But moles of H_2 actually present = 500 mol

$\therefore$ H_2 is in excess and will remain unreacted and N_2 is limiting reagent.

(*i*) 1 mol of N_2 react with H_2 to form NH_3 = 2 mol

71.43 mol of N_2 will react with H_2 to form NH_3

$= \frac{2}{1} \times 71.43$

= 142.86 mol

Mass of NH_3 produced = 142.86 × 17 = 2428.6 g

(*ii*) Hydrogen will remain unreacted.

(*iii*) Moles of H_2 initially taken = 500 mol

Moles of H_2 reacting = 214.29 mol

Moles of H_2 remaining unreacted = 500 – 214.29 = 285.71 mol

Mass of H_2 left unreacted = 285.71 × 2 = **571.42 g**

REACTIONS IN SOLUTIONS

There are many reactions which take place in solutions. In solution generally one component is present in lesser amount and is called *solute* while the other present in excess is called the *solvent. The amount of solute present in a given quantity of solvent or solution is expressed in terms of* **concentration**.

The concentration of the solution is usually expressed in the following ways :

1. Mass percentage or volume percentage

The **mass percentage** *of a component in a given solution* is the **mass of the component per 100 g of the solution**. For example, if W_A is the mass of component A and W_B is the mass of component B in a solution, then

$$\text{Mass percentage of A} = \frac{W_A}{W_A + W_B} \times 100$$

This can be expressed as w/w. For example, a 10% (*w/w*) solution of sodium chloride means that 10 g of sodium chloride is present in 90 g of water so that the total mass of the solution is 100 g or simply 10 g of sodium choride is present in 100 g of solution.

Volume percentage. In case of a liquid dissolved in another liquid, it is convenient to express the concentrations in volume percentage. The **volume percentage** is defined *as* **the volume of the component per 100 parts by volume of the solution**. For example, if V_A and V_B are the volume of two components A and B respectively in a solution, then

Volume percentage of A

$$= \frac{\text{Volume of A}}{\text{Volume of A + Volume of B}} \times 100$$

This may be expressed as *v/v*. Sometimes, we express the concentrations as weight/volume. For example, a 10% solution of sodium chloride (*w/v*) means that 10 g of sodium chloride are dissolved in 100 mL of solution.

Parts per million. When a solute is present in very minute amounts (trace quantities), the concentration is expressed in *parts per million* abbreviated as **ppm. It is the parts of a component per million parts of the solution**. It is expressed as:

$$ppm\ \text{A} = \frac{\text{Mass of component A}}{\text{Total mass of solution}} \times 10^6$$

For example, suppose a litre of public supply water contains about 3×10^{-3} g of chlorine. The mass percentage of chlorine is :

$$\text{Mass percentage of chlorine } \frac{3.0 \times 10^{-3}}{1000} \times 100$$

$$= 3 \times 10^{-4} \quad \text{(Total volume is 100 mL)}$$

The parts per million parts of chlorine is :

$$\text{or} \quad ppm \text{ of chlorine} = \frac{3 \times 10^{-3} \times 10^6}{1000} = 3$$

Thus, instead of expressing concentration of chlorine as 3×10^{-4} % it is better to express as 3 *ppm.*

Atmospheric pollution in cities due to harmful gases is generally expressed in ppm though in this case the values refer to volumes rather than masses. For example, the concentration of SO_2 in Delhi has been found to be as high as 10 ppm. This means that 10 cm^3 of SO_2 are present in 10^6 cm^3 (or 10^3 L) of air.

The concentration of atmospheric pollutants in cities is generally expressed in terms of mg/mL.

❑ Example 93.

If 11 g of oxalic acid are dissolved in 500 mL of solution (density = 1.1 g mL^{-1}), what is the mass % of oxalic acid in solution ?

Solution : 11 g of oxalic acid are present in 500 mL of solution.

Density of solution = 1.1 g mL^{-1}

Mass of solution = (500 mL) × (1.1. g mL^{-1})

= 550 g

Mass of oxalic acid = 11 g

$$\text{Mass \% of oxalic acid} = \frac{11}{550} \times 100 = \mathbf{2\%}.$$

2. Molarity of a solution

It is **the number of moles of the solute dissolved per litre of the solution**. It is represented as 'M'. Thus, a solution which contains one gram mole of the solute dissolved per litre* of the solution, is regarded as **one molar** solution. For example, 1M Na_2CO_3 (molar mass = 106) solution has 106 g of the solute present per litre of the solution.

$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Volume of solution in litres}}$$

It is convenient to express volume in cm^3 or mL so that

$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Volume of solution (in mL or cm}^3)} \times 1000$$

$$(\because 1 \text{ litre} = 1000 \text{ mL})$$

Thus, **the units of molarity are moles per litre (mol L^{-1}) or moles per cubic decimetre (mol dm^{-3})** . The symbol M is used for mol L^{-1} or mol dm^{-3} and it represents molarity.

If n_B moles of solute are present in V mL of solution, then

$$\text{Molarity} = \frac{n_B}{V} \times 1000$$

Moles of solute can be calculated as :

$$\text{Moles of solute} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}}$$

Molarity is one of the common measures of expressing concentration which is frequently used in the laboratory. However, it has one **disadvantage.** *It changes with temperature because of expansion or contraction of the liquid with temperature.*

❑ Example 94.

2.46 g of sodium hydroxide (molar mass = 40) are dissolved in water and the solution is made to 100 cm^3 in a volumetric flask. Calculate the molarity of the solution.

Solution : Amount of NaOH = 2.46 g

Volume of solution = 100 cm^3

$$\text{Moles of NaOH} = \frac{\text{Mass of NaOH}}{\text{Molar mass}}$$

$$= \frac{2.46}{40} = 0.0615 = 0.0615$$

$$\text{Molarity} = \frac{\text{Moles of NaOH}}{\text{Volume of solution}} \times 1000$$

$$= \frac{0.0615}{100} \times 1000$$

= **0.615 M.**

* In SI units, volume is expressed as dm^3 and 1 litre = 1 dm^3.

3. Molality of a solution

It is **the number of moles of the solute dissolved per 1000 g (or 1 kg) of the solvent**. It is denoted by *m*. Mathematically,

$$\text{Molality } (m) = \frac{\text{Moles of solute}}{\text{Weight of solvent in kg}}$$

$$\text{or} \quad = \frac{\text{Moles of solute}}{\text{Mass of solvent in gram}} \times 1000$$

Thus, **the units of molality are moles per kilogram *i.e.*, mol kg^{-1}.** It is represented by the symbol **m**.

If n_B moles of solute are dissolved in W grams of solvent, then

$$\text{Molality} = \frac{n_B}{W} \times 1000$$

KEY POINT

From the discussion of molarity and molality, it is evident that in molarity we consider the *volume of the solution* while in molality, we take the *mass of the solvent*. Therefore, the two are never equal. **Molality is considered better for expressing the concentration as compared to molarity because the molarity changes with temperature because of expansion or contraction of the liquid with temperature. However, molality does not change with temperature because mass of the solvent does not change with change in temperature.**

❑ **Example 95.**

Calculate the molality of a solution containing 20.7 g of potassium carbonate dissolved in 500 mL of solution (assume density of solution = 1 g mL^{-1}).

Solution : Mass of K_2CO_3 = 20.7 g
Molar mass of K_2CO_3 = 138

$$\text{Moles of } K_2CO_3 = \frac{20.7}{138} = 0.15$$

Mass of solution = (500 mL) × (1 g mL^{-1}) = 500 g
Amount of water = 500 – 20.7 = 479.3 g

$$\text{Molality} = \frac{\text{Moles of solute}}{\text{Mass of solvent in gram}} \times 1000$$

$$= \frac{0.15}{479.3} \times 1000 = \mathbf{0.313\ m.}$$

4. Mole fraction

It is the ratio of number of moles of one component to the total number of moles (solute and solvent) present in the solution. It is denoted by *x*. Let us suppose that a solution contains n_A moles of solute and n_B moles of the solvent. Then,

$$\text{Mole fraction of solute } (x_A) = \frac{n_A}{n_A + n_B}$$

$$\text{Mole fraction of solvent } (x_B) = \frac{n_B}{n_A + n_B}$$

The sum of mole fractions of all the components in solution is always equal to one as shown below :

$$x_A + x_B = \frac{n_A}{n_A + n_B} + \frac{n_B}{n_A + n_B} = 1$$

Thus, if the mole fraction of one component of a binary solution is known, that of the other can be calculated. For example, the mole fraction x_A is related to x_B as :

$x_A = 1 - x_B$ or $x_B = 1 - x_A$

It may be noted that the **mole fraction is independent of temperature.**

❑ **Example 96.**

A solution is prepared by adding 60 g of methyl alcohol to 120 g of water. Calculate the mole fraction of methanol and water.

Solution : Mass of methanol = 60 g

$$\text{Moles of methanol} = \frac{60}{32} = 1.875$$

(*Molar mass = 32*)

Mass of water = 120 g

$$\text{Moles of water} = \frac{120}{18} = 6.667$$

(*Molar mass = 18*)

Total number of moles = 1.875 + 6.667
= 8.542

$$\text{Mole fraction of methanol} = \frac{1.875}{8.542} = \mathbf{0.220}$$

$$\text{Mole fraction of water} = \frac{6.667}{8.542} = \mathbf{0.780.}$$

5. Normality

It is the number of gram equivalents of the solute dissolved per litre of the solution. It is denoted by N.

$$\text{Normality (N)} = \frac{\text{Number of gram equivalents of solute}}{\text{Volume of solution in litres}}$$

or Normality

$$= \frac{\text{Number of gram equivalents of solute}}{\text{Volume of solution in mL}} \times 1000$$

Thus, **the units of normality are gm equivalent per litre *i.e.* g equiv L^{-1}.**

Gram equivalents of solute can be calculated as :

$$\text{Gram equivalents of solute} = \frac{\text{Mass of solute}}{\text{Equivalent mass}}$$

Like molarity, **normality of a solution also changes with temperature.**

❑ **Example 97.**

Calculate the normality of solution containing 31.5 g of hydrated oxalic acid ($H_2C_2O_4 . 2H_2O$) in 1250 mL of solution.

Solution : Mass of oxalic acid = 31.5 g

$$\text{Equivalents of oxalic acid} = \frac{31.5}{63} = 0.5$$

(Eq. wt. 126/2 = 63)

Volume of solution = 1250 mL

$$\text{Normality} = \frac{0.5}{1250} \times 1000 = \mathbf{0.4\ N.}$$

▶ Relationship between Normality and Molarity of Solutions

The normality and molarity of a solution are related as :

$$\text{Normality} = \text{Molarity} \times \frac{\text{Molar mass}}{\text{Equivalent mass}}$$

For acids,

Normality = Molarity × Basicity

where basicity is the number of H^+ ions that a molecule of an acid can give in solution.

For bases,

Normality = Molarity × Acidity

where acidity is the number of OH^- ions that a molecule of a base can give in solution.

For example,

1 M H_2SO_4 solution = 2N H_2SO_4 solution (basicity = 2)

1 M H_3PO_4 solution = 3N H_3PO_4 solution (basicity = 3)

and 1 M $Ca(OH)_2$ solution = 2N $Ca(OH)_2$ solution (acidity = 2)

Alternatively, 1 N H_2SO_4 solution = 0.5 M H_2SO_4 solution (basicity = 2) and so on.

▶ It may be noted that these days the terms normality or equivalent weight are not commonly used.

Sometimes, the term **formality** is also used. It gives the number of formula masses of the solute dissolved per litre of the solution. It is represented by F.

Formality

$$= \frac{\text{Number of formula mass of the solute}}{\text{Volume of the solution in mL}} \times 1000$$

This term is used to express the concentration of ionic substances (*e.g.* NaCl, KNO_3, $CuSO_4$) which do not exist as discrete molecules. In such cases, we do not use the term mole for expressing the concentration. The sum of the atomic masses of various atoms constituting the formula of the ionic compound is called **gram formula mass** instead of molar mass.

WATCH OUT !

- Molarity of a solution changes with temperature due to accompanied changes in volume of the solution.
- Molality and mole fraction do not change with temperature.

REMEMBER

FORMULAE TO REMEMBER

$$\textbf{Molarity (M)} = \frac{\text{Moles of solute}}{\text{Vol. of solution (in mL)}} \times 1000$$

$$\textbf{Molality }(m) = \frac{\text{Moles of solute}}{\text{Mass of solvent (in g)}} \times 1000$$

$$\textbf{Normality (N)} = \frac{\text{Gram equiv. of solute}}{\text{Vol. of solution (in mL)}} \times 1000$$

$$\textbf{Normality} = \text{Molarity} \times \frac{\text{Mol. mass}}{\text{Eq. mass}}$$

$$\textbf{Mole fraction of solute } (x_B) = \frac{\text{Moles of solute}}{\text{Moles of solute} + \text{Moles of solvent}}$$

Mole fraction of solute + Mole fraction of solvent = 1

or Mole fraction of solvent = 1 – Mole fraction of solute

SOLVED EXAMPLES

Example 98.

A solution is prepared by dissolving 18.25 g of NaOH in distilled water to give 200 ml of solution. Calculate the molarity of the solution.

Solution: Moles of NaOH $= \frac{18.25}{40}$

(Molecular mass of NaOH = 40)

Vol. of solution = 200 mL

Molarity $= \frac{18.25 \times 1000}{40 \times 200}$ **= 2.28 M**

Example 99.

How many grams of Na_2CO_3 should be dissolved to make 100 cm^3 of 0.15 M Na_2CO_3 solution ?

Solution: 1000 cm^3 of 0.15 M Na_2CO_3 contain Na_2CO_3 = 0.15 mole

100 cm^3 of 0.15 M Na_2CO_3 contain Na_2CO_3

$$= \frac{0.15}{1000} \times 100 = 0.015 \text{ mole}$$

(Molar mass of Na_2CO_3 = 106)

Mass of Na_2CO_3 = 0.015 × 106

= **1.59 g**

Example 100.

A solution is prepared by dissolving 2 g of substance A in 18 g of water. Calculate the mass percentage of solute. **N.C.E.R.T.**

Solution:

Mass of solute, A = 2 g

Mass of water = 18 g

Mass of solution = 2 + 18 = 20 g

$$\text{Mass percent of A} = \frac{\text{Mass of A}}{\text{Mass of solution}} \times 100 \text{ g}$$

$$= \frac{2}{20} \times 100 = \mathbf{10\%}$$

Example 101.

Calculate the concentration of nitric acid in moles per litre in a sample which has a density, 1.41 g mL^{-1} and the mass percent of nitric acid in it being 69%. **N.C.E.R.T.**

Solution:

69 mass percent of nitric acid means that 69 g of HNO_3 are present in 100 g of solution.

$$\text{Volume of solution} = \frac{\text{Mass}}{\text{density}} = \frac{100 \text{ g}}{1.41 \text{ g mL}^{-1}}$$
$$= 70.92 \text{ mL}$$

$$\text{Moles of } HNO_3 = \frac{69}{63}$$

$$\therefore \quad \text{Molarity} = \frac{\text{Moles of } HNO_3}{\text{Volume of solution (in mL)}} \times 1000$$
$$= \frac{69/63}{70.92} \times 1000 = \mathbf{15.44\ M}$$

Example 102.

A sample of $NaNO_3$ weighing 0.38 g is placed in a 250 ml volumetric flask. The flask is then filled with water to the mark on the neck. What is the molarity of the solution ?

Solution: Mass of $NaNO_3$ dissolved = 0.38 g

$$\text{Molecular mass of } NaNO_3 = 23 + 14 + 48 = 85$$
$$\text{Moles of } NaNO_3 \text{ dissolved} = \frac{0.38}{85}$$
$$\text{Volume of solution} = 250 \text{ mL}$$
$$\text{Molarity} = \frac{0.38/85}{250} \times 1000$$
$$= \mathbf{0.018\ M}$$

Example 103.

What is the concentration of sugar ($C_{12}H_{22}O_{11}$) in mol L^{-1} if its 20 g are dissolved in enough water to make a volume upto 2 L ? **N.C.E.R.T.**

Solution: Mass of sugar = 20 g

$$\text{Molecular mass of sugar} = 12 \times 12 + 1 \times 22 + 11 \times 16$$
$$= 342$$
$$\text{Moles of sugar} = \frac{20}{342}$$
$$\text{Volume of solution} = 2 \text{ L}$$
$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Vol. of solution in L}}$$
$$= \frac{20/342}{2} = \mathbf{0.029\ mol\ L^{-1}}$$

Example 104.

If the density of methanol is 0.793 kgL^{-1}, what is the volume needed for making 2.5 L of its 0.25 M solution ? **N.C.E.R.T.**

Solution: Let us calculate moles of methanol present in 2.5 L of 0.25 M solution.

$$\text{Molarity} = \frac{\text{Moles of } CH_3OH}{\text{Volume in L}}$$
$$0.25 = \frac{\text{Moles of } CH_3OH}{2.5}$$
$$\therefore \quad \text{Moles of } CH_3OH = 0.25 \times 2.5 = 0.625 \text{ moles}$$
$$\text{Mass of } CH_3OH = 0.625 \times 32 = 20 \text{ g}$$

(Molecular mass of CH_3OH = 32)

Now 0.793×10^3 g of CH_3OH is present in 1000 mL

$$20 \text{ g of } CH_3OH \text{ will be present in} = \frac{1000}{0.793 \times 10^3} \times 20$$
$$= \mathbf{25.2\ mL}$$

Example 105.

How many moles and how many grams of sodium chloride are present in 250 mL of a 0.50 M NaCl solution ?

Solution: This can be calculated by using the formula :

$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Vol. of solution (in mL)}} \times 1000$$
$$0.50 = \frac{\text{Moles of NaCl}}{250} \times 1000$$
$$\therefore \text{Moles of NaCl} = \frac{0.50 \times 250}{1000} = \mathbf{0.125\ mol}$$

Gram molecular mass of NaCl = 23 + 35.5 = 58.5 g

∴ Mass of NaCl solution in grams

= Moles of NaCl × Molecular mass

= 0.125 × 58.5 = **7.3125 g**

Example 106.

Calculate the number of Cl^- ions in 100 ml of 0.001 M HCl solution.

Solution : Since HCl is a strong acid, it ionises completely so that the concentration of HCl is equal to that of Cl^- ions.

1000 ml of 0.001 M HCl solution contains Cl^- = 0.001 mole

100 ml of 0.001 M HCl solution contains Cl^-

$$= \frac{0.001 \times 100}{1000} = 1 \times 10^{-4} \text{ mole}$$

$$\text{No. of } Cl^- \text{ ions} = 6.022 \times 10^{23} \times 1.0 \times 10^{-4}$$
$$= \mathbf{6.022 \times 10^{19}}.$$

Example 107.

A solution of oxalic acid, $(COOH)_2.2H_2O$ is prepared by dissolving 0.63 g of the acid in 250 mL of the solution. Calculate (i) molarity and (ii) normality of the solution.

Solution : (*i*) Calculation of molarity

Molar mass of oxalic acid, $(COOH)_2 . 2H_2O$

$$= 2\,(12 + 32 + 1) + 2 \times 18$$
$$= 126 \text{ g mol}^{-1}$$

$$\text{Moles of oxalic acid} = \frac{0.63}{126} = 0.005 \text{ mol}$$
$$\text{Volume of solution} = 250 \text{ mL}$$
$$\text{Molarity} = \frac{0.005 \times 1000}{250} = \mathbf{0.02\ M}$$

(*ii*) Calculation of normality

Equivalent mass of oxalic acid

$$= \frac{\text{Mol. mass of oxalic acid}}{\text{Basicity}}$$
$$= \frac{126}{2} = 63$$

$$\text{Gram equivalents of oxalic acid} = \frac{0.63}{63} = 0.01$$

$$\text{Normality} = \frac{\text{Gram equivalents of solute}}{\text{Volume of solution (in mL)}} \times 1000$$
$$= \frac{0.01}{250} \times 1000 = \mathbf{0.04\ N}$$

Example 108.

2.82 g of glucose (molar mass = 180) are dissolved in 30 g of water. Calculate (a) the molality (b) mole fraction of glucose and water.

Solution : (*a*) **Calculation of molality of solution.**

Mass of glucose = 2.82 g

Moles of glucose = $\frac{2.82}{180}$ (Molar mass = 180)

Mass of water = 30 g

$$\text{Molality} = \frac{\text{Moles of glucose}}{\text{Mass of water}} \times 1000 = \frac{2.82 \times 1000}{180 \times 30} = \mathbf{0.522\ m.}$$

(*b*) **Calculation of mole fraction**

Moles of glucose = $\frac{2.82}{180} = 0.0157$

Moles of water = $\frac{30}{18} = 1.67$

Mole fraction of glucose = $\frac{0.0157}{0.0157 + 1.67} = \mathbf{0.009.}$

Mole fraction of water = $\frac{1.67}{0.0157 + 1.67} = \mathbf{0.991.}$

Example 109.

Calculate the molarity of pure water (density of water = 1 g mL^{-1})

Solution : Density of water = 1 g mL^{-1}

Mass of 1000 mL of water = Volume × Density = 1000 × 1 = 1000 g

Moles of water = $\frac{1000}{18} = 55.55$

Now, 55.55 moles of H_2O are present in 1000 ml or 1 L of water

∴ Molarity = **55.55 M.**

Example 110.

A solution is 25% water, 25% ethanol and 50% acetic acid by mass. Calculate the mole fraction of each component.

Solution : Let the total mass of solution = 100 g

Mass of water = 25 g

Mass of ethanol = 25 g

Mass of acetic acid = 50 g

Moles of water = $\frac{25}{18} = 1.388$ (∵ Molar mass of H_2O = 18)

Moles of ethanol = $\frac{25}{46} = 0.543$ (∵ Molar mass of C_2H_5OH = 46)

Moles of acetic acid = $\frac{50}{60} = 0.833$ (∵ Molar mass of CH_3COOH = 60)

Total number of moles = 1.388 + 0.543 + 0.833 = 2.764

Mole fraction of water = $\frac{1.388}{2.764} = \mathbf{0.502.}$

Mole fraction of ethanol = $\frac{0.543}{2.764} = \mathbf{0.196.}$

Mole fraction of acetic acid = $\frac{0.833}{2.764} = \mathbf{0.302.}$

Example 111.

A solution of glucose in water is labelled as 10% (w/w). The density of the solution is 1.20 g mL^{-1}. Calculate

(i) molality

(ii) molarity, and

(iii) mole fraction of each component in solution.

Solution : 10% (w/w) solution of glucose means that 10 g of glucose is present in 100 g of solution or in 90 g of water.

(*i*) **Calculation of molality**

Mass of glucose = 10 g

Moles of glucose = $\frac{10}{180} = 0.0556$

(Molar mass of glucose = 180)

Mass of water = 90 g

$$\therefore \quad \text{Molality} = \frac{\text{Moles of glucose}}{\text{Mass of water}} \times 1000 = \frac{0.0556}{90} \times 1000 = \mathbf{0.618\ m.}$$

(*ii*) **Calculation of molarity**

Moles of glucose = 0.0556

$$\text{Volume of solution} = \frac{\text{Mass}}{\text{Density}} = \frac{100}{1.20} = 83.3 \text{ mL}$$

$$\text{Molarity} = \frac{\text{Moles of glucose}}{\text{Vol. of solution}} \times 1000 = \frac{0.0556}{83.3} \times 1000 = \mathbf{0.667\ M.}$$

(*iii*) **Calculation of mole fraction of components**

Moles of glucose = 0.0556

Moles of water = $\frac{90}{18} = 5.0$

Total moles = 5.0 + 0.0556 = 5.0556

Mole fraction of glucose = $\frac{0.0556}{5.0556} = \mathbf{0.011.}$

Mole fraction of water = $\frac{5.0}{5.0556} = \mathbf{0.989.}$

Example 112.

A sugar syrup of weight 214.2 g contains 34.2 g of sugar ($C_{12}H_{22}O_{11}$). Calculate :

(i) molal concentration, and

(ii) mole fraction of sugar in the syrup.

Solution : (*i*) Weight of sugar syrup = 214.2 g

Weight of sugar in syrup = 34.2 g

Weight of water in syrup = 214.2 – 34.2 = 180.0 g

Moles of sugar = $\frac{34.2}{342} = 0.1$

(Molar mass = 342)

$$\text{Molality} = \frac{0.1}{180} \times 1000 = \mathbf{0.56\ m.}$$

(ii) $$\text{Moles of sugar} = \frac{34.2}{342} = 0.1$$

$$\text{Moles of water} = \frac{180}{18} = 10$$

$$\text{Mole fraction of sugar} = \frac{0.1}{10 + 0.1} = \mathbf{0.0099.}$$

Molarity Equation

To calculate the volume of a definite molarity of a solution required to prepare solution of other molarity, we can use the relation

$$M_1V_1 = M_2V_2$$

where M_1 = Initial molarity,
M_2 = Molarity of new solution
V_1 = Initial volume,
V_2 = Volume of new solution

This is known as ***molarity equation*** and is commonly used to calculate the molarity of solution after dilution.

Normality Equation

Like molarity equation, to calculate the volume of a definite normality of a solution required to prepare solution of other normality, we can use the relation :

$$N_1V_1 = N_2V_2$$

where N_1 is the initial normality of a solution having volume V_1 and N_2 is the normality of new solution having volume V_2. This is known as ***normality equation.***

Molarity equation and normality equation are **commonly used** to

(*i*) calculate the molarity or normality of a solution after mixing two or more solutions.

(*ii*) calculate the volume of the solution of given molarity or normality required to dilute to get solution of known molarity or normality. For example:

(*a*) If V_1 mL of a solution of molarity M_1 are mixed with V_2 mL of solution of molarity M_2, then molarity of the solution after mixing M_3 can be calculated as :

$$M_1V_1 + M_2V_2 = M_3(V_1 + V_2)$$

(*b*) If V_2 mL of a solution of molarity M_2 is to be prepared from a solution of molarity M_1 (concentrated) then volume of solution (say V_1) of molarity M_1 required can be calculated from the molarity equation:

$$M_1 \times V_1 = M_2 \times V_2$$

or $$V_1 = \frac{M_2 \times V_2}{M_1}$$

For example, if we want to prepare 500 mL of 0.5 M HCl solution from 10 M HCl solution, then volume of 10 M solution required can be calculated as:

$$10\ M \times V_1 = 0.5\ M \times 500\ mL$$

$$V_1 = \frac{0.5 \times 500}{10} = 25\ mL$$

$\therefore$ Volume of 10 M HCl required = 25 mL
Water to be added = 500 – 25 = 475 mL

❑ Example 113.

Calculate the volume of 0.015 M HCl solution required to prepare 250 mL of a 5.25×10^{-3} M HCl solution.

Solution: Applying molarity equation.

$$\underbrace{M_1V_1}_{\text{Initial}} = \underbrace{M_2V_2}_{\text{Final}}$$

$$0.015\ M \times V_1 = 5.25 \times 10^{-3}\ M \times 250\ mL$$

$$\therefore \quad V_1 = \frac{5.25 \times 10^{-3} \times 250}{0.015}$$

$$= \mathbf{87.5\ mL}$$

❑ Example 114.

250 mL of 1.5 M solution of sulphuric acid is diluted by adding 5L of water. What is the molarity of the diluted solution ?

Solution: Volume of diluted solution = 5000 + 250
= 5250 mL

Applying molarity equation

$$\underbrace{M_1V_1}_{\text{Initial}} = \underbrace{M_2V_2}_{\text{Final}}$$

$$1.5\ M \times 250\ mL = M_2 \times 5250\ mL$$

$$\therefore \quad M_2 = \frac{1.5 \times 250\ mL}{5250\ mL}$$

$$= \mathbf{0.0714\ M}$$

❑ Example 115.

What volume of 10 M HCl and 3 M HCl should be mixed to get 1L of 6 M HCl solution?

Solution : Suppose volume of 10 M HCl required to prepare 1 L of 6 M HCl = x litre

Volume of 3 M HCl required = $(1 - x)$ litre

Applying molarity equation :

$$\underbrace{M_1V_1}_{10\ M\ HCl} + \underbrace{M_2V_2}_{3\ M\ HCl} = \underbrace{M_3V_3}_{6\ M\ HCl}$$

$$10 \times x + 3(1 - x) = 6 \times 1$$

$$10x + 3 - 3x = 6$$

$$7x = 3 \text{ or } x = \frac{3}{7} = 0.428\ L$$

$\therefore$ Vol. of 10 M HCl required = 0.428 L = **428 mL**

Vol. of 3 M HCl required = 1 – 0.428 = 0.572 L
= **572 mL**

❑ Example 116.

Commercially available concentrated hydrochloric acid contains 38% HCl by mass.

(i) What is the molarity of the solution (density of solution = 1.19 g mL^{-1}) ?

(ii) What volume of concentrated HCl is required to make 1.0 L of an 0.10 M HCl ?

Solution: (*i*) 38% HCl by mass means that 38 g of HCl is present in 100 g of solution.

$$\text{Volume of solution} = \frac{\text{Mass}}{\text{Density}} = \frac{100}{1.19} = 84.03 \text{ mL}$$

$$\text{Moles of HCl} = \frac{38}{36.5} = 1.04$$

$$\text{Molarity} = \frac{1.04 \times 1000}{84.03} = \mathbf{12.38\ M}$$

(*ii*) The volume of this solution required to make 1.0 L of 0.10 M HCl can be calculated by applying molarity equation as

$$\underbrace{M_1V_1}_{acid_1} = \underbrace{M_2V_2}_{acid_2}$$

$$12.38 \text{ M} \times V_1 = 0.10 \text{ M} \times 1.0 \text{ L}$$

$$\therefore \quad V_1 = \frac{0.10 \times 1.0}{12.38} = 0.00808 \text{ L or} = \mathbf{8.08\ cm^3}$$

❑ **Example 117.**

Commercially available sulphuric acid contains 93% acid by mass and has a density of 1.84 g mL^{-1}. Calculate (i) the molarity of the solution (ii) volume of concentrated acid required to prepare 2.5 L of 0.50 M H_2SO_4

Solution: (*i*) 93% H_2SO_4 by mass means that 93 g of H_2SO_4 is present in 100 g of solution.

$$\text{Vol. of 100 g of solution} = \frac{\text{Mass}}{\text{Density}} = \frac{100}{1.84} = 54.3 \text{ mL}$$

$$\text{Moles of } H_2SO_4 = \frac{93}{98} = 0.95 \text{ mol}$$

$$\text{Molarity} = \frac{0.95 \times 1000}{54.3} = 17.5 \text{ M}$$

(*ii*) Applying molarity equation,

$$\underbrace{M_1V_1}_{acid_1} = \underbrace{M_2V_2}_{acid_2}$$

$$17.5 \text{ M} \times V_1 = 0.5 \text{ M} \times 2.5 \text{ L}$$

$$\therefore \quad V_1 = \frac{0.5 \times 2.5}{17.5} = 0.071 \text{ L} = \mathbf{71\ mL}$$

Converting one concentration unit into another concentration unit.

● Calculating molality from molarity

Molality is the moles of solute dissolved per 1000g of the solvent whereas molarity is the moles of solute per 1000 mL of the solution.

Molality can be calculated from molarity of the solution as :

❒ Calculate the total mass of 1L (1000 mL) of solution by using density of the solution.

❒ Calculate mass of solute from the molarity (no. of moles) in 1L of the solution.

❒ Calculate the mass of the solvent by subtracting mass of solute from the total mass of solution.

❒ From the values of moles of solute and mass of solvent calculate molality.

❑ **Example 118.**

The density of 3 M solution of NaCl is 1.25 g mL^{-1}. Calculate the molality of the solution. **N.C.E.R.T.**

Solution : 3 M solution means that 3 moles of NaCl are present in 1 L solution.

Mass of NaCl in 1 L solution = 3 × 58.5
= 175.5 g

(∵ Molecular mass of NaCl = 58.5)

Density of solution = 1.25 g mL^{-1}

Mass of 1 L solution = 1000 × 1.25 = 1250 g

Mass of water in solution = 1250 – 175.5
= 1074.5 g

$$\text{Molality} = \frac{\text{Moles of solute}}{\text{Mass of solvent (in g)}} \times 1000$$

$$= \frac{3}{1074.5} \times 1000$$

$$= \mathbf{2.79\ m}$$

● Calculating molality from mole fraction

Molality is the moles of solute dissolved per 1000g of solvent whereas mole fraction is the ratio of moles of solute present to the total moles of solute and solvent present in the solution. Mole fraction also tells us that out of total of 1 mol of solution, x_{solute} mol of solute and $(1 - x)$ mol of solvent are present.

Molality can be calculated from the mole fraction as

❒ Calculate the number of moles of solvent in 1 mol of solution.

❒ Convert moles of solvent into mass of solvent.

❒ From the values of moles of solvent and mass of solvent calculate molality.

❑ **Example 119.**

What is the molality of a solution of methanol in water in which the mole fraction of methanol is 0.25?

Solution: 1 mole of solution contains 0.25 mol of methanol and 1 – 0.25 = 0.75 mol of water.

Mass of water = 0.75 × 18 = 13.5 g

Molality of methanol solution

$$= \frac{\text{Moles of methanol}}{\text{Mass of solvent}} \times 1000$$

$$= \frac{0.25}{13.5} \times 1000$$

$$= \mathbf{18.5\ m}$$

● Calculating mole fraction from molality

Mole fraction is the ratio of moles of solute to the total number of moles of solute and solvent whereas molality is the moles of solute dissolved in 1000 g of solvent.

Mole fraction can be calculated from the molality as

❒ Moles of solute is equal to molality of solution.

❒ If we consider 1000 g of solvent, then calculate moles of solvent from 1000 g of solvent.

❒ From the values of moles of solute and moles of solvent, calculate mole fraction.

Example 120.

What is the mole fraction of the solute in 2.5 m aqueous solution?

Solution: 2.5 m aqueous solution means that 2.5 moles of solute are present in 1000 g of water. Thus,

$$\text{Moles of solute} = 2.5 \text{ mol}$$

$$\text{Moles of water} = \frac{1000}{18} = 55.6$$

$$\text{Mole fraction of solute} = \frac{2.5}{2.5 + 55.6}$$

$$= \mathbf{0.043}$$

Practice Problems

67. A sample of NaOH weighing 0.38 g is dissolved in water and the solution is made to 50.0 mL in a volumetric flask. What is the molarity of the resulting solution ? **N.C.E.R.T.**

68. The density of 3 molal solution of NaOH is 1.110 g mL^{-1}. Calculate molarity of the solution. *(NCERT Exemplar Problem)*

69. A bottle contains 500 ml of 2.4 M HCl solution. How much water should be added to dilute it to 1.6 M HCl solution ?

70. A bottle of concentrated sulphuric acid (density 1.80 g cm^{-3}) is labelled as 86% by weight. What is the molarity of the solution ?

71. 0.63 g of oxalic acid, $(COOH)_2 .2H_2O$ are dissolved in 500 ml of solution. Calculate the molarity of the solution.

72. How many moles of NaOH are contained in 27 mL of 0.15 M NaOH ?

73. Calculate the number of oxalic acid molecules in 100 ml of 0.01 M oxalic acid solution.

Answers to Practice Problems

67. 0.19 M
68. 2.97 M
69. 250 ml
70. 15.8 M
71. 0.01 M
72. 0.00405 mol
73. 6.023×10^{20}

Hints & Solutions on page 79

STOICHIOMETRY OF REACTIONS IN SOLUTIONS

Many reactions are carried out in aqueous solutions. The amounts of the product of a reaction can be calculated from the volumes of the solutions of the reactants and their concentrations. These are illustrated below:

SOLVED EXAMPLES

Example 121.

250 ml of 0.5 M sodium sulphate (Na_2SO_4) solution are added to an aqueous solution containing 10.0 g of $BaCl_2$ resulting in the formation of white precipitate of $BaSO_4$. How many moles and how many grams of barium sulphate will be obtained ?

Solution: The balanced chemical equation is :

$$BaCl_2\,(aq) + Na_2SO_4\,(aq) \longrightarrow BaSO_4\,(s) + 2NaCl\,(aq)$$

Let us first calculate moles of Na_2SO_4 and $BaCl_2$

0.5 M solution of Na_2SO_4 means that 0.5 mol of Na_2SO_4 are present in 1000 mL of solution.

$$1000 \text{ mL of solution contain } Na_2SO_4 = 0.5 \text{ mol}$$

$$250 \text{ mL of solution contain } Na_2SO_4 = \frac{0.5}{1000} \times 250$$

$$= 0.125 \text{ mol}$$

$$\text{Moles of } BaCl_2 \text{ in solution} = \frac{10}{208}$$

(Mol. mass of $BaCl_2$ = 137 + 2 × 35.5 = 208)

$$= 0.048$$

According to the balanced equation, 1 mol of $BaCl_2$ reacts with 1 mol of Na_2SO_4. Therefore, $BaCl_2$ is the limiting reactant, so only 0.048 mol of Na_2SO_4 reacts with 0.048 mol of Na_2SO_4.

Now, according to the equation,

$$1 \text{ mol of } BaCl_2 \text{ produces } BaSO_4 = 1 \text{ mol}$$

$$0.048 \text{ mol of } BaCl_2 \text{ produces } BaSO_4 = 1 \times 0.048$$

$$= \mathbf{0.048\ mol}$$

$$\text{Amount of } BaSO_4 \text{ obtained} = 0.048 \times 233$$

(Mol. mass of $BaSO_4$ = 233)

$$= \mathbf{11.18\ g}$$

Example 122.

What volume of 0.6 M HCl has enough hydrochloric acid to react exactly with 25 mL of aqueous NaOH having concentration of 0.5 M ?

Solution: The balanced chemical equation is :

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

Let us first calculate the moles of HCl and NaOH present in their solutions.

$$1000 \text{ ml of 0.5 M NaOH contains} = 0.5 \text{ mol}$$

$$25 \text{ ml of 0.5 M NaOH contain} = \frac{0.5 \times 25}{1000}$$

$$= 0.0125 \text{ mol}$$

From the equation, it is clear that 1 mol of NaOH reacts with 1 mol of HCl. Therefore, moles of HCl required for the reaction = 0.0125 mol.

Now, we are to calculate the volume of 0.6 M HCl which contains 0.0125 mol.

0.6 mol of HCl of 0.6 M concentration are present in = 1000 mL

0.125 mol of HCl of 0.6 M concentration are present in

$$= \frac{1000 \times 0.0125}{0.6}$$

$$= \mathbf{20.83\ mL}$$

Example 123.

What volume of 0.250 M HCl (aq) is required to react completely with 22.6 g of sodium carbonate according to the reaction :

$$Na_2CO_3\,(s) + 2HCl\,(aq) \longrightarrow 2NaCl\,(aq) + H_2O + CO_2$$

Solution: The balanced chemical equation is :

$$\underset{1\text{ mol}}{Na_2CO_3\,(s)} + \underset{2\text{ mol}}{2HCl\,(aq)} \longrightarrow 2NaCl\,(aq) + H_2O + CO_2$$

Moles of Na_2CO_3 present $= \frac{22.6}{106} = 0.213$

(Molar mass = 2 × 23 + 12 + 3 × 16 = 106)

Now, according to the equation,

1 mol of Na_2CO_3 requires moles of HCl = 2

0.213 mol of Na_2CO_3 requires moles of HCl = 2 × 0.213 = 0.426 mol

Now, we are to calculate the volume of 0.250 M HCl containing 0.426 moles,

0.250 mole of 0.250 M HCl is present in = 1000 mL

0.426 mol of 0.250 M HCl would be present in

$$= \frac{1000}{0.250} \times 0.426 = \mathbf{1704\ mL.}$$

Example 124.

Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction given below :

$$CaCO_3\,(s) + 2HCl\,(aq) \longrightarrow CaCl_2\,(aq) + CO_2\,(g) + H_2O\,(l)$$

What mass of $CaCl_2$ will be formed when 250 mL of 0.76 M HCl react with 1000 g of $CaCO_3$? Name the limiting reagent. Calculate the number of moles of $CaCl_2$ formed in the reaction. ***(NCERT Exemplar Problem)***

Solution:

$$CaCO_3(s) + 2HCl\,(aq) \longrightarrow CaCl_2\,(aq) + CO_2\,(g) + H_2O\,(l)$$

Mass of $CaCO_3$ = 1000 g

Moles of $CaCO_3 = \frac{1000}{100} = 10$ mol

1000 mL of 0.76 M HCl contain HCl = 0.76 mol

250 mL of 0.76 M HCl contain HCl $= \frac{0.76}{1000} \times 250$

= 0.19 mol

Now, according to the equation, 1 mol of $CaCO_3\,(s)$ requires 2 mol of HCl (*aq*). Hence, for 10 mol of $CaCO_3(s)$, moles of HCl(*aq*) required would be

$$= \frac{2}{1} \times 10 = 20 \text{ mol}$$

But we have only 0.19 mole of HCl (*aq*)

∴ HCl (*aq*) is limiting reagent.

2 mol of HCl (*aq*) form $CaCl_2$ = 1 mol

0.19 mol of HCl (*aq*) would form $CaCl_2 = \frac{1}{2} \times 0.19$

= 0.095 mol

Mass of $CaCl_2$ formed = 0.095 × 111 **= 10.54 g**

Practice Problems

74. What mass of solid AgCl is obtained when 25 ml of 0.068 M $AgNO_3$ reacts with excess of aqueous HCl?

75. What volume of 0.34 M KOH is sufficient to react with 20 ml of 0.15 M H_2SO_4 solution ?

76. Calculate the volume of 1.00 mol L^{-1} aqueous sodium hydroxide that is neutralised by 200 mL of 2.00 mol L^{-1} aqueous hydrochloric acid and the mass of sodium chloride produced. Neutralization reaction is

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

N.C.E.R.T.

77. In a reaction vessel 0.184 g of NaOH is required to be added for completing the reaction. How many millilitre of 0.150 M NaOH should be added for this requirement ?

78. 500 mL of 0.250 M Na_2SO_4 solution is added to an aqueous solution of 15.00 g of $BaCl_2$ resulting in the formation of the white precipitate of $BaSO_4$. How many grams of $BaSO_4$ are formed ? ***N.C.E.R.T.***

79. Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction :

$$CaCO_3\,(s) + 2HCl\,(aq) \longrightarrow CaCl_2\,(aq) + CO_2\,(g) + H_2O\,(l)$$

What mass of $CaCO_3$ is required to react completely with 25 mL of 0.75 M HCl ?

Answers to Practice Problems

74. 0.24 g **75.** 17.65 mL **76.** 400 mL, 23.4 g

77. 30.7 mL

78. 16.78 g

79. 0.9375 g

Hints & Solutions on page 79

add on

Conceptual Questions 3

CONCEPTUAL QA

Q. 1. Give two examples of molecules having molecular formula same as empirical formula.

Ans. CO_2, CH_4.

Q. 2. Give an example of molecule in which

(*i*) Ratio of molecular formula and empirical formula is 6 : 1.

(*ii*) Molecular weight is two times of the empirical formula weight.

(*iii*) The empirical formula is CH_2O and ratio of molecular formula weight and empirical formula weight is 6.

Ans. (*i*) C_6H_6 (*ii*) H_2O_2 (*iii*) $C_6H_{12}O_6$

Q.3. Write the empirical formula of *(i)* glucose *(ii)* sucrose.

Ans. (*i*) CH_2O (*ii*) $C_{12}H_{22}O_{11}$

Q. 4. What are the the SI units of molarity ?

Ans. mol dm^{-3}

Q. 5. 1.615 g of anhydrous $ZnSO_4$ was left in moist air. After a few days its weight was found to be 2.875 g. What is the molecular formula of hydrated salt ?

(At. masses : Zn = 65.5, S = 32, O = 16, H = 1)

Ans. Molecular mass of anhydrous $ZnSO_4$

$$= 65.5 + 32 + 4 \times 16 = 161.5$$

1.615 g of anhydrous $ZnSO_4$ combines with water

$$= 2.875 - 1.1615 = 1.260 \text{ g}$$

1.615 g of anhydrous $ZnSO_4$ combine with water = 1.260 g

161.5 g of anhydrous $ZnSO_4$ combine with $= \dfrac{1.260}{1.615} \times 161.5 = 126$ g

No. of moles of water $= \dfrac{126}{18} = 7$

Formula : $ZnSO_4.7H_2O$.

Q. 6. Desnsity of water 1000 kg m^{-3} corresponds to .... g cm^{-3}. Complete the statement.

Ans. 1.

Q.7 How are 0.50 m Na_2CO_3 and 0.50 M Na_2CO_3 different ?

Ans. 0.50 m Na_2CO_3 means that 0.50 moles of Na_2CO_3 are dissolved in 1000 g of water.

0.50 M Na_2CO_3 solution means that 0.50 moles of Na_2CO_3 are dissolved in 1000 mL of solution.

Q. 8. Calculate the amount of carbon dioxide that could be produced when

(*i*) 1 mol of carbon is burnt in air

(*ii*) 1 mol of carbon is burnt in 16 g dioxygen

(*iii*) 2 moles of carbon is burnt in 16 g of dioxygen.

Ans. $C + O_2 \longrightarrow CO_2$

1 mol of carbon reacts with 1 mol of oxygen to form 1 mole of CO_2.

(*i*) 1 mol of CO_2 or 44 g of CO_2.

(*ii*) Since 16 g of dioxygen *i.e.* 0.5 mol of O_2 are present, it is a limiting reagent. 0.5 mol of O_2 will form 0.5 mol of CO_2 *i.e.* 22 g.

(*iii*) 16 g or 0.5 mol of O_2 is limiting reagent. 0.5 mol of O_2 will form 0.5 mol of CO_2 i.e. 22 g

Q. 9. In the reaction :

$$2A + 4B \longrightarrow 3C + 4D$$

when 5 moles of A react with 6 moles of B, then

(*i*) Which is the limiting reagent ?

(*ii*) Calculate the amount of C formed.

Ans. According to the equation, 2 moles of A require 4 moles of B for the reaction.

For 5 moles of A, the moles of B required

$$= \frac{4}{2} \times 5 = 10 \text{ mol}.$$

But we have only 6 moles of B.

Hence,

(*i*) B is the limiting reagent.

(*ii*) 4 moles of B give 3 mole of C.

$\therefore$ 6 moles of B will give $= \frac{3}{4} \times 6 = 4.5 \text{ mol of C}$.

Advanced Level PROBLEMS

Accelerate Your Potential (for JEE Advance)

Problem 8. 1.0 g of a mixture of carbonates of calcium and magnesium gave 240 mL of CO_2 at N.T.P. Calculate the percentage composition of the mixture.

Solution

$$\underset{100\text{ g}}{CaCO_3} \longrightarrow CaO + \underset{22400\text{ mL}}{CO_2}$$

$$\underset{84\text{ g}}{MgCO_3} \longrightarrow MgO + \underset{22400\text{ mL}}{CO_2}$$

Let $CaCO_3$ present in the mixture = x g

Now 100 g of $CaCO_3$ give CO_2 = 22400 mL

x g of $CaCO_3$ will give $CO_2 = \frac{22400 \times x}{100} = 224\,x$ g

Similarly, 84 g of $MgCO_3$ give CO_2 = 22400 mL

$(1-x)$ g of $MgCO_3$ give $CO_2 = \frac{22400 \times (1-x)}{84} = 266.7\,(1-x)$

Total CO_2 evolved at N.T.P. = 240 mL

Now, $224\,x + 266.7\,(1-x) = 240$

$224x - 266.7x = 240 - 266.7$

$-42.7x = -26.7$

$$x = \frac{26.7}{42.7} = 0.625$$

Wt. of $CaCO_3$ = 0.625 g

Wt. of $MgCO_3$ = 0.375 g

$$\% \, CaCO_3 = \frac{0.625}{1} \times 100 = 62.5$$

$$\% \, MgCO_3 = \frac{0.375 \times 100}{1} = \mathbf{37.5}$$

Problem 9. 1.0 g of an alloy of aluminium and magnesium when treated with excess of dil. HCl form magnesium chloride and aluminium chloride and hydrogen collected over mercury at 0°C has a volume of 1.20 L at 0.92 atmospheric pressure. Calculate the composition of the alloy.

Solution Let us first calculate the volume of H_2 at N.T.P.

Applying $\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$

$p_1 = 0.92$ atm $p_2 = 1$ atm

$V_1 = 1.20$ L $V_2 = ?$

$T_1 = 273$ K $T_2 = 273$ K

$$\therefore \quad V_2 = \frac{p_1V_1T_2}{p_2T_1} = \frac{0.92 \times 1.20 \times 273}{1 \times 273} = 1.104 \text{ L}$$

Let mass of Al in alloy = x g

Mass of Mg in alloy = $(1-x)$ g

$$\underset{\substack{2 \times 27 \\ = 54\text{ g}}}{2Al} + 6HCl \longrightarrow AlCl_3 + \underset{\substack{3 \times 22.4 \\ = 67.2\text{ L}}}{3H_2}$$

$$\underset{24\text{ g}}{Mg} + 2HCl \longrightarrow MgCl_2 + \underset{22.4\text{ L}}{H_2}$$

54 g of Al give H_2 at N.T.P. = 67.2 L

x g of Al give H_2 at N.T.P. $= \frac{67.2 \times x}{54}$

Similarly, 24 g of Mg give H_2 at N.T.P.= 22.4 L

$(1-x)$ g of Mg will give H_2 at N.T.P.$= \frac{22.4}{24} \times (1-x)$

Total H_2 liberated

$$\frac{67.2x}{54} + \frac{22.4(1-x)}{24} = 1.104$$

Solving for x, we get

$x = 0.5486$ g

$\therefore$ Mass of Al in the alloy = 0.5486 g

$$\% \text{ of Al} = \frac{0.5486}{1} \times 100 = \mathbf{54.86}$$

% of Mg in alloy = 100 – 54.86 = **45.14.**

Problem 10. A mixture of sodium iodide and sodium chloride when treated with sulphuric acid gave sodium sulphate equal to the weight of the original mixture. Find the percentage composition of the mixture.

Solution

$$\underset{\substack{2(23+127) \\ = 300}}{2NaI} + H_2SO_4 \longrightarrow \underset{\substack{46+32+64 \\ = 142}}{Na_2SO_4} + 2HI$$

$$2NaCl + H_2SO_4 \longrightarrow Na_2SO_4 + 2HCl$$
$$\underset{= 117}{2(23+35.5)} \qquad \underset{= 142}{46+32+64}$$

$$\begin{aligned}
\text{Let the wt. of mixture} &= 1 \text{ g} \\
\text{Wt. of NaI in the mixture} &= x \text{ g} \\
\text{Wt. of NaCl in the mixture} &= (1-x) \text{ g} \\
\text{Now, 300 g of NaI give } Na_2SO_4 &= 142 \text{ g} \\
x \text{ g of NaI give } Na_2SO_4 &= \frac{142}{300} \times x
\end{aligned}$$

Similarly,

$$\begin{aligned}
\text{117 g of NaCl give } Na_2SO_4 &= 142 \text{ g} \\
(1-x) \text{ g of NaCl give } Na_2SO_4 &= \frac{142}{117} \times (1-x)
\end{aligned}$$

Weight of Na_2SO_4 formed = Wt. of original mixture

$$\frac{142}{300} x + \frac{142}{117} \times (1-x) = 1$$

Solving for x, we get

$$\begin{aligned}
x &= 0.2886 \text{ g} \\
\therefore \quad \text{Wt. of NaI} &= 0.2886 \\
\text{\% NaI in mixture} &= \frac{0.2886}{1} \times 100 = \mathbf{28.86\%} \\
\text{\% NaCl in mixture} &= 100 - 28.86 = \mathbf{71.14\%}
\end{aligned}$$

Problem 11. How many mL of H_2SO_4 of density 1.8 g/mL containing 92.5% by volume of H_2SO_4 should be added to 1 litre of 40% solution of H_2SO_4 (density 1.30 g/mL) in order to prepare 50% solution of H_2SO_4 (density 1.4 g/mL) ?

Solution Molarity of solution containing 92.5% H_2SO_4 may be calculated as :

$$\text{Molarity } (M_1) = \frac{\text{Vol. of } H_2SO_4 \times \text{Density} \times 1000}{98 \times 100}$$
$$= \frac{92.5 \times 1.8 \times 1000}{98 \times 100} = 16.99 \text{ M}$$

Molarity of solution containing 40% H_2SO_4,

$$M_2 = \frac{40 \times 1.3 \times 1000}{98 \times 100} = 5.31 \text{ M}$$

Molarity of required solution containing 50% H_2SO_4,

$$M = \frac{50 \times 1.4 \times 1000}{98 \times 100} = 7.14 \text{ M}$$

Let V litre of solution with molarity (M_1) is added to 1 litre of solution with molarity (M_2) to prepare (1 + V) litre of solution with molarity M, then

$$\begin{aligned}
M_1V_1 + M_2V_2 &= MV \\
16.99 \times V + 5.31 \times 1 &= 7.14\,(1 + V) \\
9.85V &= 1.83 \\
V &= \frac{1.83}{9.85} = 0.186 \text{ L} \\
\text{or} \quad V &= 0.186 \text{ litre or } = \mathbf{186 \text{ mL.}}
\end{aligned}$$

Problem 12. You are given one litre of 0.15 M HCl and one litre of 0.40 M HCl. What is the maximum volume of 0.25 M HCl which you can make from these solutions without adding any water ?

Solution Since no water is added, the volume of 0.25 M HCl cannot be more than 2 litres. Let x litre of 0.40 M HCl (more concentrated) be added to 1 litre of 0.15 M HCl.

Applying molarity equation :

$$\begin{aligned}
M_1V_1 + M_2V_2 &= M_3V_3 \\
0.15 \times 1 + 0.40 \times x &= 0.25 \times (1 + x)
\end{aligned}$$

(Total volume V_3 becomes $1 + x$ litres)

$$\begin{aligned}
0.15 + 0.40\,x &= 0.25 + 0.25\,x \\
0.40x - 0.25\,x &= 0.25 - 0.15 \\
0.15\,x &= 0.10 \\
x &= \frac{0.10}{0.15} = 0.667 \text{ litre} \\
\text{Total volume of 0.25 M solution} &= 1 + x \\
&= 1 + 0.667 = \mathbf{1.667 \text{ L}}
\end{aligned}$$

Problem 13. A mixture of FeO and Fe_3O_4 when heated in air to a constant mass gains 5% by mass. Calculate the composition of the mixture (atomic mass of Fe = 55.8).

Solution The chemical equations are:

$$2FeO + \frac{1}{2} O_2 \longrightarrow Fe_2O_3 \qquad \dots (i)$$
$$\underset{= 143.6}{2\,(55.8 + 16)} \qquad \underset{= 159.6}{2 \times 55.6 + 3 \times 16}$$

$$2Fe_3O_4 + \frac{1}{2} O_2 \longrightarrow 3Fe_2O_3 \qquad \dots (ii)$$
$$\underset{= 462.8}{2\,(55.6 \times 3 + 4 \times 16)} \qquad \underset{= 478.8}{3(2 \times 55.6 + 3 \times 16)}$$

$$\begin{aligned}
\text{Let the weight of mixture} &= 100 \text{ g} \\
\text{Weight of FeO} &= x \text{ g} \\
\text{Weight of } Fe_3O_4 &= (100 - x) \text{ g}
\end{aligned}$$

According to equation (i),

$$\begin{aligned}
\text{143.6 g of FeO produce } Fe_2O_3 &= 159.6 \text{ g} \\
x \text{ g of FeO produce } Fe_2O_3 &= \frac{159.6}{143.6} \times x
\end{aligned}$$

According to equation (ii)

$$\begin{aligned}
\text{462.8 g of } Fe_3O_4 \text{ produce } Fe_2O_3 &= 478.8 \text{ g} \\
(100 - x) \text{ g of } Fe_3O_4 \text{ produce } Fe_2O_3 &= \frac{478.8}{462.8} \times (100 - x)
\end{aligned}$$

Total mass of Fe_2O_3 formed by heating

$$\frac{159.6}{143.6} x + \frac{478.8}{462.8} (100 - x)$$

$$\begin{aligned}
\text{Gain in weight of mixture} &= \frac{100 \times 5}{100} = 5 \text{ g} \\
\text{Total } Fe_2O_3 \text{ formed} &= 100 + 5 = 105 \text{ g} \\
\therefore \frac{159.6}{143.6} x + \frac{478.8}{462.8} (100 - x) &= 105 \\
\text{Solving, we get} \quad x &= 19.92 \\
\therefore \quad \text{Mass of FeO} &= 19.92 \\
\text{Mass of } Fe_3O_4 &= 100 - 19.92 = 80.08 \\
\text{Mass \% FeO} &= \frac{19.92}{100} \times 100 = \mathbf{19.92} \\
\text{Mass \% } Fe_3O_4 &= \mathbf{80.08.}
\end{aligned}$$

Problem 14. Gastric juice contains about 3 mg of HCl per millilitre. If a person produces about 225 mL of gastric juice per day, how many antacid tablets each containing 250 mg of $Al(OH)_3$ are needed to neutralise all the HCl produced in one day ?

Solution HCl and $Al(OH)_3$ react as

$$\underset{\substack{36.5 \times 3 \\ = 109.5}}{3HCl} + \underset{78}{Al(OH)_3} \longrightarrow AlCl_3 + 3H_2O$$

Volume of gastric juice produced per day = 225 mL

Mass of HCl produced per day $= 225 \times 3 \times 10^{-3}$
$= 0.675$ g

Now 109.5 g of HCl require $Al(OH)_3$ = 78 g

0.675 g of HCl will require $Al(OH)_3 = \dfrac{78}{109.5} \times 0.675$
$= 0.481$ g

Number of tablets required $= \dfrac{0.48}{0.250} = 1.92$
= 2 tablets.

Problem 15. A mixture of formic acid and oxalic acid is heated with concentrated H_2SO_4. The gas produced is collected and on its treatment with KOH solution, the volume of the solution decreases by 1/6th. Calculate the molar ratio of the two acids in the original mixture.

Solution Let x moles of formic acid and y moles of oxalic acid are heated.

$$\underset{x \text{ mol}}{HCOOH} \xrightarrow{\text{Conc. } H_2SO_4} H_2O + \underset{x \text{ mol}}{CO}$$

$$\underset{y \text{ mol}}{\begin{array}{c} COOH \\ | \\ COOH \end{array}} \xrightarrow{\text{Conc. } H_2SO_4} H_2O + \underset{y \text{ mol}}{CO} + \underset{y \text{ mol}}{CO_2}$$

Total moles of gaseous mixture
= moles of CO + moles of CO_2
$= x + y + y = x + 2y$

KOH absorbs only CO_2 *i.e.* volume occupied by (y) moles. Under same conditions, ratios of volumes will be proportional to molar ratio of gases

$$\frac{\text{Moles of } CO_2}{\text{Moles of both gases}} = \frac{y}{x + 2y} = \frac{1}{6}$$

$6y = x + 2y$ or $4y = x$

or $\dfrac{x}{y} = 4$

$\therefore$ Ratio of HCOOH : $\begin{array}{c} COOH \\ | \\ COOH \end{array}$ = **4 : 1.**

Problem 16. (*a*) Calculate the number of chloride ions in 100 mL of 0.01 M $AlCl_3$ solution.
(*b*) What will be the change in number of chloride ions if the solution is diluted by 100 mL water?

Solution (*a*) Molarity $= \dfrac{\text{Moles of } AlCl_3}{\text{Vol. of solution}} \times 1000$

$\therefore$ Moles of $AlCl_3 = \dfrac{\text{Molarity} \times \text{Vol. of Solution}}{1000}$

$= \dfrac{0.01 \times 100}{1000} = 0.001$ mol

Now, 1 molecule of $AlCl_3$ contains 3 Cl^- ions

$\therefore$ 1 mol of $AlCl_3$ contains 3 mol of Cl^- ions

0.001 mol of $AlCl_3$ contains Cl^- ions $= 0.001 \times 3$
$= 0.003$ mol

1 mol of Cl^- ions $= 6.022 \times 10^{23}$ Cl^- ions

0.003 mol of Cl^- ions $= 6.022 \times 10^{23} \times 0.003$
$= \mathbf{1.81 \times 10^{21}}$ Cl^- ions

(*b*) No change in number of Cl^- ions.

Solution File — Hints & Solutions for Practice Problems

3. Each side of cube = 76 mm = 7.6 cm

Volume of cubic block of ice = (side)^3
$= (7.6 \text{ cm})^3 = 439 \text{ cm}^3$

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

$\therefore$ Mass = Density × Volume
$= (0.921 \text{ g cm}^{-3}) \times (439 \text{ cm}^3)$
= 404.3 or $\mathbf{4.0 \times 10^2}$ **g**

Rounded off to two significant figures as in 7.6.

5. (*a*) $\dfrac{108}{7.2}$ =15. (upto 2 significant numbers)

(*b*) $(1.6 \times 10^2)^2 = 2.56 \times 10^4$

or $= 2.6 \times 10^4$ (upto 2 significant numbers)

(*c*) $\dfrac{(1.35 \times 10^{-6})(0.4)}{5.6} = 0.096 \times 10^{-6} = 0.1 \times 10^{-6}$

or $= 1 \times 10^{-7}$ (upto one significant figure)

(*d*) $\dfrac{3.25 \times 0.08621}{4.002} = 0.0700$ (upto 3 significant figures)

(*e*) (1.0042 – 0.0034) (1.23)

(1.0008) (1.23) = 1.23 (upto 3 significant figures)

6. (*i*) Area of square = (side)^2
$= (1.2)^2 = 1.44 \text{ m}^2$

Correct answer = $\mathbf{1.4 \ m^2}$ (upto 2 significant places).

(*ii*) Volume of sphere $= \dfrac{4}{3}\pi r^3$

Solution File

$= \frac{4}{3} \times \frac{22}{7} \times (1.6)^3$

$= 17.164$

$=$ **17 cm³** (upto two significant figures).

(*iii*) Length of rectangle $= \frac{\text{Area}}{\text{Breadth}} = \frac{10.25}{2.5}$

$=$ **4.1 m** (upto 2 significant figures).

7. (*i*) $(1.20 \times 10^{-6}) + (6.00 \times 10^{-5})$

$0.120 \times 10^{-5} + 6.00 \times 10^{-5} = \mathbf{6.12 \times 10^{-5}}$.

(upto second place of decimal as in 6.00×10^{-5})

(*ii*) $(2.164 \times 10^5)^{1/2} = (21.64 \times 10^4)^{1/2} = 4.6519 \times 10^2$

$= \mathbf{4.652 \times 10^2}$ (upto 4 significant figures)

(*iii*) $(9.13 \times 10^{-2})(7.006 \times 10^{-3}) = 63.9648 \times 10^{-5}$

$= \mathbf{64.0 \times 10^{-5}}$ (upto 3 significant figures)

or $= 6.40 \times 10^{-4}$

(*iv*) $4.00 \times 10^{-2} + 3.26 \times 10^{-3} + 1 \times 10^{-6}$

4.00×10^{-2}

0.326×10^{-2}

$\frac{0.0001 \times 10^{-2}}{4.3261 \times 10^{-2}} = \mathbf{4.33 \times 10^{-2}}$

(upto second place of decimal).

8. (*i*) $\frac{2.36 \times 0.07251}{2.130}$ = upto **three significant figures** because 2.36 contains 3 significant figures.

(*ii*) $\frac{(28.2-21.2)(1.79 \times 10^6)}{1.62}$ = upto **two significant figures** because difference of 28.2 – 21.2 *i.e.* 7.0 contains 2 significant figures.

10. $(3.0 \times 10^8 \text{ ms}^{-1}) \times (2.00 \times 10^{-9}\text{s}) = 6.0 \times 10^{-1} = 0.60$ m answer should be in two significant figures as in 3.0×10^8.

11. Unit factor $= \frac{0.1(\text{nm})}{1(\text{Å})}$

$\therefore$ 5896 Å $= 5896\text{Å} \times \frac{0.1(\text{nm})}{1(\text{Å})}$

$=$ **589.6 nm.**

12. 1m = 10^2 cm

$1\text{m}^3 = (10^2 \text{ cm})^3 = \mathbf{10^6 \text{ cm}^3}$.

13. (*i*) 1kg = 1000 g; 1 litre = 10^3 mL

Unit factor : 1 $= \frac{1000\text{g}}{\text{kg}}$; Unit factor : 1 $= \frac{1 \text{ litre}}{10^3\text{mL}}$

$\therefore \frac{4.86 \text{ kg}}{\text{L}} = \frac{4.86 \text{ kg}}{\text{L}} \times \frac{1000 \text{ g}}{\text{kg}} \times \frac{\text{litre}}{10^3\text{mL}} = \mathbf{4.86 \text{ g/mL}}$.

(*ii*) 1.86 km to cm

1 km = 10^5 cm

Unit factor : 1 $= \frac{10^5 \text{ cm}}{(\text{km})}$

1.86 km = 1.86 km $\times \frac{10^5 \text{ cm}}{(\text{km})} = \mathbf{1.86 \times 10^5 \text{ cm}}$.

(*iii*) 1 m = 10^6 μm

Unit factor : 1 $= \frac{10^6 \ \mu\text{m}}{\text{m}}$

$\therefore 6.92 \times 10^{-7} \text{ m} \times \frac{10^6 \ \mu\text{m}}{\text{m}} = \mathbf{0.692 \ \mu m}$

6.92 × 10⁻⁷ m to Angstroms

1Å = 10^{-10} m

Unit factor : 1 $= \frac{1\text{Å}}{10^{-10}\text{m}}$

$\therefore 6.92 \times 10^{-7} \text{ m} \times \frac{1\text{Å}}{10^{-10}\text{m}} = \mathbf{6920Å}$.

(*iv*) 1 litre $=10^3$ cm³

Unit factor : 1 $= \frac{1 \text{ litre}}{10^3\text{cm}^3}$

$9.2 \times 10^{-3} \text{ cm}^3 = 9.2 \times 10^{-3} \text{ cm}^3 \times \frac{1 \text{ litre}}{10^3\text{cm}^3}$

$= \mathbf{9.2 \times 10^{-6}}$ **litre.**

14. Capacity of tank = length × breadth × depth

$= 0.8 \text{ m} \times 0.1 \text{ m} \times 50 \times 10^{-3}$ m

$= 0.004$ or 4×10^{-3} m³

1 litre = 10^{-3} m³

Unit factor : 1 $= \frac{1 \text{ litre}}{10^{-3}\text{m}^3}$

$\therefore 4 \times 10^{-3} \text{ m}^3 = 4 \times 10^{-3} \text{ m}^3 \times \frac{1 \text{ litre}}{10^{-3}\text{m}^3} =$ **4 litre.**

15. 100 L = 100×10^3 cm³ (1L = 10^3 cm³)

$= \mathbf{10^5 \text{ cm}^3}$.

16. (*i*) 500 mg in kilogram

500 mg = 500×10^6 g

1 kg = 10^3 g

Unit factor : 1 $= \frac{1\text{kg}}{10^3\text{g}}$

$500 \times 10^6 \ g = 500 \times 10^6 \times \frac{1\text{kg}}{10^3\text{g}} = 500 \times 10^3$ kg

$= \mathbf{5.0 \times 10^5}$ **kg.**

(*ii*) 1 kg = 10^3 g

Unit factor : 1 $= \frac{1\text{kg}}{10^3\text{g}}$

$\therefore 3.34 \times 10^{-24} \text{ g} = 3.34 \times 10^{-24} \text{ g} \times \frac{1\text{kg}}{10^3\text{g}}$

$= \mathbf{3.34 \times 10^{-27}}$ **kg.**

17. (*i*) 1 pm = 10^{-12} m, 1m = 10^3 mm

$\therefore$ 1pm = 10^{-12} m = 10^{-9} mm

Unit factor : 1 $= \frac{1 \text{ pm}}{10^{-9}\text{mm}}$

1 μs = 10^{-6} s

Solution File

$$\text{Unit factor : } 1 = \frac{1\ \mu s}{10^{-6} s} \text{ or } = \frac{10^{-6} s}{1\ \mu s}$$

$$1.54\ mms^{-1} = 1.54\ mms^{-1} \times \frac{1\ pm}{10^{-9} mm} \times \frac{10^{-6} s}{1\ \mu s}$$

$$= \mathbf{1.54 \times 10^3\ pm\ \mu s^{-1}.}$$

(*ii*) $1g = 1000\ mg$

$$\text{Unit factor : } 1 = \frac{1000\ mg}{1g}$$

$$10\ dL = 1L$$

$$\text{Unit factor : } 1 = \frac{10\ dL}{1L} = \frac{1dL^{-1}}{10\ L^{-1}}$$

$$\therefore\ 25\ g\ L^{-1} = 25 gL^{-1} \times \frac{1000\ mg}{1g} \times \frac{1dL^{-1}}{10\ L^{-1}}$$

$$= \mathbf{2.5 \times 10^3 mg\ dL^{-1}.}$$

(*iii*) $1L = 1dm^3$

$$\text{Unit factor : } 1 = \frac{1dm^3}{1L}$$

$$10^3 dm^3 = 1m^3$$

$$\text{Unit factor : } 1 = \frac{1m^3}{10^3 dm^3}$$

$$\therefore\ 25\ L = 25\ L \times \frac{1dm^3}{1L} \times \frac{1m^3}{10^3 dm^3} = \mathbf{2.5 \times 10^{-2}\ m^3.}$$

(*iv*) $1\mu g = 10^{-6}\ g$

$$\text{Unit factor : } 1 = \frac{1\mu g}{10^{-6} g}$$

$$1\ m^3 = 10^6\ cm^3$$

$$\text{Unit factor : } 1 = \frac{10^6 cm^3}{1m^3}$$

$$1\ m^3 = (10^6)^3 \mu m$$

$$\text{Unit factor : } 1 = \frac{10^{-18} \mu m}{1m^3} \text{ or } \frac{1m^3}{10^{18} \mu m^3}$$

$$\therefore\ \frac{2.66\ g}{cm^3} \times \frac{1\ \mu g}{10^{-6} g} \times \frac{10^6 cm^3}{1m^3} \times \frac{1m^3}{10^{18} \mu m}$$

$$= \mathbf{2.66 \times 10^{-6}\ \mu g\ \mu m^{-3}.}$$

(*v*) $1L = 1000\ mL$

$$\text{Unit factor : } 1 = \frac{1000\ mL}{1L}$$

$$1h = 3600\ s$$

$$\text{Unit factor : } 1 = \frac{1h}{3600\ s}$$

$$\therefore\ \frac{4.2\ L}{h^2} = \frac{4.2\ L}{h^2} \times \left(\frac{1000\ mL}{1L}\right) \times \left(\frac{1\ h}{3600\ s}\right)^2$$

$$= \mathbf{3.2 \times 10^{-4}\ mL\ s^{-2}}$$

18. (*i*) $1\ nm = 10^{-9}\ m$

$$\text{Unit factor : } 1 = \frac{10^{-9} m}{1\ nm}$$

$$\therefore\ 7\ nm = 7nm \times \left(\frac{10^{-9} m}{1\ nm}\right) = \mathbf{7 \times 10^{-9}\ m.}$$

(*ii*) $1\ pm = 10^{-12}\ m$

$$1 = \frac{10^{-12} m}{1\ pm}$$

$$\therefore\ 41\ pm = 41\ pm \times \left(\frac{10^{-12} m}{1\ pm}\right) = \mathbf{41 \times 10^{-12}\ m.}$$

19.
$$\text{Unit factor : } 1 = \frac{24\ h}{1\ day}$$

$$1\ hr = 60\ min$$

$$\text{Unit factor : } 1 = \frac{60\ min}{1\ hr}$$

$$1\ min = 60s$$

$$\text{Unit factor : } 1 = \frac{60s}{1\ min}$$

$$\therefore\ 2\ days = 2day \times \left(\frac{24\ h}{1\ day}\right) \times \left(\frac{60\ min}{1\ h}\right) \times \left(\frac{60\ s}{1\ min}\right)$$

$$= \mathbf{172800\ s}$$

20. (*i*) $1\ pm = 10^{-12}\ m$

$$\text{Unit factor : } 1 = \frac{10^{-12}}{1\ pm}$$

$$28.7\ pm = 28.7\ pm \times \left(\frac{10^{-12} m}{1\ pm}\right) = \mathbf{2.87 \times 10^{-11} m}$$

(*ii*) $1\ \mu s = 10^{-6}\ s$

$$\text{Unit factor : } 1 = \frac{10^{-6} s}{1\ \mu s}$$

$$\therefore\ 15.15\ \mu s = 15.15 \mu s \times \left(\frac{10^{-6} s}{1\ \mu s}\right) = \mathbf{1.515 \times 10^{-5}\ s}$$

(*iii*) $1\ mg = 10^{-6}\ g$

$$\text{Unit factor : } 1 = \frac{10^{-6} kg}{1\ mg}$$

$$\therefore\ 25365\ mg = 25365\ mg \times \left(\frac{10^{-6} kg}{1\ mg}\right)$$

$$= \mathbf{2.5365 \times 10^{-2}\ kg}$$

21. Mass of reactants = 4.2 g + 10.0 g = 14.2 g
Mass of products = 2.2 g + 12.0 g = 14.2 g
∴ Mass of reactants = Mass of products

22. In H_2O_2,
Mass of hydrogen = 5.93 g
Mass of oxygen = 100 – 5.93 = 94.07 g
In water,
Hydrogen = 11.2 g
Oxygen = 100 – 11.2 = 88.8 g

Solution File

Let us fix the mass of hydrogen = 1.0 g

In H_2O_2,

Mass of oxygen combining with 5.93 g of hydrogen = 94.07 g

Mass of oxygen combining with 1 g of hydrogen

$$= \frac{94.07}{5.93} = \mathbf{15.86\ g}$$

In H_2O,

Mass of oxygen combining with 11.2 g of hydrogen = 88.8 g

Mass of oxygen combining with 1 g of hydrogen

$$= \frac{88.8}{11.2} = 7.93\ g$$

The ratio of masses of oxygen combining with 1g of hydrogen

15.86 : 7.93

2 : 1

This is simple ratio and therefore, it illustrates the law of multiple proportions.

23. Let us fix the mass of A to be 1.0 g

In compound X,

0.3 g of A combines with B = 0.4 g

$$1.0 \text{ g of A will combine with B} = \frac{0.4}{0.3} \times 1.0 = \mathbf{1.33\ g}$$

In compound Y,

18 g of A combine with B = 48 g

$$1 \text{ g of A will combine with B} = \frac{48}{18} \times 1 = \mathbf{2.66\ g}$$

The ratio of masses of B which combine with fixed weight of A (*i.e.* 1 g).

1.33 : 2.66

1 : 2

This is a simple ratio and hence it illustrates the law of multiple proportions.

24. Let us fix the mass of carbon as 1 g.

In compound A,

75 g of carbon combine with hydrogen = 25 g

1 g of carbon combines with hydrogen

$$= \frac{25}{75} = \mathbf{0.333\ g}$$

In compound B,

85.7 g of carbon combine with hydrogen = 14.3 g

1 g of carbon combines with hydrogen

$$= \frac{14.3}{85.7} = \mathbf{0.167\ g}$$

In compound C,

92.3 g of carbon combine with hydrogen = 7.7 g

1 g of carbon combines with hydrogen

$$= \frac{7.7}{92.3} = \mathbf{0.083\ g}$$

The ratio of masses of hydrogen which combine with fixed mass of carbon (i.e. 1 g)

0.333 : 0.167 : 0.083

4 : 2 : 1

This is simple whole number ratio and therefore, it illustrates the law of multiple proportions.

25. *First experiment*

Mass of copper oxide = 1.375 g

Copper left = 1.098 g

Mass of oxygen present = 1.375 – 1.098 = 0.277 g

$$\% \text{ of oxygen in CuO} = \frac{0.277 \times 100}{1.375} = \mathbf{20.14}$$

Second experiment

Mass of copper taken = 1.178 g

Mass of CuO formed = 1.476 g

Mass of oxygen present = 1.476 – 1.178 = 0.298

$$\% \text{ of oxygen in CuO} = \frac{0.298 \times 100}{1.476} = \mathbf{20.19}$$

Percentage of oxygen in the two compounds is the same, therefore, these results illustrate the law of constant composition.

26. (*i*) 22400 mL of SO_2 at N.T.P = 64 g

$$224 \text{ mL of } SO_2 \text{ at N.T.P} = \frac{64}{22400} \times 224 = 0.64 \text{ g}$$

Mass of sulphur present = 0.32 g

Mass of oxygen present = 0.64 – 0.32 = 0.32 g

$$\% \text{ of sulphur} = \frac{0.32}{0.64} \times 100 = \mathbf{50\%}$$

(*ii*) Percentage of sulphur in second sample = **50%**

Since percentage of sulphur in two compounds is the same, therefore, these results illustrate the law of definite proportions.

27. Mass of Cu = 25.45%

Mass of H_2O = 36.07%

100 g of crystalline copper sulphate sample contain copper = 25.45

40 g of crystalline copper sulphate sample will contain copper $= \frac{25.45}{100} \times 40 = \mathbf{10.18\ g}$

28. *In copper sulphide,*

Percentage of Cu = 66.5

Percentage of sulphur = 100 – 66.5 = 33.5

66.5 g of Cu combine with 33.5 g of sulphur in copper sulphide.

In copper oxide,

Percentage of Cu = 79.9

Percentage of oxygen = 100 – 79.9 = 20.1

Solution File

79.9 g of Cu combine with 20.1 g of oxygen in copper oxide.

Let us fix the mass of copper as 1 g

In copper sulphide,

66.5 g of copper combine with sulphur = 33.5 g

1 g of copper will combine with sulphur

$$= \frac{33.5}{66.5} = \mathbf{0.503\ g}$$

In copper oxide,

79.9 g of Cu combine with oxygen =2 0.1 g

1 g of Cu will combine with oxygen

$$= \frac{20.1}{79.9} = \mathbf{0.251\ g}$$

The ratio of weights of sulphur and oxygen which combine with the fixed mass of copper (1 g)

0.503 : 0.251

2 : 1 ...(*i*)

In sulphur trioxide,

Percentage of sulphur = 40

Percentage of oxygen = 60

The ratio of sulphur and oxygen in sulphur trioxide

40 : 60

2 : 3 ...(*ii*)

The ratios (*i*) and (*ii*) are

$$\frac{2}{1} : \frac{2}{3} \text{ or } \mathbf{3:1}$$

This is a simple ratio. Hence law of reciprocal proportions is illustrated.

29. *Calculation of weights of oxygen in each oxide.*

Mass of each oxide taken = 1 g

18 g of H_2O ≡16 g of oxygen

∴ 0.1254 g of water in first oxide contains oxygen

$$= \frac{16}{18} \times 0.1254 = 0.1115 \text{ g}$$

0.2263 g of water in second oxide contains oxygen

$$= \frac{16}{18} \times 0.2263 = 0.2011 \text{ g}$$

In first oxide

Mass of oxygen = 0.1115 g

Mass of metal =1 – 0.1115 = 0.8885 g

In second oxide

Mass of oxygen = 0.2011 g

Mass of metal = 1 – 0.2011 = 0.7989 g

Let us fix the weight of metal as 1 g

First oxide

Mass of oxygen which combines with 0.8885 g of metal = 0.1115 g

Mass of oxygen which combines with 1 g of metal

$$= \frac{0.1115}{0.8885} \times 1 = \mathbf{0.1255\ g}$$

Second oxide

Mass of oxygen which combines with 0.7989 g of metal = 0.2011 g

Mass of oxygen which combines with 1 g of metal

$$= \frac{0.2011}{0.7989} = \mathbf{0.2517\ g}$$

The ratio of masses of oxygen which combine with fixed mass of metal (1 gram)

0.1255 : 0.2517

1 : 2

Since this is a simple ratio, therefore the above data illustrates the law of multiple proportions.

30. Average atomic mass of Si

$$= \frac{28 \times 92.25 + 29 \times 4.65 + 30 \times 3.10}{100}$$

$$= \mathbf{28.11\ u.}$$

31. (*a*) 1.6 gram atoms of oxygen = 16 × 1.6 = **25.6 g.**

(*b*) 5.6 gram atoms of sulphur $= \frac{32}{1} \times 5.6$

$= \mathbf{179.2\ g.}$

(*c*) 2.4 gram atoms of iodine = 127 × 2.4 = **304.8 g.**

32. (*i*) 2.5 gram molecules of H_2S = 34 × 2.5 = **85 g.**

(*ii*) 3.6 gram molecules of glucose = 180 × 3.6 = **648 g**

(Molecular mass of glucose = 12 × 6 + 1 × 12 + 16 × 6 = 180

33. (*i*)Atomic mass of iron = 55.8

55.8 g of iron ≡ 1 gram atom

669.6 g of iron $= \frac{1}{55.8} \times 669.6 =$ **12 gram atom.**

(*ii*) Molecular mass of C_2H_5OH

= 2 × 12 + 5 × 1 + 1 × 16 + 1 × 1 = 46

46 g of C_2H_5OH = 1 gram molecule

73.6 g of C_2H_5OH $= \frac{1}{46} \times 73.6$

= **1.6 gram molecule.**

34. (*a*) 2.6 gram atoms of sulphur = 2.6 × 32 = 83.2 g

(*b*) 2.6 gram molecule of sucrose = 2.6 × 342 = 889.2g

(*c*) 2.60 g of iodine

∴ **Maximum mass is of 2.6 gram molecule of sucrose (*b*)**

35. (*i*) 6..022 × 10^{23} atoms of calcium have mass = 40 g

1 atom of calcium has mass $= \frac{40}{6.022 \times 10^{23}}$

$= \mathbf{6.64 \times 10^{-23}\ g.}$

(*ii*) Mass of 6.022 × 10^{23} molecules of SO_2 = 64 g

Mass of 1 molecule of SO_2 $= \frac{64}{6.022 \times 10^{23}}$

$= \mathbf{1.06 \times 10^{-22}\ g.}$

36. (*i*) 1 mole atoms of carbon = 6.022 × 10^{23} atoms

0.5 mole atoms of carbon = 6.022 × 10^{23} × 0.5

= **3.01 × 10^{23} atoms.**

(*ii*) 32 g of sulphur = 6.022 × 10^{23} atoms

Solution File

3.2 g of sulphur = $\frac{6.022\times10^{23}\times3.2}{32}$

= **6.022×10^{22} atoms.**

(*iii*) 180 g of glucose = 6.022×10^{23} molecules

18 g of glucose = $\frac{6.022\times10^{23}\times18}{180}$

= 6.022×10^{22} molecules

1 molecule of glucose = 6 + 12 + 6 = 24 atoms

6.022×10^{22} molecules of glucose

= $6.022 \times 10^{22} \times 24$ =**1.445×10^{24} atoms.**

(*iv*) 1 mole molecules of oxygen = 6.022×10^{23} molecules

0.20 mol molecules of oxygen = $6.022 \times 10^{23} \times 0.2$

= 1.2046×10^{23} molecules

1 molecule of oxygen = 2 atoms

1.2046×10^{23} molecules = $1.2046 \times 10^{23} \times 2$

= **2.409×10^{23} atoms.**

37. 40 g of calcium = 6.022×10^{23} atoms

15 g of calcium contain atoms

= $\frac{6.022\times10^{23}}{40} \times 15 = 2.258 \times 10^{23}$ atoms.

Now,number of sodium atoms = 2.258×10^{23}

6.022×10^{23} atoms of sodium has mass = 23 g

2.258×10^{23} atoms of sodium has mass

= $\frac{23}{6.022\times10^{23}} \times 2.258 \times 10^{23}$ = **8.624 g.**

38. (*i*) 28 g of nitrogen at N.T.P occupy volume = 22.4 L

1.4 g of nitrogen at N.T.P will occupy volume

= $\frac{22.4}{28} \times 1.4$ = **1.12 L.**

(*ii*) 6.022×10^{23} molecules of O_2 at N.T.P occupy volume = 22.4 L

6.022×10^{21} molecules of O_2 at N.T.P will occupy

volume = $\frac{22.4\times6.022\times10^{21}}{6.022\times10^{23}}$

= **0.224 L**

(*iii*) 1 mol of NH_3 occupy volume = 22.4 L

0.2 mol of NH_3 occupy volume = 22.4 × 0.2

= **4.48 L.**

39. Avogadro number of rupees = 6.022×10^{23} rupees

Rupees spent per second = 10^6

Rupees spent per year = $10^6 \times 60 \times 60 \times 24 \times 365$

∴ $10^6 \times 60 \times 60 \times 24 \times 365$ rupees are spent in = 1year

6.022×10^{23} rupees are spent in

= $\frac{1}{10^6\times60\times60\times24\times365} \times 6.022 \times 10^{23}$

=1.91×10^{10} years.

40. 1 gram atom corresponds to mass of 6.022×10^{23} atoms

Mass of 6.022×10^{23} atoms = $6.644 \times 10^{-23} \times 6.022 \times 10^{23}$

= 40.01 g

∴ 1 gram atom = 40.01 g

40.01 g of X = 1 gram atom

80×10^3 g of X = $\frac{1}{40.01} \times 80 \times 10^3$

= **1999.5 gram atoms.**

41. 22.4 L of ozone (O_3) at N.T.P = 6.022×10^{23} molecules

5.60 L of ozone at N.T.P = $\frac{6.022\times10^{23}}{22.4} \times 5.6$

= **1.506×10^{23} molecules.**

Now,1 molecule of ozone = 3 atoms

1.506×10^{23} molecules of ozone = $1.506\times10^{23}\times3$

= **4.518×10^{23} atoms.**

42. Gold present in 300 mg gold ring = $\frac{300\times20}{24}$

= 250 mg

Now, number of atoms present in 197 g

= 6.022×10^{23}

No. of atoms present in 250×10^{-3} g

= $\frac{6.022\times10^{23}\times250\times10^{-3}}{197}$

= **7.64×10^{20}.**

43. Molecular mass of KCl = 39 + 35.5 = 74.5 g

74.5 g of KCl = 1 mole atoms of K

1000 g of KCl = $\frac{1\times1000}{74.5}$ = 13.42 mole atoms of K

Now, molecular mass of K_2O = 39 × 2 + 16 = 94 g

2 mol atoms of K in K_2O weigh = 94 g

13.42 mol atoms of K in K_2O weigh = $\frac{94}{2} \times 13.42$

= 631 g = **0.631 kg.**

44. Gram molecular mass of $H_2C_2O_4 . 2H_2O$ = 126 g

No. of water molecules in 1 mol of oxalic acid = 2 mol

No. of water molecules in 126 g of oxalic acid = $2 \times 6.022 \times 10^{23}$

∴ No. of water molecules in 630×10^{-3} g of oxalic acid

= $\frac{2\times6.022\times10^{23}\times630\times10^{-3}}{126}$

= **6.022×10^{21}.**

45. Volume of CO_2 present in air = $1000 \times \frac{3}{100}$

= 0.03 mL

22400 ml of CO_2 at S.T.P. contain = 6.022×10^{23} molecules

0.3 mL of CO_2 at S.T.P. contain = $\frac{6.022\times10^{23}\times0.3}{22400}$

= **8.07×10^{18} molecules.**

46. Mass of dot = 10^{-6}

12 g of C contain atoms = 6.022×10^{23}

10^{-6} g of C contain atoms = $\frac{6.022\times10^{23}}{12} \times 10^{-6}$

= **5.019×10^{16} atoms.**

Solution File

47. 1 mol of CCl_4 = 154 g

154 g of CCl_4 contains $6.022 \times 10^{23} \times 4$ Cl atoms

or $6.022 \times 10^{23} \times 4$ Cl atoms are present in 154 g of CCl_4

$$1 \times 10^{25} \text{ Cl atoms are present in } = \frac{154 \times 1 \times 10^{25}}{6.022 \times 10^{23} \times 4} = 639.2 \text{ g}$$

$$\text{Volume of } CCl_4 = \frac{\text{Mass}}{\text{Density}} = \frac{639.2}{1.5} = 426 \text{ cc}$$

or $= \mathbf{0.426}$ **litre.**

48. *Number of C atoms in 1.0 g of C-14 isotope*

14 g of C-14 isotope contains C atoms = 6.022×10^{23}

1 g C-14 isotope contain C atoms

$$= \frac{6.022 \times 10^{23}}{14} \times 1 = 4.30 \times 10^{22}$$

Number of C atoms in 1.0 of C-12 isotope

12 g of C-12 isotope contains C atoms $= 6.022 \times 10^{23}$

1g of C-12 isotope contain C atoms

$$= \frac{6.022 \times 10^{23}}{12} \times 1 = 5.02 \times 10^{22}$$

Difference in number of C atoms

$$= 5.02 \times 10^{22} - 4.30 \times 10^{22} = \mathbf{7.2 \times 10^{21}}.$$

49. Molecular weight of $Na_2CO_3.10H_2O$ = 286

Mass of oxygen present in $Na_2CO_3.10H_2O$ = 208

Now 1 mol of $Na_2CO_3.10H_2O$ = 286 g

0.1 mol of $Na_2CO_3.10H_2O$ = 28.6 g

Wt. of oxygen present in 286 g of Na_2CO_3 = 208

Wt. of oxygen present in 28.6 g of $Na_2CO_3.10H_2O$

$$= \frac{208}{286} \times 28.6 = \mathbf{20.8 \text{ g.}}$$

50. (*i*) Molecular mass of $MgSO_4$

$= 24 + 32 + 4 \times 16 = 120$

Mass of Mg = 24

$$\text{Mass percentage of Mg} = \frac{24}{120} \times 100 = \mathbf{20\%}$$

Mass of S = 32

$$\text{Mass percentage of S} = \frac{32}{120} \times 100 = \mathbf{26.67\%}$$

Mass of O = $4 \times 16 = 64$

$$\text{Mass percentage of O} = \frac{64}{120} \times 100 = \mathbf{53.33\%}.$$

(*ii*) Molar mass of $(NH_4)_2Cr_2O_7 = 2 \times (14 + 4) + 2 \times 52 + 7 \times 16 = 252$

No. of parts of cation $NH_4^+ = 2 \times (14 + 4) = 36$

$$\text{Percentage of cation} = \frac{36}{252} \times 100 = \mathbf{14.29\ \%}$$

51. (*i*) **Calculation of empirical formula**

Element	Percentage	Atomic mass	Moles of atoms	Atomic ratio	Simplest whole no. ratio
C	40.68	12	$\frac{40.68}{12} = 3.39$	$\frac{3.39}{3.39} = 1$	2
H	5.08	1	$\frac{5.08}{1} = 5.08$	$\frac{5.08}{3.39} = 1.50$	3
O	100 – (40.68 + 5.08) = 54.24	16	$\frac{54.24}{16} = 3.39$	$\frac{3.39}{3.39} = 1$	2

Empirical formula = $C_2H_3O_2$

Empirical formula mass = $2 \times 12 + 3 \times 1 + 2 \times 16 = 59$

(*ii*) **Calculation of molecular formula**

Molecular mass = 2 × V.D. = 2 × 59 = 118

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{118}{59} = 2$$

Molecular formula = $2(C_2H_3O_2) = C_4H_6O_4$.

Solution File

52. (*i*) **Calculation of empirical formula**

Element	Percentage	Atomic mass	Mole ratio	Atomic ratio
C	54.2	12	$\frac{54.2}{12} = 4.50$	$\frac{4.50}{2.29} = 2$
H	9.2	1	$\frac{9.2}{1} = 9.2$	$\frac{9.2}{2.29} = 4$
O	36.6	16	$\frac{36.6}{16} = 2.29$	$\frac{2.29}{2.29} = 1$

Empirical formula = $\mathbf{C_2H_4O}$

(*ii*) **Calculation of molecular formula**

Empirical formula mass = $2 \times 12 + 4 \times 1 + 1 \times 16 = 44$

Molecular mass = 88

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{88}{44} = 2$$

∴ Molecular formula = $2(C_2H_4O) = \mathbf{C_4H_8O_2}$.

53. (*i*) **Calculation of empirical formula**

Element	Percentage	Moles of atoms	Atomic ratio	Simplest ratio
N	30.43	14	$\frac{30.43}{14} = 2.17$	$\frac{2.17}{2.17} = 1$
O	69.57	16	$\frac{69.57}{16} = 4.35$	$\frac{4.35}{2.17} = 2$

Empirical formula = $\mathbf{NO_2}$

(*ii*) **Calculation of molecular formula**

Empirical formula mass = $14 + 2 \times 16 = 46$

Molecular mass = 92

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{92}{46} = 2$$

Molecular formula = $2\,(NO_2) = \mathbf{N_2O_4}$.

54. (*i*) **Calculation of empirical formula**

Element	Percentage	Atomic mass	Moles of atoms	Simplest ratio
Na	36.5	23	$\frac{36.5}{23} = 1.59$	$\frac{1.59}{0.79} = 2$
H	0.8	1	$\frac{0.8}{1} = 0.8$	$\frac{0.8}{0.79} = 1$
P	24.6	31	$\frac{24.6}{31} = 0.79$	$\frac{0.79}{0.79} = 1$
O	38.1	16	$\frac{38.1}{16} = 2.38$	$\frac{2.38}{0.79} = 3$

Empirical formula = $\mathbf{Na_2HPO_3}$

(*ii*) **Calculation of molecular formula**

Empirical formula mass = $2 \times 23 + 1 \times 1 + 1 \times 31 + 3 \times 16 = 126$

Molecular mass = 126

Solution File

$$n = \frac{126}{126} = 1$$

Molecular formula = Na_2HPO_3 **Name : Sodium hydrogen phosphite.**

55. (*i*) **Calculation of empirical formula**

Element	Percentage	Atomic mass	Atomic ratio	Simple atomic ratio
Mg	20.0	24	$\frac{20.0}{24} = 0.83$	$\frac{0.83}{0.83} = 1$
S	26.66	32	$\frac{26.66}{32} = 0.83$	$\frac{0.83}{0.83} = 1$
O	53.33	16	$\frac{53.33}{16} = 3.33$	$\frac{3.33}{0.83} = 4$

Empirical formula = $\mathbf{MgSO_4}$

(*ii*) **Calculation of molecular formula**

Empirical formula mass = 24 + 32 + 4 x 16 = 120

Molecular mass = 120

$$n = \frac{120}{120} = 1$$

Molecular formula of anhydrous salt = $MgSO_4$

Let molecular formula mass = 100

Loss of weight due to dehydration = 51.2%

Molecular mass of anhydrous salt = 100 – 51.2 = 48.8

Now, if molecular mass of anhydrous salt is 48.8, then that of hydrated salt is = 100

If molecular mass of anhydrous salt is 120, then that of hydrated salt is = $\frac{100}{48.8} \times 120 = 245.9$

Loss in weight due to dehydration = 245.9 – 120 = 125.9

No. of water molecules in hydrated sample = $\frac{125.9}{18} = 7$

Molecular formula of hydrated salt (crystalline salt) = $\mathbf{MgSO_4 . 7H_2O}$.

56.

Compound	Percentage	Molecular formula mass	Relative no.	Simple ratio	Simplest whole no. ratio
CuO	44.82	79.5	$\frac{44.82}{79.5} = 0.564$	$\frac{0.564}{0.564} = 1$	1
SiO_2	34.83	60	$\frac{34.83}{60} = 0.580$	$\frac{0.580}{0.564} = 1.03$	1
H_2O	20.35	18	$\frac{20.35}{18} = 1.12$	$\frac{1.12}{0.564} = 2.0$	2

Empirical formula = $\mathbf{CuO . SiO_2 . 2H_2O}$.

Solution File

57.

Element / Compound	Percentage	At. mass/ Molecular mass	Relative no.	Simple ratio	Simplest whole no. ratio
Fe	20	56	$\frac{20}{56} = 0.36$	$\frac{0.36}{0.36} = 1$	1
S	11.5	32	$\frac{11.5}{32} = 0.36$	$\frac{0.36}{0.36} = 1$	1
O	23.1	16	$\frac{23.1}{16} = 1.44$	$\frac{1.44}{0.36} = 4$	4
H_2O	45.4	18	$\frac{45.4}{18} = 2.52$	$\frac{2.52}{0.36} = 7$	7

Empirical formula = $\mathbf{FeSO_4 . 7H_2O}$.

60. $\underset{\substack{2\times56+3\times16\\=160\\1\text{ kg}}}{Fe_2O_3} + 3H_2 \longrightarrow \underset{\substack{2\times56\\=112\\?}}{2Fe} + 3H_2O$

160 g of Fe_2O_3 on reduction gives iron = 112 g

1000 g of Fe_2O_3 on reduction will give iron

$= \frac{112}{160} \times 1000 = 700$ g

$=$ **0.7 kg.**

61. $\underset{\substack{342g\\34\times24\text{ g}}}{C_{12}H_{22}O_{11}} + \underset{\substack{12\times32g\\?}}{12O_2} \longrightarrow 12CO_2 + 11H_2O$

Energy requirement in a day = 34 × 24 g of sucrose

Amount of oxygen needed for burning 342 g of sucrose = 12 × 32 g

Amount of oxygen needed for burning 34 × 24 g of sucrose $= \frac{12\times32}{342} \times 34 \times 24$

$=$ **916.2 g.**

62. $\underset{100\text{ g}}{CaCO_3} + 2HCl \longrightarrow CaCl_2 + H_2O + \underset{\substack{22.4\text{ L at}\\\text{N.T.P.}}}{CO_2}$

22.4 L of CO_2 at N.T.P. is obtained from $CaCO_3$

= 100 g

100 L of CO_2 at N.T.P. will be obtained from $CaCO_3$

$= \frac{100}{22.4}\times100 = 446.4$ g

Impure marble required $= \frac{446.4\times100}{96.5}$ **= 462.6 g.**

63. $\underset{\substack{22.4\text{ L}\\5.6\text{ L}}}{CH_4} + 2O_2 \longrightarrow \underset{\substack{1\text{mol}\\?}}{CO_2} + 2H_2O$

22.4 L of methane on ignition give = 1 mol of CO_2

5.6 L of methane on ignition will give CO_2

$= \frac{1}{22.4}\times5.6$ **= 0.25 mol.**

64. Let us first calculate the theoretical yield of BI_3

$\underset{\substack{23+10.8+4\times1\\=37.8\text{ g}}}{NaBH_4} + 4I_2 \longrightarrow \underset{\substack{10.8+3\times127\\=391.8\text{ g}}}{BI_3} + NaI + 4HI$

37.8 g of $NaBH_4$ give BI_3 = 391.8 g

64 g of $NaBH_4$ will give BI_3 $= \frac{391.8}{37.8}\times64 = 663.36$ g

Experimental yield = 15.0 g

Percentage yield $= \frac{15.0\times100}{663.36}$

$=$ **2.27%.**

65. $\underset{\substack{5(39+14+2\times16)\\=425\text{ g}}}{5KNO_2} + \underset{\substack{2(39+54.9\times4\times16)\\=315.8\text{ g}}}{2KMnO_4} + 3H_2SO_4 \longrightarrow 5KNO_3 + 2MnSO_4 + K_2SO_4 + 3H_2O$

For oxidising 425 g of KNO_2, $KMnO_4$ needed

= 315.8 g

For oxidising 11.4 g of KNO_2, $KMnO_4$ will be needed

$= \frac{315.8}{425}\times11.4$

$=$ **8.47 g.**

66. $\underset{\substack{2(39+54.9+4\times16)\\=315.8\text{ g}}}{2KMnO_4} + 8H_2SO_4 + \underset{\substack{10(39.5+127)\\=1665\text{ g}}}{10KI} \longrightarrow 6K_2SO_4 + 2MnSO_4 + 5I_2 + 8H_2O$

1665 g of KI require for conversion to I_2 = 315.8 g $KMnO_4$

75 g of KI will require for conversion to I_2

$= \frac{315.8\times75}{1665}$ **= 14.22 g**

158 g of $KMnO_4$ is present in = 1000 mL

14.23 g of $KMnO_4$ will be present in

$= \frac{1000}{158}\times14.22$ **= 90.05 mL.**

67. Mass of NaOH = 0.38 g

Moles of NaOH $= \frac{0.38}{40}$

Volume of solution = 50.0 mL

Molarity $= \frac{0.38/40}{50.0} \times 1000$

$=$ **0.19 M.**

Solution File

68. 3 molal solution of NaOH means that 3 mol of NaOH are present in 1000 g of solvent.

$$\text{Mass of 3 mol of NaOH} = 3 \times 40 = 120 \text{ g}$$
$$\text{Mass of solution} = 1000 + 120 = 1120 \text{ g}$$
$$\text{Volume of solution} = \frac{1120\,\text{g}}{1.110\,\text{g mL}^{-1}} = 1009.0 \text{ mL}$$
$$\text{Molarity} = \frac{3}{1009.0} \times 1000 = \mathbf{2.97\ M}$$

69. $M_1V_1 = M_2V_2$

$$2.4 \times 500 = 1.6 \times V_2$$

or $$V_2 = \frac{2.4 \times 500}{1.6} = 750 \text{ ml}$$

Volume of water to be added = 750 – 500 = **250 mL.**

70. 86% by weight of H_2SO_4 means that 86 g of H_2SO_4 is present in 100 g of solution

$$\text{Volume of 100 g solution} = \frac{100}{1.80} = 55.6 \text{ ml}$$
$$\text{Moles of } H_2SO_4 = \frac{86}{98}$$
$$\text{Molarity} = \frac{86/98}{55.6} \times 1000 = \mathbf{15.8\ M.}$$

71. Molecular mass of $(COOH)_2 . 2H_2O = 126$

$$\text{Moles of oxalic acid} = \frac{0.63}{126} = 0.005 \text{ mol}$$
$$\text{Molarity} = \frac{0.005}{500} \times 1000 = \mathbf{0.01\ M.}$$

72. $$\text{Molarity} = \frac{\text{Moles of NaOH}}{\text{Vol. of solution}} \times 1000$$

Molarity = 0.15 M, vol. of solution = 27 mL

$$\therefore \quad 0.15 = \frac{\text{Moles of NaOH}}{27} \times 1000$$
$$\therefore \text{ Moles of NaOH} = \frac{27 \times 0.15}{1000} = \mathbf{0.00405\ mol}.$$

73. $$\text{Moles of oxalic acid} = \frac{\text{Molarity} \times \text{Volume in mL}}{1000}$$
$$\text{Moles of oxalic acid} = \frac{0.01 \times 100}{1000} = 0.001$$

1 mole of oxalic acid = 6.022×10^{23} molecules

0.001 mole of oxalic acid = $6.022 \times 10^{23} \times 0.001$

$= \mathbf{6.022 \times 10^{20}}$.

74. $AgNO_3 + HCl \longrightarrow AgCl + HNO_3$

1000 mL of 0.068 M $AgNO_3$ contain $AgNO_3$ = 0.068 mol

25 mL of 0.068 $AgNO_3$ contain $AgNO_3$

$$= \frac{0.068}{1000} \times 25 = 0.0017 \text{ mol}$$

1 mol of $AgNO_3$ give 1 mol of AgCl and therefore, 0.0017 mol of $AgNO_3$ will form 0.0017 mol of AgCl,

Amount of AgCl obtained = 0.0017 × 143.5

= **0.244 g .**

75. $2KOH + H_2SO_4 \longrightarrow K_2SO_4 + 2H_2O$

1000 mL of 0.15 M H_2SO_4 contain = 0.15 mol

20 mL of 0.15M H_2SO_4 contain

$$= \frac{0.15}{1000} \times 20 = 0.003 \text{ mol}$$

According to equation, 2 mol of KOH react with 1 mol of H_2SO_4 so that 0.003 mol of H_2SO_4 will react with 2 × 0.003 = 0.006 mol of KOH

0.34 mol of 0.34 M KOH are present in = 1000 mL

0.006 mol of 0.34 M KOH are present in

$$= \frac{1000}{0.34} \times 0.006 = \mathbf{17.65\ mL.}$$

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

1000 mL of 2.0 mol L^{-1} contain NaOH = 2.0 mol

200 mL of 2.0 mol L^{-1} contain NaOH

$$= \frac{2.0}{1000} \times 200 = 0.4 \text{ mol}$$

1 mol of NaOH is neutralised by 1 mol of HCl

0.4 mol of NaOH will be neutralised by 0.4 mol of HCl

1.0 mol of 1.0 mol L^{-1} is present in 1000 mL

0.4 mol of 1.0 mol L^{-1} is present in $\frac{1000}{1.0} \times 0.4$

= **400 mL**

1 mol of NaOH on reaction with HCl give = 58.5 g NaCl

0.4 mol of NaOH on reaction with HCl will give

$$= \frac{58.5}{1} \times 0.4 = \mathbf{23.4\ g.}$$

77. Molecular mass of NaOH = 23 + 16 + 1 = 40

1 M solution of NaOH contains 40 g L^{-1}

0.150 M solution of NaOH contains 4.0 × 0.150

= 6 g L^{-1}

6 g of NaOH is present in 1000 mL

0.184 g of NaOH is present in

$$= \frac{1000 \times 0.184}{6} = \mathbf{30.7\ mL\ of\ NaOH.}$$

78. $BaCl_2 (aq) + Na_2SO_4 (aq) \longrightarrow BaSO_4 (s) + 2NaCl$

1000 mL of 0.250 M Na_2SO_4 contain Na_2SO_4
$= 0.250$ mol

500 mL of 0.250 M Na_2SO_4 cantain Na_2SO_4

$$= \frac{0.250}{1000} \times 500 = 0.125 \text{ mol}$$

$$\text{Moles of } BaCl_2 = \frac{15}{208} = 0.0721 \text{ mol}$$

∴ $BaCl_2$ is limiting reactant.

Molecular mass of $BaSO_4 = 233$

1 mol of $BaCl_2$ react with Na_2SO_4 to give
$= 233$ g $BaSO_4$

0.072 mol of $BaCl_2$ will react with Na_2SO_4 to give
$= 233 \times 0.072 =$ **16.78 g.**

79. The balanced chemical equation is :

$$CaCO_3 (s) + 2\,HCl\,(aq) \longrightarrow CaCl_2\,(aq) + CO_2\,(g) + H_2O\,(l)$$

Let us first calculate the moles of HCl in 25 mL of 0.75 M HCl

1000 mL of 0.75 M HCl contain HCl = 0.75 mol

$$25 \text{ mL of } 0.75 \text{ M HCl contain HCl} = \frac{0.75}{1000} \times 25$$
$$= 0.01875 \text{ mol}$$

2 mol of HCl completely react with 1 mol of $CaCO_3$

0.01875 mol of HCl will completely react with

$$= \frac{1}{2} \times 0.01875$$
$$= 0.009375 \text{ mol}$$

Mass of 0.009375 mol of $CaCO_3 = 0.009375 \times 100$
= 0.9375 g

Chapter Summary

Key Terms & Laws

- **Precision** refers for the closeness of the set of values obtained from identical measurements of a quantity.
- **Accuracy** refers to the closeness of a single measurement to its true value.
- All numbers, small or large are expressed as a number between 1.000 .... and 9.999 multiplied or divided by 10, an appropriate number of times.
  $\Rightarrow 1426.2 = 1.4262 \times 10^3$
  In **scientific notation,** a number is generally expressed in the form $N \times 10^n$
  where N is a number (called digit term) between 1.000 .... and 9.999 and n is a number called an exponent.
- The **significant figures** in a number are all the certain digits plus one doubtful digit.
- In **addition and subtraction,** the final result should be reported to the same number of decimal places as the number with the minimum number of decimal places.
- In **multiplication and division,** the final result should be reported as having the same number of significant digits as the number with least number of significant digits.
- **Law of conservation of mass (Lavoisier in 1774).** During any physical or chemical change, the total mass of the products formed is equal to the total mass of reactants consumed.
- **Law of constant composition (Proust in 1798).** A chemical compound always contains same elements combined together in same proportion by mass.
- **Law of multiple proportions (John Dalton in 1803).** When two elements combine together to form two or more than two compounds then the masses of one of the elements that combine with fixed mass of the other bear a simple whole number ratio to one another.
- **Law of reciprocal proportions (Richter in 1792).** When two different elements combine separately with a fixed mass of a third element, then the ratio of their masses in which they do so is either the same or some whole number multiple of the ratio in which they combine with each other.
- **Gay Lussac's law (Gay Lussac in 1808).** When gases react with each other they do so in volumes which bear a simple whole number ratio to one another and to the volume of the products, if they are also gases, provided all volumes are measured under similar condition of temperature and pressure.
- **Avogadro's law.** Equal volume of all gases under similar conditions of temperature and pressure contain equal number of molecules.
- The smallest particle of an element that takes part in chemical reactions is **atom.**
- The smallest particle of a substance that has independent existence is **molecule.**
- One twelveth (1/12) of the mass of an atom of carbon (C-12) is **atomic mass unit.** It is equal to 1.66×10^{-27} kg.
- **Atomic mass** is the average relative mass of an atom of element as compared with mass of a carbon atom (C-12) taken as 12 a.m.u.
- **Molecular mass** is the average relative mass of a molecule of the substance as compared with mass of a carbon atom (C-12) taken as 12 a.m.u.
- Atomic mass expressed in grams is **gram atomic mass** and molecular mass expressed in grams is **gram molecular mass.**
- **One mole** is 6.022×10^{23} specified particles.

- Volume occupied by one mole of a gaseous substance is called gram molecular volume (G.M.V.). Its value is 22.4 L at N.T.P.
- **Mass percentage composition** of a compound gives the mass of each element expressed as percentage of the total mass.
- **Empirical formula** is the formula which gives the simplest whole number ratio of atoms of different elements present in the molecule of a compound.
  Molecular formula is the formula which gives the actual number of atoms present in the molecule of the compound. It is whole number multiple of empirical formula.
- **Molarity** is the number of moles of solute dissolved per litre of solution. It is expressed as mol dm^{-3} or M.
  Molarity changes with temperature because volume of the solution changes with temperature.
- **Limiting reactant** is the reactant which is completely consumed in the reaction.

QUICK CHAPTER ROUND UP

Matter
anything which occupies space and has mass.

Physical classification

- **solids.** definite shape and definite volume
- **liquids.** definite volume but no definite shape
- **gases.** neither definite shape nor definite volume

$$\text{solid} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{liquid} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{gas}$$

Chemical classification

Pure substances

- **Element:** simplest form of matter
- **Compound:** contain more than one kind of elements combined in a definite proportion

Mixtures

- **Homogeneous:** Uniform composition throughout
- **Heterogeneous:** composition is not uniform throughout

- The **significant figures** in a number are all the certain digits plus one doubtful digit.
- In **addition and subtraction**, the final result should be reported to the same number of decimal places as the number with the minimum number of decimal places.
- In **multiplication and division**, the final result should be reported as having the same number of significant digits as the number with least number of significant digits.

Smallest particle of element which may or may not have independent existence ⇒ **atom.**
Smallest particle of a substance (element or compound) which has independent existence ⇒ **molecule**

1 MOLE = 1.6022 × 10^{23} particles

1 MOLE → 6.022×10^{23} particles
1 MOLE → 22.4 L of a gas at N.T.P.
1 MOLE → 1 gram atom of an element
1 MOLE → 1 gram molecule of a substance

$$\text{Atomic mass} = \frac{\text{Mass of an atom}}{\frac{1}{12}\text{th mass of } ^{12}C}$$

$$\text{Molecular mass} = \frac{\text{Mass of a molecule}}{\frac{1}{12}\text{th mass of } ^{12}C}$$

- One twelveth (1/12) of the mass of an atom of carbon (C-12) is atomic mass unit. It is equal to 1.66×10^{-27} kg.

One mole of atoms = 6.022×10^{23} atoms
= Gram atomic mass of the element

$$\text{Moles of an element} = \frac{\text{Mass of element}}{\text{Atomic mass}}$$

$$\text{Mass of one atom} = \frac{\text{Atomic mass}}{6.022 \times 10^{23}}$$

One mole of molecules = 6.022×10^{23} molecules
= Gram molecular mass

$$\text{Moles of a compound} = \frac{\text{Mass of compound}}{\text{Molar mass}}$$

$$\text{Mass of one molecule} = \frac{\text{Molecular mass}}{6.022 \times 10^{23}}$$

Volume occupied by 1 mole of a gas at N.T.P. = 22.4 L

- Atomic mass expressed in grams is **gram atomic mass** or **gram atom** and molecular mass expressed in grams is **gram molecular mass** or **gram molecule**.

- **Mass percentage composition** of a compound gives the mass of each element expressed as percentage of the total mass.

$$\text{Mass \% of an element} = \frac{\text{Mass of element in 1 mole of compound}}{\text{Molar mass}} \times 100$$

- The formula which gives the simplest whole number ratio of atoms of different elements present in the molecule of a compound is **empirical formula**.
- The **molecular formula** gives the actual number of atoms present in the molecule of the compound. It is whole number multiple of empirical formula.
- Molecular formula = n (Empirical formula)

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

Molarity equation :

$$M_1V_1 = M_2V_2$$

$$\text{Molarity (M)} = \frac{\text{Moles of solute}}{\text{Vol. of solution (in mL)}} \times 1000$$

$$\text{Molality (m)} = \frac{\text{Moles of solute}}{\text{Mass of solvent (in g)}} \times 1000$$

$$\text{Normality (N)} = \frac{\text{Gram equiv. of solute}}{\text{Vol. of solution (in mL)}} \times 1000$$

$$\text{Normality} = \text{Molarity} \times \frac{\text{Mol. mass}}{\text{Eq. mass}}$$

$$\text{Mole fraction of solute}(x_B) = \frac{\text{Moles of solute}}{\text{Moles of solute + Moles of solvent}}$$

Mole fraction of solute + Mole fraction of solvent = 1

or Mole fraction of solvent=1 – Mole fraction of solute.

Limiting reactant ▸ the reactant which reacts completely in a reation.

NCERT FILE Solved

NCERT Textbook Exercises

Q. 1. Calculate the molecular mass of the following:

(*i*) H_2O (*ii*) CO_2 (*iii*) CH_4

Ans. (*i*) Molecular mass of H_2O

= 2 (1.008 u) + (16.00 u) = 18.016 u

(*ii*) Molecular mass of CO_2

= 1 (12.01 u) + 2(16.00 u) = 44.01 u

(*iii*) Molecular mass of CH_4

= 1 (12.01 u) + 4(1.008 u) = 16.042u

Q. 2. Calculate the mass percent of different elements present in sodium sulphate (Na_2SO_4).

Ans. Molar mass of Na_2SO_4

= 2 × At. mass of Na + At. mass of S + 4 × At. mass of O

= 2 × 23.0 + 32 + 4 × 16 = 142

$$\text{Mass \% of sodium} = \frac{2 \times 23}{142} \times 100 = 32.39\%$$

$$\text{Mass \% of sulphur} = \frac{32}{142} \times 100 = 22.53\%$$

$$\text{Mass \% of oxygen} = \frac{4 \times 16}{142} \times 100 = \mathbf{45.07\%}$$

Q. 3. Determine the empirical formula of an oxide of iron which has 69.9% iron and 30.1% dioxygen by mass. (Atomic masses : Fe = 55.85 amu, O = 16.00 amu).

Ans. *Refer Solved Example 63 (Page 53).*

Q. 4. Calculate the amount of carbon dioxide that would be produced when

(*i*) 1 mole of carbon is burnt in air

(*ii*) 1 mole of carbon is burnt in 16 g of dioxygen

(*iii*) 2 moles of carbon are burnt in 16 g of dioxygen

Ans. The balanced chemical equation of combustion of carbon in dioxygen (or air) is

$$\underset{1\text{ mol}}{C(s)} + \underset{1\text{ mol}}{O_2(g)} \longrightarrow \underset{1\text{ mol}}{CO_2(g)}$$

(*i*) In air, carbon will be completely burnt. 1 mol of carbon will give 1 mol of CO_2 or = 44g

(*ii*) Since only 16 g of dioxygen is available (0.5 mol), it will combine with only 0.5 mol of carbon. Dioxygen is the limiting reagent. Thus, 0.5 mol of carbon will be burnt to give 0.5 mol of CO_2 or = 22 g

(*iii*) In this case also, dioxygen is limiting reagent and only 0.5 mol of carbon will be burnt. It will produce 22 g of CO_2.

Q. 5. Calculate the mass of sodium acetate (CH_3COONa) required to make 500 mL of 0.375 molar aqueous solution. Molar mass of sodium acetate is 82.0245 g mol^{-1}.

Ans. 0.375 M aqueous solution means that 0.375 mol of sodium acetate are present in 1000 mL of solution. 500 mL of the solution should contain sodium acetate

$$= \frac{0.375}{2} \text{ mol}$$

Molar mass of sodium acetate = 82.0245 g mol^{-1}

$$\text{Mass of sodium acetate required} = \frac{0.375}{2} \times 82.0245$$

$$= \mathbf{15.38\ g}$$

Alternatively, it may be solved as :

$$\text{Molarity} = \frac{\dfrac{\text{Mass of sodium acetate}}{\text{Molar mass}}}{\text{Volume}} \times 1000$$

$$0.375 = \frac{\dfrac{\text{Mass of sodium acetate}}{82.0245}}{500} \times 1000$$

$\therefore$ Mass of sodium acetate $= \dfrac{0.375 \times 82.0245 \times 500}{1000}$

$= \mathbf{15.38\ g}$

Q. 6. Calculate the concentration of nitric acid in moles per litre in a sample which has a density, 1.41 g mL^{-1} and the mass per cent of nitric acid in it being 69%.

Ans. *Refer Solved Example 101 (Page 69).*

Q. 7. How much copper can be obtained from 100 g of copper sulphate ($CuSO_4$) ? (Atomic mass of Cu = 63.5 amu).

Ans. 1 mol of $CuSO_4$ contains 1 mol (1 gram atom) of Cu

Molar mass of $CuSO_4$ = 63.5 + 32 + 4 × 16
= 159.5 *g*

$\therefore$ 159.5 g of $CuSO_4$ will give Cu = 63.5 g

100 g of $CuSO_4$ will give Cu $= \dfrac{63.5}{159.5} \times 100$

$= \mathbf{39.81\ g}$

Q. 8. Determine the molecular formula of an oxide of iron in which the mass per cent of iron and oxygen are 69.9 and 30.1 respectively. Given that the molecular mass of iron oxide is 159.8 and atomic masses : Fe = 55.85 amu and O = 16.00 amu.

Ans. Calculation of empirical formula = Fe_2O_3 (See Solved Example 65)

Empirical formula mass of Fe_2O_3
= 2 × 55.85 + 3 × 16.00
= 111.7 + 48.00 = 159.7 g mol^{-1}

Molecular formula mass = 159.8 g mol^{-1}

$$n = \frac{\text{Molecular formula mass}}{\text{Empirical formula mass}}$$

$$= \frac{159.8}{159.7} = 1$$

$\therefore$ Molecular formula = $(Fe_2O_3) = Fe_2O_3$

Q. 9. Calculate the atomic mass (average) of chlorine using the following data :

	% Natural abundance	Molar mass
^{35}Cl	75.77	34.9689
^{37}Cl	24.23	36.9659

Ans. Average atomic mass of Cl

$$= \frac{75.77 \times 34.9689 + 24.23 \times 36.9659}{100}$$

$= \mathbf{35.453}$

Q. 10. In three moles of ethane (C_2H_6), calculate the following :

(*i*) Number of moles of carbon atoms
(*ii*) Number of moles of hydrogen atoms
(*iii*) Number of molecules of ethane

Ans. (*i*) 1 mole of C_2H_6 contains 2 moles of carbon

$\therefore$ Number of moles of carbon in 3 moles of C_2H_6 = 6

(*ii*) 1 mole of C_2H_6 contain 6 mole atoms of hydrogen

$\therefore$ Number of moles of hydrogen atoms in 3 moles of C_2H_6 = 3 × 6 = 18

(*iii*) 1 mole of C_2H_6 = 6.022 × 10^{23} molecules

$\therefore$ Number of molecules in 3 moles of C_2H_6 = 3 × 6.022 × 10^{23}

$= \mathbf{1.807 \times 10^{24}}$ **molecules**

Q. 11. What is the concentration of sugar ($C_{12}H_{22}O_{11}$) in mol L^{-1} if its 20g are dissolved in enough water to make a final volume upto 2 L ?

Ans. *Refer Solved Example 103 (Page 70).*

Q. 12. If the density of methanol is 0.793 kg L^{-1}, what is its volume needed for making 2.5 L of its 0.25 M solution ?

Ans. *Refer Solved Example 104 (Page 70).*

Q. 13. Pressure is defined as force per unit area of the surface. The SI unit of pressure, pascal is shown below :

$$\mathbf{1\ Pa = 1Nm^{-2}}$$

If mass of air at sea level is 1034 g cm^{-2}, calculate the pressure in pascal.

Ans. Pressure is force (i.e., weight) acting per unit area

But weight = *mg*

$\therefore$ Pressure = Weight per unit area

$$= \frac{1034\text{g}}{\text{cm}^2} \times 9.8\ \text{ms}^{-2}$$

$$= \frac{1034\text{ g}}{\text{cm}^2} \times 9.8\ \text{ms}^{-2} \times \frac{1\text{ kg}}{1000\text{ g}}$$

$$\times \frac{100\text{ cm} \times 100\text{ cm}}{1\text{ m} \times 1\text{ m}} \times \frac{1\text{N}}{\text{kg m s}^{-2}} \times \frac{1\text{Pa}}{1\text{Nm}^{-2}}$$

$= \mathbf{1.01332 \times 10^5\ Pa}$

Q. 14. What is the SI unit of mass ? How is it defined?

Ans. The SI unit of mass is kilogram (kg). Kilogram is defined as the mass of platinum-iridium (Pt-Ir) block, stored at the International Bureau of Weights and Measures in France. Thus it is the mass of the international prototype of the kilogram.

Q. 15. Match the following prefixes with their multiples :

	Prefixes	**Multiples**
(*i*)	**micro**	**10^6**
(*ii*)	**deca**	**10^9**
(*iii*)	**mega**	**10^{-6}**
(*iv*)	**giga**	**10^{-15}**
(*v*)	**femto**	**10**

Ans. (*i*) micro – 10^{-6} (*ii*) deca – 10 (*iii*) mega –10^6
(*iv*) giga – 10^9 (*v*) femto – 10^{-15}

Q. 16. What do you mean by significant figures ?

Ans. The significant figures in a number are all the certain digits plus one doubtful digit. It depends upon the precision of the scale or instrument used for the measurement. For example, if volume of a liquid is reported to be 18.25 mL, the digits 1, 8 and 2 are certain while 5 is doubtful. So, it has four significant figures (three certain plus one doubtful).

Q. 17. A sample of drinking water was found to be severely contaminated with chloroform, $CHCl_3$, supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass).

(*i*) Express this in per cent by mass.
(*ii*) Determine the molality of chloroform in the water sample.

Ans. (*i*) 15 ppm means that 15 parts of chloroform is present in 10^6 parts.

$\therefore$ % by mass of $CHCl_3 = \frac{15}{10^6} \times 100 = 1.5 \times 10^{-3}\%$

(*ii*) Molar mass of $CHCl_3 = 12 + 1 + 3 \times 35.5 = 119.5$

$$\text{Moles of } CHCl_3 = \frac{1.5 \times 10^{-3} g}{119.5 g\ mol^{-1}} = 1.255 \times 10^{-5}$$

Mass of water = 100 g

$$\therefore \quad \text{Molality} = \frac{1.225 \times 10^{-5}}{100} \times 1000$$

$$= \mathbf{1.255 \times 10^{-4}\ m}$$

Q. 18. Express the following in the scientific notation.
(*i*) 0.0048 (*ii*) 234,000 (*iii*) 8008
(*iv*) 500.0 (*v*) 6.0012

Ans. *Refer Solved Example 2 (Page 15).*

Q.19. How many significant figures are present in the following ?
(*i*) 0.0025 (*ii*) 208 (*iii*) 5005
(*iv*) 126,000 (*v*) 500.0 (*vi*) 2.0034

Ans. (*i*) 2 (*ii*) 3 (*iii*) 4
(*iv*) 3 (*v*) 4 (*vi*) 5

Q. 20. Round up the following upto three significant figures :
(*i*) 34.216 (*ii*) 10.4107 (*iii*) 0.04597 (*iv*) 2808

Ans. (*i*) 34.2 (*ii*) 10.4 (*iii*) 0.0460 (*iv*) 2810

Q. 21. The following data were obtained when dinitrogen and dioxygen react together to form different compounds :

	Mass of dinitrogen	Mass of dioxygen
(*i*)	14 g	16 g
(*ii*)	14 g	32 g
(*iii*)	28 g	32 g
(*iv*)	28 g	80 g

(*a*) Which law of chemical combination is obeyed by the above experimental data ? Give its statement.

(*b*) Fill in the blanks in the following conversions :

(*i*) 1 km = ... mm = ... pm
(*ii*) 1 mg = ... kg = ... ng
(*iii*) 1 mL = ... L = ... dm^3

Ans. (***a***) Fixing the mass of dinitrogen as 14 g, the masses of dioxygen combined will be

16 : 32 : 16 : 40

or 2 : 4 : 2 : 5

This is a simple whole number ratio and hence, the data illustrate the **law of multiple proportions**.

The law of multiple proportions states that when two elements combine to form two or more than two compounds, then the mass of one of the elements which combine with a fixed mass of the other bear a simple whole number ratio.

(***b***) (*i*)

$$1 \text{ km} = 1 \text{ km} \times \frac{1000 \text{ m}}{1 \text{ km}} \times \frac{100 \text{ cm}}{1 \text{ m}} \times \frac{10 \text{ mm}}{1 \text{ cm}} = 10^6 \text{mm}$$

$$1 \text{ km} = 1 \text{ km} \times \frac{1000 \text{ m}}{1 \text{ km}} \times \frac{1 \text{pm}}{10^{-12} \text{ m}} = 10^{15} \text{ pm}$$

Correct answer : 10^6, 10^{15}

(*ii*) $1 \text{mg} = 1 \text{mg} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 10^{-6} \text{ kg}$

$$1 \text{ mg} = 1 \text{ mg} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ ng}}{10^{-9} \text{ g}} = 10^6 \text{ ng}$$

Correct answer : 10^{-6}, 10^6

(*iii*) $1 \text{ mL} = 1 \text{mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 10^{-3} \text{ L}$

$$1 \text{ mL} = 1 \text{ cm}^3 \times \left(\frac{1 \text{ dm}}{10 \text{ cm}}\right)^3 = 10^{-3} \text{ dm}^3$$

Correct answer : 10^{-3}, 10^{-3}

Q. 22. If the speed of light is $3.0 \times 10^8\ ms^{-1}$, calculate the distance covered by light in 2.00 ns.

Ans. Distance covered by light = Speed × Time

$$= 3.0 \times 10^8 \text{ ms}^{-1} \times 2.00 \text{ ns} \times \left(\frac{10^{-9}}{1 \text{ ns}}\right)$$

$$= 6.00 \times 10^{-1} \text{ m} = \mathbf{0.600\ m}$$

Q. 23. In a reaction

$$A + B_2 \longrightarrow AB_2$$

identify the limiting reagent if any in the following reaction mixtures :

(*i*) 300 atoms of A + 200 molecules of B
(*ii*) 2 mol of A + 3 mol of B
(*iii*) 100 atoms of A + 100 molecules of B
(*iv*) 5 mol of A + 2.5 mol of B
(*v*) 2.5 mol of A + 5 mol of B

Ans. *Refer Solved Example 88 (Page 65).*

Q. 24. Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation :

$$N_2 (g) + 3H_2 (g) \longrightarrow 2NH_3 (g)$$

(*i*) Calculate the mass of ammonia produced if 2.00×10^3 g of dinitrogen reacts with 1.00×10^3g of dihydrogen.
(*ii*) Will any of the two reactants remain unreacted ?
(*iii*) If yes, which one and what would be its mass ?

Ans. *Refer Solved Example 92 (Page 66).*

Q. 25. How are 0.50 m Na_2CO_3 and 0.50 M Na_2CO_3 different ?

Ans. Molar mass of $Na_2CO_3 = 2 \times 23 + 12 + 3 \times 16$
$= 106 \text{ g mol}^{-1}$

Mass of 0.5 mol of $Na_2CO_3 = 106 \times 0.5 = 53$ g

0.50 m Na_2CO_3 solution means that 0.5 mol or 53g of Na_2CO_3 are present in 1000 g of solvent.

0.50 M Na_2CO_3 solution means that 0.5 mol or 53 g of Na_2CO_3 are present in 1 L of the solution.

Q. 26. If ten volumes of dihydrogen gas react with five volumes of dioxygen gas, how many volumes of water would be produced ?

Ans. Dihydrogen (H_2) reacts with dioxygen (O_2) as

$$2\,H_2\,(g) + O_2\,(g) \longrightarrow 2H_2O\,(g)$$

2 volumes of dihydrogen react with 1 volume of O_2 to produce 2 volumes of water vapour. Thus,

10 volumes of dihydrogen will react completely with 5 volumes of dioxygen to produce 10 volumes of water.

10 volumes of water would be produced.

Q. 27. Convert the following into basic units :
(*i*) 28.7 pm (*ii*) 15.15 μs (*iii*) 25365 mg

Ans. (*i*) $28.7\text{ pm} = 28.7\text{ pm} \times \dfrac{10^{-12}\text{m}}{1\text{pm}} = 2.87 \times 10^{-11}\text{ m}$

(*ii*) $15.15\ \mu\text{s} = 15.15\ \mu\text{s} \times \dfrac{10^{-6}\text{s}}{1\mu\text{s}} = 1.515 \times 10^{-5}\text{ s}$

(*iii*) $25365\text{ mg} = 25365\text{ mg} \times \dfrac{1\text{g}}{1000\text{ mg}} \times \dfrac{1\text{kg}}{1000\text{ g}}$

$= 2.5365 \times 10^{-2}\text{ kg}$

Q. 28. Which one of the following will have largest number of atoms ?
(*i*) 1 g Au (*s*) (*ii*) 1g Na(*s*) (*iii*) 1g Li (*s*)
(*iv*) 1g of Cl_2 (*g*)

Ans. No. of atoms can be calculated as :

(*i*) $1\text{g Au} = \dfrac{1}{197} \times 6.022 \times 10^{23} = \dfrac{6.022}{197} \times 10^{23}$

(*ii*) $1\text{g Na} = \dfrac{6.022}{23} \times 10^{23}$

(*iii*) $1\text{g Li} = \dfrac{6.022 \times 10^{23}}{7}$

(*iv*) $1\text{g } Cl_2 = \dfrac{2 \times 6.022 \times 10^{23}}{71} = \dfrac{6.022 \times 10^{23}}{35.5}$

It is clear that 1 g Li contains largest number of atoms.

Q. 29. Calculate the molarity of a solution of ethanol in water in which the mole fraction of ethanol is 0.040.

Ans. Mole fraction of ethanol

$$x(C_2H_5OH) = \frac{n(C_2H_5OH)}{n(C_2H_5OH) + n(H_2O)}$$

To calculate molarity, we need to calculate moles of ethanol in 1 L of solution or nearly 1L of water because the solution is dilute.

No. of moles of water in 1 L of water $= \dfrac{1000}{18}$

$= 55.55$ mole

$$\therefore \quad \frac{n(C_2H_5OH)}{n(C_2H_5OH) + 55.55} = 0.040$$

$$n\,(C_2H_5OH) = 0.04\,[n\,(C_2H_5OH)] + 2.222$$

or $\quad 0.96\,n\,(C_2H_5OH) = 2.222$

$$\therefore \quad n = \frac{2.222}{0.96} = 2.31$$

∴ 2.31 moles of ethanol are present in 1 L of solution and hence, molarity of the solution is **2.31 M**

Q. 30. What will be the mass of one ^{12}C atom in g ?

Ans. 1 mole of ^{12}C atoms $= 6.022 \times 10^{23}$ atoms $= 12$g

∴ 6.022×10^{23} atoms of ^{12}C have mass = 12g

1 atom of ^{12}C will have mass $= \dfrac{12}{6.022 \times 10^{23}}$

$= 1.9927 \times 10^{-23}$ g

Q. 31. How many significant figures should be present in the answer of the following calculations ?

(*i*) $\dfrac{0.02856 \times 298.15 \times 0.112}{0.5785}$

(*ii*) 5 × 5.364

(*iii*) 0.0125 + 0.7864 + 0.0215

Ans. *Refer Solved Example 8 (Page 17).*

Q. 32. Use the data given in the following table to calculate the molar mass of naturally occurring argon isotopes :

Isotope	**Isotopic molar mass**	**Abundance**
^{36}Ar	**35.96755 g mol^{-1}**	**0.337%**
^{38}Ar	**37.96272 g mol^{-1}**	**0.063 %**
^{40}Ar	**39.9624 g mol^{-1}**	**99.600 %**

Ans. *Refer Solved Example 26 (Page 36).*

Q. 33. Calculate the number of atoms in each of the following :
(*i*) 52 mol of Ar (*ii*) 52 u of He (*iii*) 52 g of He

Ans. *Refer Solved Example 41 (Page 42).*

Q. 34. A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g of carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (mesured at S.T.P.) of this welding gas is found to weigh 11.6 g. Calculate (*i*) empirical formula (*ii*) molar mass of the gas and (*iii*) molecular formula.

Ans. *Refer Solved Example 66 (Page 55).*

Q. 35. Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction :

$$CaCO_3 + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

What mass of $CaCO_3$ is required to react completely with 25 mL of 0.75 M HCl?

Ans. Let us calculate mass of HCl in 25 mL of 0.75 M HCl

1000 mL of 0.75 M HCl contain HCl = 0.75 × 36.5 g

∴ 25 mL of 0.75 M HCl will contain HCl

$$= \frac{0.75 \times 36.5}{1000} \times 25 = 0.684\text{ g}$$

$$\underset{100\text{g}}{CaCO_3} + \underset{2 \times 36.5\text{ g}}{2HCl} \longrightarrow CaCl_2 + CO_2 + H_2O$$

2 × 36.5 g of HCl react completely with $CaCO_3$ = 100 g

0.684 g of HCl will react completely with $CaCO_3$

$$= \frac{100}{2 \times 36.5} \times 0.684 = \mathbf{0.937\ g}$$

Q. 36. Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid (HCl) according to the reaction :

$$MnO_2(s) + 4HCl\ (aq) \longrightarrow MnCl_2\ (aq) + 2H_2O(l) + Cl_2(g)$$

How many grams of HCl react with 5.0 g of manganese dioxide ?

Ans. *Refer Solved Example 74 (Page 61).*

NCERT Exemplar Problems

Note: *Objective Questions from Exemplar Problems are given in Competition File, page 123*

Subjective Questions

Short Answer Questions *Carrying 2 or 3 marks*

Q. 1. What will be the mass of one atom of C-12 in grams ?

Ans. 1 mole of carbon atoms = 12g = 6.022×10^{23} atoms.

6.022×10^{23} atoms of carbon-12 have mass = 12 g

$$\therefore 1 \text{ atom of carbon has mass} = \frac{12}{6.022 \times 10^{23}} = 1.99 \times 10^{-23} \text{ g.}$$

Q. 2. How many significant figures should be present in the answer of the following calculations ?

$$\frac{2.5 \times 1.25 \times 3.5}{2.01}$$

Ans. Least precise number in the calculations is 2.5 or 3.5, which has 2 significant figures. Hence the answer should have **two** significant figures.

Q. 3. What is the symbol for SI unit of mole ? How is the mole defined ?

Ans. Symbol for SI unit of mole is mol.

It is defined as the amount of a substance that contains as many particles or entities as there are atoms in 12 g of ^{12}C isotope.

Q. 4. What is the difference between molality and molarity ?

Ans. Molality is the number of moles of solute present in one kilogram of solvent whereas molarity is the number of moles of solute dissolved in one litre of solution.

Molality is independent of temperature whereas molarity depends on temperature.

Q. 5. Calculate the mass percent of calcium, phosphorus and oxygen in calcium phosphate $Ca_3(PO_4)_2$.

Ans. Molecular mass of $Ca_3(PO_4)_2$

$$= 3 \times 40 + 2(31 + 4 \times 16)$$
$$= 310$$

$$\text{Mass of Ca} = 3 \times 40 = 120$$

$$\text{Mass percentage of Ca} = \frac{120}{310} \times 100 = 38.71\%$$

$$\text{Mass of P} = 2 \times 31 = 62$$

$$\text{Mass percentage of P} = \frac{62}{310} \times 100 = 20\%$$

$$\text{Mass of O} = 8 \times 16 = 128$$

$$\text{Mass percentage of O} = \frac{128}{310} \times 100 = 41.29\%$$

Q. 6. 45.4L of dinitrogen reacted with 22.7L of dioxygen and 45.4L of nitrous oxide was formed. The reaction is given below :

$$2N_2(g) + O_2(g) \longrightarrow 2N_2O(g)$$

Which law is being obeyed in this experiment ? Write the statement of the law.

Ans. Gases are reacting together to form gaseous products.

Ratio of volumes of gases :

$$N_2 : O_2 : N_2O = 45.4 : 22.7 : 45.4$$
$$= 2 : 1 : 2$$

which is a simple whole number ratio. Hence the experiment proves Gay Lussac's law of gaseous volumes. This law states that gases combine or are produced in a chemical reaction in a simple whole number ratio by volume provided that all gases are at the same temperature and pressure.

Q. 7. If two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in whole number ratio.

(*a*) Is this statement true ?

(*b*) If yes, according to which law ?

(*c*) Give one example related to this law.

Ans. (*a*) Yes, the given statement is true.

(*b*) According to the law of multiple proportions.

(*c*) Hydrogen and oxygen combine to form water (H_2O) and hydrogen peroxide (H_2O_2) as :

$$\underset{2\text{ g}}{H_2} + \underset{16\text{ g}}{\frac{1}{2}O_2} \longrightarrow \underset{\substack{\text{Water}\\18\text{ g}}}{H_2O}$$

$$\underset{2\text{ g}}{H_2} + \underset{32\text{ g}}{O_2} \longrightarrow \underset{\substack{\text{Hydrogen peroxide}\\34\text{ g}}}{H_2O_2}$$

Masses of oxygen, (*i.e.*, 16 g in H_2O and 32 g in H_2O_2) which combine with fixed mass of hydrogen (2 g) are in the ratio; 16 : 32 or 1 : 2. This is a simple whole number ratio.

Q. 8. Calculate the average atomic mass of hydrogen using the following data :

Isotope	% Natural abundance	Molar mass
^{1}H	99.985	1
^{2}H	0.015	2

Ans. Average atomic mass $= \dfrac{1 \times 99.985 + 2 \times 0.015}{99.985 + 0.015}$

$= 1.00015$ u.

Q. 9. Hydrogen gas is prepared in the laboratory by reacting dilute HCl with granulated zinc. Following reaction takes place.

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

Calculate the volume of hydrogen gas liberated at STP when 32.65 g of zinc reacts with HCl. 1 mol of a gas occupies 22.7 L volume at STP; Atomic mass of Zn = 65.3 u.

Ans. 1 mol of Zn *i.e.*, 63.5 g of Zn produces 22.7 L of H_2

$\therefore$ 32.65 g of Zn will liberate $H_2 = \dfrac{22.7}{63.5} \times 32.65$

$= \mathbf{11.35\ L}$.

Q. 10. The density of 3 molal solution of NaOH is 1.110 g mL^{-1}. Calculate the molarity of the solution.

Ans. 3 molal solution of NaOH means that 3 mol of NaOH are dissolved in 1000 g of solvent.

Mass of solute $= 3 \times 40 = 120$ g

$\therefore$ Mass of solution = Mass of Solvent + Mass of Solute

$= 1000$ g + 120 g

$= 1120$ g

Density of solution $= 1.110$ g mL^{-1}

Volume of solution $= \dfrac{1120}{1.110}$ mL

$= 1009.00$ mL

Molarity $= \dfrac{\text{Moles of solute}}{\text{Vol. of solution (in mL)}} \times 1000$

$= \dfrac{3}{1009} \times 1000 = \mathbf{2.97\ M}$.

Q. 11. Volume of a solution changes with change in temperature, then, will the molality of the solution be affected by temperature ? Give reason for your answer.

Ans. No, molality of solution does not change with temperature because it involves masses of solute and solvent which do not change with temperature.

Q. 12. If 4 g of NaOH dissolves in 36 g of H_2O, calculate the mole fraction of each component in the solution. Also, determine the molarity of solution (specific gravity of solution is 1 g mL^{-1}).

Ans. Mass of NaOH = 4 g

Number of moles of NaOH $= \dfrac{4\ g}{40\ g} = 0.1$ mol

Mass of H_2O = 36 g

Number of moles of $H_2O = \dfrac{36\ g}{18\ g} = 2$ mol

Mole fraction of NaOH $= \dfrac{0.1}{2 + 0.1} = \dfrac{0.1}{2.1} = 0.047$

Mole fraction of water $= \dfrac{2}{2 + 0.1} = \dfrac{2}{2.1} = 0.95$

Mass of solution = Mass of water + Mass of NaOH

= 36 g + 4 g = 40 g

Volume of solution $= 40 \times 1 = 40$ mL

(Since specific gravity of solution is 1 g mL^{-1})

Molarity $= \dfrac{\text{Moles of solute}}{\text{Volume of solution (in mL)}} \times 1000$

$= \dfrac{0.1}{40} \times 1000 = 2.5$ M.

Q. 13. The reactant which is entirely consumed in reaction is known as limiting reagent. In the reaction 2A + 4B → 3C + 4D, when 5 moles of A react with 6 moles of B, then

(*i*) which is the limiting reagent ?

(*ii*) calculate the amount of C formed ?

Ans. *Refer Conceptual Questions* $\boxed{3}$, *Q.9* (*Page 77*).

Long Answer Questions *carrying 5 marks*

Q. 14. A vessel contains 1.6 g of dioxygen at STP (273.15K, 1 atm pressure). The gas is now transferred to another vessel at constant temperature, where pressure becomes half of the original pressure. Calculate

(*i*) volume of the new vessel.

(*ii*) number of molecules of dioxygen.

Ans. (*i*) Moles of oxygen $= \dfrac{1.6}{32} = 0.05$ mol

1 mol of O_2 at STP has volume = 22.4 L

0.05 mol of O_2 at STP has volume $= 22.4 \times 0.05$

= 1.12 L

$V_1 = 1.12$ L $p_1 = 1$ atm

$V_2 = ?$ $p_2 = \dfrac{1}{2} = 0.5$ atm.

According to Boyle's law (unit 4)

$p_1V_1 = p_2V_2$

or $V_2 = \dfrac{p_1V_1}{p_2} = \dfrac{1\ \text{atm} \times 1.12\ \text{L}}{0.5\ \text{atm}} = 2.24$ L

(*ii*) No. of molecules in 1.6 g or 0.05 mol

$= 6.022 \times 10^{23} \times 0.05 = \mathbf{3.011 \times 10^{22}}$.

Q. 15. Calcium carbonate reacts wth aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction given below :

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

What mass of $CaCl_2$ will be formed when 250 mL of 0.76 M HCl reacts with 1000 g of $CaCO_3$? Name the limiting reagent. Calculate the

number of moles of $CaCl_2$ formed in the reaction.

Ans. *Refer Solved Example 124 (Page 75).*

Q. 16. Define the law of multiple proportions. Explain it with two examples. How does this law point to the existence of atoms ?

Ans. *Refer Text (Page 24).*

Q. 17. A box contains some identical red coloured balls, labelled as A, each weighing 2 grams. Another box contains identical blue coloured balls, labelled as B, each weighing 5 grams. Consider the combinations AB, AB_2, A_2B and A_2B_3 and show that law of multiple proportions is applicable.

Ans.

Combination of A and B	AB	AB_2	A_2B	A_2B_3
Mass of A (in g)	2	2	4	4
Mass of B (in g)	5	10	5	15

Masses of B which combine with fixed mass of A (say 4 g) are :

10 g , 20 g , 5 g , 15 g

Ratio 2 : 4 : 1 : 3

This is a simple whole number ratio. Hence the law of multiple proportions is applicable.

Check your Retention Power

QUICK MEMORY TEST

A. Say True or False

1. Properties of a compound are average of the properties of its constituent atoms.
2. There is no difference in writing mass of an object as 7.00 g or 7.0 g.
3. The number of ozone molecules present in 1 mole of ozone are 6.022×10^{23}.
4. There is no difference between mass and weight of a substance.
5. Balancing of chemical equations is in accordance with the law of conservation of mass.
6. Homogeneous mixtures have sharp melting and boiling points.
7. The figure 0.02450 has 4 significant figures.
8. There is no difference between 0.005 or 5.00×10^{-3} g.
9. The empirical and molecular formula of sucrose is same.
10. 0.5 mole of S_8 and 0.5 mol of P_4 have same number of polyatomic molecules.
11. Molarity of a solution changes with temperature but molality does not.
12. The empirical formula of glucose is CH_2O.
13. 1 gram atom of C and 1 gram atom of sulphur have same mass.
14. When 3.0 g of H_2 react with 29.0g of O_2 to form water, O_2 is the limiting reactant.
15. Nitrogen and oxygen combine to form N_2O, NO and NO_2. This is in accordance with law of reciprocal proportions.

B. Complete the missing links

1. AZT is used for the victims of ....
2. The prefix pico stands for ....
3. Pascal are the units for the physical quantity ....
4. The number of significant figures in 0.00030 is ....
5. Decimal equivalent of 2/3 is .... upto three significant figures.
6. The empirical formula of hydrogen peroxide is ....
7. The law which does not follow from Dalton's atomic theory is ....
8. The mass of a molecule of carbon-14 dioxide ($^{14}CO_2$) is .... g.
9. An atom of sulphur is .... times heavier than an atom of carbon.
10. The ratio of atoms of hydrogen in 1 mole of methane and 1 mole of sucrose ($C_{12}H_{22}O_{11}$) is ....
11. .... mol of N_2 are needed to produce 3.8 mol of NH_3 by reaction with hydrogen.
12. If mole fraction of sodium chloride in sodium chloride aqueous solution is 0.35, then mole fraction of water in the solution is ....
13. The molarity of 0.5 N H_2SO_4 solution is .... M.
14. Amount of glucose ($C_6H_{12}O_6$) required to prepare 100 mL of 0.1 M solution is ....
15. The empirical formula of benzene is ....

C. Choose the correct alternative

1. When 50g of N_2 and 10g of H_2 are mixed to produce NH_3; the limiting reagent is *N_2/H_2*.
2. Out of molarity and mole fraction; *molarity/mole fraction* is independent of temperature.
3. 1 gram mole of CO_2 has *less/more* mass than 1 gram mole of CO.
4. Empirical and molecular formula are *same/different* of acetylene.

5. Law of multiple proportion was stated by *Dalton/Proust*.
6. 1 femtometer is equal to $10^{-15}/10^{-18}$ m.
7. The units kg $m^{-1}s^{-2}$ are of *force/pressure*.
8. 0.0030 has *two/four* significant figures.
9. The sum 4.524 + 2.3 + 6.24 is equal to *13.1/13.06*.
10. Mass of 1*cc* of H_2 at STP is 9.0×10^{-5} *g/1g*.

Answers QUICK MEMORY TEST

A. Say True or False

1. **False :** Properties of a compound are entirely different from those of the constituent atoms.
2. **False :** because 7.00 g is more precise than 7.0 g.
3. True
4. **False :** Mass of a substance is a constant quantity but its weight varies from place to place.
5. True
6. **False :** Homogeneous as well as heterogeneous mixtures do not have sharp melting and boiling points.
7. True
8. **False :** 0.006 g contains one significant figure while 5.00×10^{-3} contains 3 significant figures.
9. True
10. True
11. True
12. True
13. False
14. False
15. False.

B. Complete the missing links

1. AIDS 2. 10^{-12} 3. Pressure
4. two 5. 0.667 6. HO
7. Gay-Lussac law of gaseous volumes
8. 7.64×10^{-23} 9. 8/3 10. 2 : 11 11. 1.9
12. 0.65 13. 0.25 14. 1.8 g 15. CH

C. Choose the correct alternative

1. H_2 2. mole fraction 3. more 4. different
5. Dalton 6. 10^{-15} 7. pressure 8. two
9. 13.1 10. 9.0×10^{-5} g.

HOTS Higher Order Thinking Skills & Advanced Level

QUESTIONS WITH ANSWERS

Q.1. In the combustion of methane in air, what is the limiting reactant and why ?

Ans. Methane is the limiting reactant because the other reactant is oxygen of the air which is always present in excess. Thus, the amounts of carbon dioxide and water formed will depend upon the amount of CH_4 burnt.

Q.2. What is kg-mole ? How many electrons are present in 1 kg mole of methane (CH_4)?

Ans. One kg-mole is the molecular mass of the substance expressed in kilograms. It is also called kilomole (k mol). One k mol contains 6.022×10^{26} particles. Thus,

1kg-mole of CH_4 contains 6.022×10^{26} molecules.

Since one molecule of methane contains 10 electrons and therefore, 1 kg mole of CH_4 contains $10 \times 6.022 \times 10^{26} = 6.022 \times 10^{27}$ electrons.

Q.3. Which aqueous solution has higher concentration : 1 molar or 1 molal solution of the same solute. Give reason.

Ans. 1 molar aqueous solution has higher concentration than 1 molal solution.

A molar solution contains one mole of solute in one litre of solution while a one molal solution contains one mole of solute in 1000 g of solvent.

If density of water is 1, then one mole of solute is present in 1000 mL of water in 1 molal solution while one mole of solute is present in less than 1000 mL of water in 1 molar solution (1000 mL solution = amount of solute + amount of solvent). Thus, 1 molar solution is more concentrated.

Q.4. Will the molarity of a solution at 50°C be same, less or more than molarity at 25°C?

Ans. Molarity at 50°C of a solution will be less than that at 25°C because molarity decreases with increase in temperature. This is because volume of the solution increases with increase in temperature but number of moles of solution remain the same.

Q.5. Is the law of constant composition true for all types of compounds ? Explain why or why not.

Ans. No, law of constant composition is not true for all types of compounds. It is true for only those compounds which are obtained from one isotope. For example, carbon exists in two common isotopes : ^{12}C and ^{14}C. When it forms $^{12}CO_2$, the ratio of masses is 12 : 32 or 3 : 8. However, when it is formed from ^{14}C i.e., $^{14}CO_2$, the ratio will be 14 : 32 *i.e.*, 7 : 16, which is not same as in the first case.

Q.6. Why is molality preferred over molarity in expressing the concentration of a solution ?

Ans. Molality is the number of moles of a solute present in 1000 g of the solvent while molarity is the number of moles of solute present in 1000 mL of the solution. Thus, molality involves only masses which donot change with temperature whereas molarity involves volume which changes with temperature and hence molality is preferred over molarity.

Q.7. What is the difference in expressing a weight of a solid as 36.5×10^3 g and 36.50×10^3 g ?

Ans. 36.5×10^3 g has three significant figures while 36.50×10^3 has four significant figures. Hence 36.50 represents greater accuracy than 36.5.

Q.8. How many significant figures are there in 'π' ?

Ans. Infinite number.

Q.9. In calculations involving more than one arithmetic operation, rounding off to the proper number of significant figures may be done once at the end if all the operations are multiplication and/ or division or if they are all additions and/or subtractions but not if they are combinations of additions or subtractions with multiplications or divisions. Explain.

Ans. There are different rules for the number of significant digits in the answer to an addition or subtraction and to a multiplication or division. Therefore, they must be applied separately when a mixed calculation is performed.

Q.10. Calculate the molarity of water if its density is 1000 kg/m^3.

Ans. Molarity of water means the number of moles of water in 1 litre of water.

1 L of water = 1000 cm^3 = 1000 g

$$(\because 1000 \text{ kg/m}^3 = 1 \text{g/cm}^3)$$

$$1000 \text{ g of water} = \frac{1000}{18}$$

$$= 55.56 \text{ moles}$$

$$\therefore \quad \text{Molarity} = \mathbf{55.56\ M.}$$

Q.11. Sulphuric acid is generally available in market as 18.0 M solution. How would you prepare 250 mL of 0.50 M aqueous H_2SO_4?

Ans. Applying molarity equation,

$$M_1V_1 = M_2V_2$$

Volume of 18 M (V_2) H_2SO_4 required to prepare 250 mL (V_1) of concentration 0.50 M (M_1).

$$0.50 \times 250 = 18 \times V_2$$

$$\therefore \quad V_2 = \frac{0.50 \times 250}{18} = 6.94 \text{ mL}$$

Volume of 18 M H_2SO_4 required = 6.94 mL

Volume of water required = 250 – 6.94 = **243.06 mL**

Q.12. A compound (molecular mass = 246) has the following data :

Element	% Composition	Relative no. of atoms
A	9.76	0.406
B	13.01	0.406
C	26.01	1.625
D	51.22	2.846

From the data find out

(*i*) atomic masses of the elements A, B, C and D,

(*ii*) simple ratio,

(*iii*) molecular formula of the compound.

Ans. Step I. *To calculate the atomic masses :*

The relative number of atoms

$$= \frac{\text{Percentage of element}}{\text{Atomic mass}}$$

$$\text{Atomic mass} = \frac{\text{Percentage of element}}{\text{Relative number of atoms}}$$

$$\therefore \quad \text{Atomic mass of A} = \frac{9.76}{0.406} = 24,$$

$$\text{Atomic mass of B} = \frac{13.01}{0.406} = 32$$

$$\therefore \quad \text{Atomic mass of C} = \frac{26.01}{1.625} = 16,$$

$$\text{Atomic mass of D} = \frac{51.22}{2.846} = 18$$

Step II. *To calculate the simple ratio of atoms.*

Element	Relative no. of atoms	Simple atomic ratio
A	0.406	$\frac{0.406}{0.406} = 1$
B	0.406	$\frac{0.406}{0.406} = 1$
C	1.625	$\frac{1.625}{0.406} = 4$
D	2.846	$\frac{2.846}{0.406} = 7$

Thus, the atomic ratio of A : B : C : D is 1: 1 : 4 : 7.

Hence, the empirical formula = $\mathbf{ABC_4D_7}$.

Step III. *Calculation of molecular formula*

Empirical formula mass = 24 + 32 + 4 × 16 + 7 × 18

= 246

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{246}{246} = 1$$

$$\text{Molecular formula of compound} = (ABC_4D_7)_1$$

$$= \mathbf{ABC_4D_7}$$

Q.13. A compound on analysis gave the following percentage composition : Sodium = 18.59%, Sulphur = 25.80%, Hydrogen = 4.03% and Oxygen = 51.58%. Calculate the molecular formula of crystalline salt on the assumption that all the hydrogen atoms in the compound are present in combination with oxygen as water of crystallisation. The molecular weight of the compound is 248 a.m.u.

Solution.

Step I. *Calculation of empirical formula*

Element	Percentage of element	At. mass	Moles of atoms	Mole ratio
Na	18.59	23	$\frac{18.59}{23} = 0.80$	$\frac{0.80}{0.80} = 1$
S	25.80	32	$\frac{25.80}{32} = 0.80$	$\frac{0.80}{0.80} = 1$
H	4.03	1	$\frac{4.03}{1} = 4.03$	$\frac{4.03}{0.80} = 5$
O	51.58	16	$\frac{51.58}{16} = 3.22$	$\frac{3.22}{0.80} = 4$

Thus, the simple ratio of Na : S : H : O is 1 : 1 : 5 : 4. Therefore, the empirical formula of the compound is $NaSH_5O_4$.

Step II. *Calculation of molecular formula*

Empirical formula mass $= 23 + 32 + 5 \times 1 + 4 \times 16 = 124$

Molecular mass $= 248$

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{248}{124} = 2$$

$\therefore$ Molecular formula $= (NaSH_5O_4)_2 = Na_2S_2H_{10}O_8$

Since all the 10 hydrogen atoms are present as water molecules (H_2O), the total number of water molecules is $5H_2O$.

Molecular formula = $\mathbf{Na_2S_2O_3 \cdot 5H_2O}$

Q.14. The vapour density of a mixture of NO_2 and N_2O_4 is 38.3 at 26.7°C. Calculate the number of moles of NO_2 in 100 g of the mixture.

Ans. V.D. of mixture of NO_2 and $N_2O_4 = 38.3$

Mol. wt. of mixture of NO_2 and $N_2O_4 = 38.3 \times 2 = 76.6$

(Mol. wt. = 2 × V.D.)

Let NO_2 present in 100 g of mixture $= x$

N_2O_4 present in 100 g of mixture $= 100 - x$

Mol. wt. of $NO_2 = 46$, Mol. wt. of $N_2O_4 = 92$

Now, $$\frac{x}{46} + \frac{100-x}{92} = \frac{100}{76.6}$$

Solving for x, we get $x = 20.1$ g

$\therefore$ Moles of NO_2 in the mixture $= \frac{20.1}{46}$

$= \mathbf{0.437}$

Q.15. A solid mixture (5.0 g) consisting of lead nitrate and sodium nitrate is heated below 600°C until the weight of the residue was constant. If the loss in weight is 28.0%, calculate the amount of lead nitrate and sodium nitrate in the mixture.

Ans. Let the amount of $NaNO_3$ in the mixture $= x$ g

Amount of $Pb(NO_3)_2$ in the mixture $= (5 - x)$ g

$$\underset{\substack{2(23+14+3\times16)\\ = 170\text{ g}}}{2NaNO_3} \longrightarrow 2NaNO_2 + \underset{\substack{2\times16\\ = 32}}{O_2\uparrow}$$

$$\underset{\substack{2(207+28+96)\\ = 662\text{ g}}}{2Pb(NO_3)_2} \longrightarrow 2PbO + \underset{\substack{2\times16\\ = 32}}{O_2\uparrow} + \underset{\substack{4\times(14+32)\\ = 184\text{ g}}}{4NO_2\uparrow}$$

170 g of $NaNO_3$ give $O_2 = 32$ g

x g of $NaNO_3$ give $O_2 = \frac{32}{170} \times x$ g

Similarly,

662 g of $Pb(NO_3)_2$ give O_2 and NO $= 216$ g

$(5 - x)$ g of $Pb(NO_3)_2$ give gases $= \frac{216}{662} \times (5 - x)$ g

Total loss on heating $= \frac{32x}{170} + \frac{216}{662}(5 - x)$

Actual loss on heating = 28% of 5 g $= \frac{28 \times 5}{100} = 1.4$ g

$\therefore \quad \frac{32x}{170} + \frac{216}{662}(5 - x) = 1.4$

Solving for x, we get $x = 1.676$ g

$\therefore$ Wt. of $NaNO_3$ = **1.676 g**

Wt. of $Pb(NO_3)_2 = 5 - 1.676$

= **3.324 g.**

Q.16. A sample of hard water contains 20 mg of Ca^{2+} ions per litre. How many milliequivalents of Na_2CO_3 would be required to soften 1 litre of sample ?

Ans. $$\underset{40\text{ g}}{Ca^{2+}} + \underset{106\text{ g}}{Na_2CO_3} \longrightarrow CaCO_3 + 2Na^+$$

40 g of Ca^{2+} react with $Na_2CO_3 = 106$ g

20×10^{-3} g of Ca^{2+} react with $Na_2CO_3 = \frac{106}{40} \times 20 \times 10^{-3}$

$= 5.3 \times 10^{-2}$ g

Eq. wt. of $Na_2CO_3 = \frac{\text{Mol. wt.}}{2} = \frac{106}{2} = 53$

Equivalents of $Na_2CO_3 = \frac{5.3 \times 10^{-2}}{53} = 1 \times 10^{-3}$ equiv.

Milliequivalents of $Na_2CO_3 = \frac{1 \times 10^{-3}}{10^{-3}}$

= **1 milliequivalent.**

Q.17. Igniting MnO_2 converts it quantitatively to Mn_3O_4. A sample of pyrolusite is of the following composition :

MnO_2 = 80%, SiO_2 and other inert contents = 15%, rest being water.

The sample is ignited in air to constant weight. What is the percentage of Mn in the ignited sample ? (At. wt. of Mn = 55).

Ans. $$\underset{\substack{3(55+2\times16)\\ = 261\text{ g}}}{3MnO_2} \longrightarrow \underset{\substack{3\times55+4\times16\\ = 229\text{ g}}}{Mn_3O_4} + O_2$$

Let the amount of pyrolusite ignited = 100 g

Wt. of MnO_2 = 80 g

Wt. of SiO_2 and other inert contents = 15 g

Wt. of water = 100 – (80 + 15) = 5 g

Now, 261 g of MnO_2 gives Mn_3O_4 = 229 g

80 g of MnO_2 give Mn_3O_4 $= \frac{229}{261} \times 80$

$= 70.19$ g.

During ignition, water present in pyrolusite is removed while SiO_2 and other inert contents remain as such.

Total wt. of residue = 70.19 + 15 = 85.19 g

$\therefore$ Percentage of Mn_3O_4 in the residue $= \frac{70.19}{85.19} \times 100$

$= 82.39\%$

Now 3 Mn = Mn_3O_4

3×55 $\quad$ $3 \times 55 + 4 \times 16$

$= 165$ $\quad$ $= 229$

$\therefore$ 229 g of Mn_3O_4 contain Mn = 165

82.39 g of Mn_3O_4 contain Mn $= \frac{165}{229} \times 82.39$

$= 59.36$ g

$\therefore$ Percentage of Mn in ignited sample = **59.36%**

Q.18. The density of gold is 19.3 g cm^{-3}. Calculate the diameter of a solid gold sphere having a mass of 422 g.

Ans. Volume of gold sphere $= \frac{\text{Mass}}{\text{Density}}$

$= \frac{422 \text{ g}}{19.3 \text{ g cm}^{-3}}$

$= 21.865 \text{ cm}^3$

Volume of sphere $= \frac{4}{3}\pi r^3$

or radius, $r = \left(\text{Volume} \times \frac{3}{4\pi}\right)^{1/3}$

$= \left(\frac{21.865 \times 3 \times 7}{4 \times 22}\right)^{1/3}$

$= (5.218)^{1/3} = 1.73$ cm

$\therefore$ Diameter $= 2 \times r$

$= 1.73 \times 2 =$ **3.46 cm**

Q.19. Lithium oxide is used to remove water from air according to the reaction :

$$\mathbf{Li_2O(s) + H_2O(g) \longrightarrow 2LiOH(s)}$$

If 72 kg of water is to be removed and 35 kg of Li_2O is available

(*i*) which reactant is limiting ?

(*ii*) how many kg of excess reactant is left ?

Ans. Moles of water $= \frac{72 \times 10^3}{18}$

$= 4 \times 10^3$ mol

Moles of Li_2O $= \frac{35 \times 10^3}{30}$

$= 1.167 \times 10^3$ mol

(*i*) Li_2O is limiting reagent because 1 mole of Li_2O reacts with 1 mole of water.

(*ii*) Moles of water which will react $= 1.167 \times 10^3$ mol

Moles of water left unreacted

$= 4 \times 10^3 - 1.167 \times 10^3$

$= 2.833 \times 10^3$

Mass of water left $= 2.833 \times 10^3 \times 18$

= 50.99 kg

Q.20. P_4O_6 and P_4O_{10} are formed by burning P_4 with O_2 as :

$$\mathbf{P_4 + 3O_2 \longrightarrow P_4O_6}$$

$$\mathbf{P_4 + 5O_2 \longrightarrow P_4O_{10}}$$

What are the masses of P_4O_6 and P_4O_{10} that will be produced by the combustion of 2.0 g of P_4 in 2.0 g of oxygen leaving no P_4 and O_2 ?

Ans. P_4O_6 and P_4O_{10} are formed as :

$$P_4 + 3O_2 \longrightarrow P_4O_6$$

$$P_4 + 5O_2 \longrightarrow P_4O_{10}$$

Let x be the mass of P_4 that is converted into P_4O_6 so that

Mass of P_4 which is converted to $P_4O_{10} = 2 - x$

Mass of oxygen required for forming P_4O_6

$= \frac{x}{4 \times 31} \times 96$

Mass of oxygen required for forming P_4O_{10}

$= \frac{2 - x}{4 \times 31} \times 160$

Total oxygen required,

$$\left(\frac{x}{4 \times 31} \times 96\right) + \left(\frac{2 - x}{4 \times 31} \times 160\right) = 2.0$$

or $\frac{96x}{124} + \frac{320}{124} - \frac{160x}{124} = 2.0$

or $-\frac{64x}{124} = 2.0 - \frac{320}{124}$

or $-\frac{64x}{124} = -\frac{72}{124}$

$x = \frac{72}{64} = 1.125$ g

Mass of P_4O_6 formed

$= \frac{\text{Mass of } P_4}{\text{Molar mass of } P_4} \times \text{Molar mass of } P_4O_6$

$$= \frac{1.125}{124} \times 220 = 1.996\,\text{g}$$

Mass of P_4O_{10} formed

$$= \frac{\text{Mass of } P_4}{\text{Molar mass of } P_4} \times \text{Molar mass of } P_4O_{10}$$

$$= \frac{0.875}{124} \times 284 = 2.004\,\text{g}$$

Revision Exercises

Very Short Answer Questions *carrying 1 mark*

1. Define Avogadro's law.
2. What is meant by a.m.u. ?
3. Define significant figure.
4. What is a mole ?
5. Define precision.
6. State the law of definite proportions.
7. Define atomic mass of an element.
8. Does a balanced chemical equation obey the law of conservation of mass ?
9. Define molarity of a solution.
10. Express decimal equivalent of 2/7 to three significant figures.
11. Is the molar volume of NH_3 different from that of CO_2 ?
12. Name a monoatomic gas ? What is its valency ?
13. Define limiting reagent.
14. Write 0.000623 cm in a scientific notation.
15. Define law of multiple proportions.
16. What is gram molecular mass ? Give one example.
17. Give one example each of a molecule in which empirical formula and molecular formula are

    (*i*) same and (*ii*) different.
18. Define mole in terms of number.
19. Balance the equation :

    $CaF_2 + H_2SO_4 + H_3BO_3 \longrightarrow CaSO_4 + BF_3 + H_2O$
20. How many atoms of carbon are present in 0.1 mole of $C_{12}H_{22}O_{11}$?
21. How many hydrogen atoms are present in 60 a.m.u. of ethane ?
22. What is meant by one gram of atom of iron ?
23. What is the S.I. unit of density ?
24. Name the law which deals with the ratios of the volumes of the gaseous reactants and products.
25. Which isotope of C is used for expressing relative atomic masses of elements ?
26. An atom of an element is 13 times heavier than the mass of a carbon atom. What is its mass in a.m.u. ?
27. What is the standard for the molecular weights of molecules ?
28. What is the ratio of molar volumes of SO_2 and SO_3 ?
29. State law of reciprocal proportions.
30. What volume will 250 g of mercury occupy ? (Density of mercury = 13.6 g cm^{-3}).

Short Answer Questions *Carrying 2 or 3 marks*

1. What do you understand by the terms element, compound and mixture ? Give two examples in each case.
2. Explain the term mole. What does one mole of ammonia represent ?
3. Give the SI units for (*i*) volume (*ii*) speed and (*iii*) force.
4. What do you understand by the terms (*i*) empirical formula and (*ii*) molecular formula ? How are they related to each other ? Illustrate with an example.
5. Define molarity. What does 1 M solution of sodium carbonate mean ?
6. Classify the following into elements, compounds or mixtures :

    (*i*) Water (*ii*) milk (*iii*) tea (*iv*) iron (*v*) sugar (*vi*) smoke (*vii*) sulphur (*viii*) 22 carat gold (*ix*) iodised table salt (*x*) gasoline.
7. What are homogeneous and heterogeneous mixtures ? Which of the following are homogeneous ?

    (*a*) tap water (*b*) wood (*c*) soil (*d*) smoke (*e*) cloud.
8. When two substances A and B are mixed together in a pestel and mortar, a large amount of heat is evolved and a new substance C is formed. C has the properties different from A and B. Is C an element, compound or a mixture?
9. State the following :

    (*i*) atomic mass (*ii*) gram atomic mass (*iii*) gram molar volume.
10. How would you recover

    (*i*) iodine from a mixture of iodine and salt?

    (*ii*) sulphur from a mixture of carbon and sulphur ?
11. State Avogadro's hypothesis. In what way, has it given support to Dalton atomic theory ?
12. How can you deduce the atomicity of hydrogen with the help of Avogadro's hypothesis ?
13. State the following laws of chemical combination and give one example in each case

    (*i*) Law of constant composition.

    (*ii*) Law of multiple proportions.

14. What do you understand by a balanced chemical equation ? What quantitative information does a balanced chemical equation convey ?

15. Explain (*i*) molarity (*ii*) limiting reagent.

16. Write the balanced chemical equations for the following reactions :

(*i*) Manganese dioxide and concentrated hydrochloric acid.

(*ii*) Sodium thiosulphate and iodine.

(*iii*) Copper and dilute nitric acid.

(*iv*) Sulphur dioxide and hydrogen sulphide.

17. Write the empirical formulae of the compounds having the following 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.

18. Balance the following equations :

(*i*) $H_3PO_3 \longrightarrow H_3PO_4 + PH_3$

(*ii*) $Ca + H_2O \longrightarrow Ca(OH)_2 + H_2$

(*iii*) $Fe_2(SO_4)_3 + NH_3 + H_2O \longrightarrow Fe(OH)_3 + (NH_4)_2SO_4$

19. What do you understand by the term formula mass ? How does it differ from molecular mass ?

20. Which of the following has (*i*) maximum (*ii*) minimum mass ?

(*a*) 1 gram atom of C

(*b*) 1 a.m.u. of an atom

(*c*) 1 gram mole of sulphur dioxide

(*d*) 6.02×10^{20} atoms of nitrogen.

Long Answer Questions *carrying 5 marks*

1. State the law of conservation of mass. How is it verified experimentally ?

2. What are laws of chemical combinations ? Discuss any three laws in detail.

3. Why is it necessary to balance a chemical equation ? Outline briefly the various steps for balancing a chemical equation by hit and trial method.

4. Write short notes on

(*i*) Limiting reagent (*ii*) Avogadro hypothesis

(*iii*) Dalton's atomic theory.

5. How is mole related to

(*i*) mass (*ii*) volume and (*iii*) number of molecules of a substance ?

6. Complete the following :

	CaH_2 +	$2H_2O$ $\longrightarrow$	$Ca(OH)_2$ +	$2H_2$
(*a*)	2 moles		....	....
(*b*)	.... g	........g	.........g	........2g.
(*c*)	..... H atoms	..... H atoms	6×10^{23} H atoms	..... H atoms
(*d*)	6×10^{20} total atoms	.....total atoms	.....total atoms	.....total atoms

7. What are the main postulates of Dalton's atomic theory ? What were its limitations ? How has the theory been modified ?

8. Define Avogadro number and mole. What is their importance ?

9. What are the essentials of a chemical equation ? What is the information conveyed by a chemical equation ?

10. Explain the following :

(*a*) Gay Lussac law

(*b*) Law of definite composition

(*c*) Empirical and molecular formula

(*d*) Relation between mole and volume of gases

(*e*) Limiting reagent.

Numerical Problems *carrying 5 marks*

1. State the number of significant figures in each of the following :

(*a*) 0.0037 (*b*) 0.00601

(*c*) 1.0001 (*d*) 0.00236

(*e*) 1.06×10^{-3}

2. Express the following numbers to three significant figures :

(*a*) 6.0263 (*b*) 2.3652

(*c*) sixty thousand (*d*) 2.861×10^5

3. Express the result of the following calculations to appropriate number of significant figures :

(*a*) $\dfrac{7.5 \times 206.8}{0.0512 \times 1002}$

(*b*) 4.20 + 1.6523 + 0.015

(*c*) (1.0042 – 0.0034) (1.23)

(*d*) $\dfrac{8.5 \times 208.9}{0.054 \times 9291.6}$

4. Carbon and oxygen combine to form two oxides having the composition : 1st oxide C = 42.9% and II oxide C = 27.3%. Show that the data is in agreement with the law of multiple proportions.

5. Calculate the amount in grams of :

(*a*) 2.5 gram atoms of nitrogen

(*b*) 3.6 gram mole of carbon dioxide

6. Calculate (*i*) number of molecules present in 2.24 dm^3 of carbon dioxide at N.T.P.

(*ii*) mass of an atom of oxygen

(*iii*) number of oxygen atoms in 2 mol of ozone

(*iv*) volume occupied by 4.4 g of SO_2 at N.T.P.

7. One atom of nickel weighs 9.75×10^{-23} g. Calculate the atomic mass of nickel.

8. How many molecules are present in 1 kg of hydrogen ?

9. Calculate the total charge of a mole of electrons if the electrical charge on a single electron is 1.60×10^{-19} C.

10. The volume of a drop of rain was found to be 0.448 ml at N.T.P. How many molecules of water and number of atoms of hydrogen are present in this drop ?

11. Calculate the number of molecules of oxygen in 150 ml of it at 20°C and 750 mm pressure.

12. Assuming the atomic mass of a metal M to be 56, calculate the empirical formula of its oxide containing 70.0% M.

13. How many moles of hydrogen, phosphorus and oxygen are there in 0.4 moles of phosphoric acid (H_3PO_4) ?

14. A solution has been prepared by dissolving 5.6 g of KOH in 250 mL of it. Calculate the molarity of the solution.

15. A chemist wishes to prepare 6.022×10^{24} molecules of SO_2 according to the reaction :

$$S + O_2 \longrightarrow SO_2$$

How many gram atoms of S and how many grams of O does he need ?

16. A sample of iron has a mass of 1.68 g. Calculate (*a*) the number of moles of iron present, (*b*) the number of atoms of iron in the sample.

17. 2.5 g of an impure sample of sodium bicarbonate when heated strongly gave 300 ml of carbon dioxide measured at 27°C and 760 mm pressure. Calculate the percentage purity of the sample.

18. Which is cheaper ?

40% HCl at the rate of ₹ 6 per kg or 80% H_2SO_4 at the rate of ₹ 3.5 per kg required to neutralise 7 kg of KOH.

19. The compound adrenaline is released in the human body in times of stress. It was found by experiment to have the composition 56.8 % C, 6.50 % H, 28.4% O and 8.28% N. What is the empirical formula of adrenaline?

20. What volume of concentrated aqueous sulphuric acid which is 98.0% H_2SO_4 by mass and has a density of 1.84 g mL^{-1} is required to prepare 10.0 L of 0.200 M H_2SO_4 solution ?

Hints & Answers *for* Revision Exercises

Very Short Answer Questions

10. 0.286

11. Both have same molar volume

12. He, valency = 0

14. 6.23×10^{-4}

17. (*i*) CO_2 (*ii*) C_6H_6

19. $3CaF_2 + 3H_2SO_4 + 2H_3BO_3 \longrightarrow 3CaSO_4 + 2BF_3 + 6H_2O$

20. 7.224×10^{23}

21. 12

22. One gram atom of iron means atomic mass of iron expressed in grams. It is equal to 56 g.

23. kg m^{-3}

24. Gay Lussac law

25. ^{12}C

26. 13 × 12 = 156 a.m.u.

27. C^{12}

28. 1 : 1

29. $\dfrac{250}{13.6} = 18.4 \text{ g cm}^{-3}$

Short Answer Questions

3. (*i*) m^3 (*ii*) m s^{-1} (*iii*) kg m s^{-2} or N

6. Compounds : (*i*), (*v*) ;
Mixtures : (*ii*), (*iii*), (*vi*), (*viii*), (*x*);
Elements (*iv*), (*vii*).

7. Tap water and clouds are homogeneous.

18. (*i*) $4H_3PO_3 \rightarrow 3H_3PO_4 + PH_3$
(*ii*) $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$
(*iii*) $Fe_2(SO_4)_3 + 3NH_3 + 3H_2O \longrightarrow 2Fe(OH)_3 + 3(NH_4)_2SO_4$

20. (*i*) (*c*) (*ii*) (*b*).

Long Answer Questions

6. (*a*) 4, 2, 4 moles (*b*) 21, 18, 37 g
(*c*) 6×10^3, 12×10^3, 12×10^3 H atoms
(*d*) 12×10^{20}, 10×10^{20}, 8×10^{20}

Numerical Problems

1. (*a*) 2, (*b*) 3, (*c*) 5, (*d*) 3, (*e*) 3

2. (*a*) 6.03, (*b*) 2.36, (*c*) 6.00×10^4, (*d*) 2.86

3. (*a*) $\dfrac{7.5 \times 206.8}{0.0512 \times 1002} = 30.232 = \mathbf{30.}$

7.5 contains only two significant figures and therefore, the result is reported to 2 significant figures.

(*b*) 4.20
1.6523
0.015
5.8673**= 5.87** [upto two decimal places (as in 4.20]

(*c*) (1.0042 – 0.0034) (1.23)
(1.0008) (1.23) = 1.230984 = **1.23** [upto 3 significant figures]

(*d*) $\dfrac{8.5 \times 208.9}{0.054 \times 9291.6} = 3.5389 = \mathbf{3.5}$ (upto 2 significant figures as in 8.5)

4. *Ist oxide*

Mass of carbon = 42.9 g
Mass of oxygen = 100 – 42.9 = 57.1 g

Second oxide

Mass of carbon = 27.3 g
Mass of oxygen = 100 – 27.3 = 72.7 g

Let us fix the mass of carbon as 1g

In first oxide,

Mass of oxygen combining with 42.9 g of carbon
= 57.1 g

Mass of oxygen combining with 1g of carbon

$$= \frac{57.1}{42.9} = 1.33 \text{ g}$$

In second oxide,

Mass of oxygen combining with 27.3 g of carbon

$= 72.7$ g

Mass of oxygen combining with 1g of carbon

$$= \frac{72.7}{27.3} = 2.66 \text{ g}$$

Ratio of oxygen combining with 1g of carbon

1.33 : 2.66

1 : 2

This is a simple whole number ratio and therefore, it is in agreement with **law of multiple proportions.**

5. (*a*) 1 gram atom of nitrogen = 14 g

2.5 gram atoms of nitrogen $= 14 \times 2.5 =$ **35 g.**

(*b*) 1 gram mole of carbon dioxide = 44 g

3.6 gram mole of carbon dioxide $= 44 \times 3.6 =$ **158.4 g.**

6. (*i*) 22.4 dm^3 of CO_2 at N.T.P. has molecules

$= 6.02 \times 10^{23}$

2.24 dm^3 of CO_2 at N.T.P. has molecules

$$= \frac{6.02 \times 10^{23}}{22.4} \times 2.24$$

$= \mathbf{6.02 \times 10^{22}}$.

(*ii*) Mass of 6.02×10^{23} atoms of oxygen = 16 g

Mass of 1 atom of oxygen

$$= \frac{16}{6.02 \times 10^{23}} = \mathbf{2.66 \times 10^{-23} \text{ g.}}$$

(*iii*) No. of ozone molecules in 2 mol $= 6.02 \times 10^{23} \times 2$

One ozone molecule = 3 atoms of oxygen

No. of oxygen atoms in 2 mol of ozone

$= 6.02 \times 10^{23} \times 2 \times 3$

$= \mathbf{3.61 \times 10^{24}}$.

(*iv*) Molecular mass of SO_2 = 64

64 g of SO_2 at N.T.P. occupy volume = 22.4 L

4.4 g of SO_2 at N.T.P. occupy volume

$$= \frac{22.4}{64} \times 4.4$$

$=$ **1.54 L.**

7. Atomic mass is the mass of 6.022×10^{23} atoms

Mass of 6.022×10^{23} atoms of nickel

$= 9.75 \times 10^{-23} \times 6.022 \times 10^{23} = 58.7$ g

$\therefore$ Atomic mass = **58.7.**

8. 2 g of hydrogen contain molecules $= 6.022 \times 10^{23}$

1000 g of hydrogen contain molecules

$$= \frac{6.022 \times 10^{23} \times 1000}{2} = \mathbf{3.01 \times 10^{26}}.$$

9. 1 mole of electrons $= 6.022 \times 10^{23}$

Total charge on 1 mole of electrons

$= 1.60 \times 10^{-19} \times 6.022 \times 10^{23}$

$= \mathbf{9.64 \times 10^4}$ **C.**

10. Mass of drop of rain $= 0.448 \times 1 = 0.448$ g

18 g of water contain molecules $= 6.022 \times 10^{23}$

0.448 g of water contains molecules

$$= \frac{6.022 \times 10^{23} \times 0.448}{18}$$

$= \mathbf{1.5 \times 10^{22}}$.

1 molecule of water = 2 H atoms

No. of hydrogen atoms $= 1.5 \times 10^{22} \times 2 = \mathbf{3.0 \times 10^{22}}$ **atoms.**

11. Volume of oxygen at N.T.P.

$p_1 = 750$ mm $p_2 = 760$ mm

$V_1 = 150$ mL $V_2 =$?

$T_1 = 273 + 20 = 293$ K $T_2 = 273$ K

Applying $\dfrac{p_1 V_1}{T_1} = \dfrac{p_2 V_2}{T_2}$ or $V_2 = \dfrac{p_1 V_1 T_2}{T_1 p_2}$

$$\therefore V_2 = \frac{750 \times 150 \times 273}{293 \times 760} = 137.9 \text{ mL}$$

22400 mL of oxygen at N.T.P. has molecules

$= 6.022 \times 10^{23}$

137.9 mL of oxygen at N.T.P. has molecules

$$= \frac{6.022 \times 10^{23} \times 137.9}{22400}$$

$= \mathbf{3.71 \times 10^{21}}$.

12.

Element	Percentage	Atomic mass	Atomic ratio	Simple atomic ratio	Simplest whole no. ratio
M	70	56	$\frac{70}{56} = 1.25$	$\frac{1.25}{1.25} = 1$	2
O	30	16	$\frac{50}{16} = 1.875$	$\frac{1.875}{1.25} = 1.5$	3

Empirical formula = $\mathbf{M_2O_3}$.

13. 0.4 moles of $H_3PO_4 \equiv 0.4 \times 3 = 1.2$ mol of H

$\equiv 0.4 \times 1 = 0.4$ mol of P

$\equiv 0.4 \times 4 = 1.6$ mol of O

14. Moles of KOH $= \dfrac{5.6}{56} = 0.1$

Volume = 250 mL

Molarity $= \dfrac{0.1}{250} \times 1000 =$ **0.4 M.**

15. Moles of SO_2 required $= \dfrac{6.022\times10^{24}}{6.022\times10^{23}} = 10$ mole

$$S + O_2 \longrightarrow SO_2$$

1 mole of SO_2 is prepared from = 1 gram atom of S

10 mole of SO_2 are prepared from = 10 gram atoms of S

1 mole of SO_2 require O_2 = 32 g

10 mole of SO_2 require O_2 = 32 × 10 = **320 g.**

16. (*a*) 56 g of iron = 1 mole

1.68 g of iron $= \dfrac{1}{56} \times 1.68 = \mathbf{0.03\ mol}$

(*b*) 1 mole of iron $= 6.022 \times 10^{23}$ atoms

0.03 mole of iron $= 6.022 \times 10^{23} \times 0.03 = \mathbf{1.8 \times 10^{22}}$.

17. $2NaHCO_3 \longrightarrow Na_2CO_3 + CO_2 + H_2O$

2 (23 + 1 + 12 + 48) = 168 g   22.4 L at N.T.P.

Volume of CO_2 at N.T.P.

p_1 = 760 mm, p_2 = 760 mm

V_1 = 300 mL, V_2 = ?

T_1 = 273 + 27 = 300 K, T_2 = 273

Applying $\dfrac{p_1V_1}{T_1} = \dfrac{p_2V_2}{T_2}$ or $V_2 = \dfrac{p_1V_1T_2}{p_2T_1}$

$$\therefore \quad V_2 = \frac{760\times300\times273}{760\times300} = 273 \text{ mL}$$

Calculation of pure $NaHCO_3$

22400 ml of CO_2 at N.T.P. are produced by heating 168 g of $NaHCO_3$

273 ml of CO_2 at N.T.P. are produced by heating

$$NaHCO_3 = \frac{168}{22400}\times273 = 2.0475$$

Wt. of sample taken = 2.5 g

$$\% \text{ Purity} = \frac{2.0475}{2.5}\times100 = \mathbf{81.9\%}.$$

18. $KOH + HCl \longrightarrow KCl + H_2O$

56   36.5

56 g of KOH require HCl for neutralisation = 36.5 g

7×10^3 g of KOH require HCl for neutralisation

$$= \frac{36.5}{56}\times7\times10^3$$

$$= 4.56 \times 10^3 \text{ g} = 4.56 \text{ kg}$$

$$\text{Wt. of 40\% HCl} = \frac{4.56\times100}{40} = 11.4 \text{ kg}$$

Cost = 11.4 × 6 = ₹ 68.4

$2KOH + H_2SO_4 \longrightarrow K_2SO_4 + 2H_2O$

112 g   98 g

112 g of KOH require H_2SO_4 for neutralisation = 98 g

7×10^3 g of KOH require H_2SO_4 for neutralisation

$$= \frac{98}{112} \times 7 \times 10^3 \text{ g} = 6.125 \text{ kg}$$

$$\text{Wt. of 80\% } H_2SO_4 = \frac{6.125\times100}{80} = 7.66 \text{ kg}$$

Cost = 7.66 × 3.50 = ₹ 26.81

∴ **It is cheaper to neutralise KOH by H_2SO_4 than HCl.**

19.

Element	Percentage	Atomic mass	Moles of atoms	Mole ratio	Simple whole no. ratio
C	56.8	12	$\frac{50.8}{12} = 4.73$	$\frac{4.73}{0.59} = 8$	8
H	6.50	1	$\frac{6.50}{1} = 6.50$	$\frac{6.50}{0.59} = 11$	11
O	28.4	16	$\frac{28.4}{16} = 1.77$	$\frac{1.77}{0.59} = 3$	3
N	8.28	14	$\frac{8.28}{14} = 0.59$	$\frac{0.59}{0.59} = 1$	1

Empirical formula = $\mathbf{C_8H_{11}O_3N}$.

20. 98.0% H_2SO_4 means that 98 g of H_2SO_4 is present in 100 g of solution.

$$\text{Moles of } H_2SO_4 = \frac{98}{98} = 1 \text{ mol}$$

$$\text{Volume of solution} = \frac{100}{1.84} = 54.35 \text{ mL}$$

$$\text{Molarity} = \frac{1}{54.35}\times1000 = \mathbf{18.4\ M}$$

Now applying molarity equation,

$$M_1V_1 = M_2V_2$$

$$18.4 \times V_1 = 0.20 \times 10.0\text{L}$$

$$\therefore \quad V_1 = \frac{0.20\times10}{18.4} = 0.1087\text{L}$$

$$= \mathbf{108.7\ mL.}$$

Competition File

Additional Useful Information and Objective Questions

ADDITIONAL USEFUL INFORMATION

CONCEPT OF EQUIVALENT MASS

The **equivalent mass of a substance** is defined as *the number of parts by mass of a substance which combines with or displaces directly or indirectly 1.008 parts by mass of hydrogen or 8 parts by mass of oxygen or 35.5 parts by mass of chlorine.*

The following examples clearly illustrate the definition of equivalent mass :

- *Chlorine combines with hydrogen to form hydrogen chloride; HCl as :*

$$\underset{2}{H_2} + \underset{71}{Cl_2} \longrightarrow 2\,HCl$$

Hence, 2 parts by mass of hydrogen combine with 71 parts by mass of chlorine.

or 1 part by mass of hydrogen will combine with 35.5 parts by mass of chlorine.

Hence, **equivalent mass of chlorine is 35.5.**

[For simplicity 1 part by mass of hydrogen may be taken, instead of 1.008 parts]

- *Magnesium wire burns in oxygen gas to form magnesium oxide :*

$$\underset{2 \times 24}{Mg} + \underset{32}{O_2} \longrightarrow 2MgO$$

32 parts by mass of oxygen combine with 48 parts by mass of magnesium.

8 parts by mass of oxygen will combine with 12 parts by mass of magnesium.

Hence, **equivalent mass of magnesium is 12.**

- *Zinc metal reacts with dil. H_2SO_4 to liberate hydrogen gas :*

$$\underset{65}{Zn} + \underset{98}{H_2SO_4} \longrightarrow ZnSO_4 + \underset{2}{H_2}$$

2 parts by mass of hydrogen is liberated by 65 parts by mass of zinc.

1 part by mass of hydrogen is liberated by 32.5 parts by mass of zinc.

Hence, **equivalent of mass of zinc = 32.5**

Similarly,

2 parts by mass of hydrogen ≡ 98 parts by mass of H_2SO_4

1 part by mass of hydrogen ≡ 49 parts by mass of H_2SO_4

Hence, **equivalent mass of H_2SO_4 = 49**

Equivalent mass can also be calculated from the formulae of the compounds.

For example :

- **$AlCl_3$:** 3 × 35.5 parts by mass of chlorine combine with 27 parts by mass of Al.

35.5 parts by mass of chlorine will combine with 9 parts by mass of Al

∴ **Equivalent mass of Al = 9**

- **FeO :** 16 parts by mass of oxygen ≡ 56 parts by mass of Fe

8 parts by mass of oxygen ≡ 28 parts by mass of Fe

∴ **Equivalent mass of Fe = 28.**

Gram equivalent mass

The mass of a substance in grams which is numerically equal to its equivalent mass is called **gram equivalent mass** or **one gram equivalent**. It is simply the equivalent mass of a substance expressed in grams. For example, 1 gram equivalent of oxygen is 8 gram because equivalent mass of oxygen is 8.

$$\textbf{Gram-equivalents} = \frac{\textbf{Mass in gm}}{\textbf{Equivalent Mass}}$$

▸ It may be noted that the elements with *variable valency* generally exhibit variable equivalent masses. The equivalent mass of an element is related to its valency as,

$$\text{Equivalent mass} = \frac{\text{Atomic mass}}{\text{Valency}}$$

Since atomic mass of an element is always constant, but its valency may change in different compounds, therefore, its equivalent mass changes with change in valency in its compounds.

Equivalent mass of acids, bases and salts

Equivalent mass of an acid *is the number of parts by mass of it which can give one part by mass of H^+ ions in its aqueous solutions or one part by mass of replaceable hydrogen.*

The number of H^+ ions that one molecule of an acid can give in its aqueous solution is known as its **basicity**.

Therefore,

$$\text{Equivalent mass of an acid} = \frac{\text{Molecular mass}}{\text{Basicity}}$$

Competition File

For example,

$HCl \rightleftharpoons H^+ + Cl^-$ Basicity = 1

$$\text{Eq. mass} = \frac{\text{Mol. mass}}{1}$$

$H_2SO_4 \rightleftharpoons 2H^+ + SO_4^{2-}$ Basicity = 2

$$\text{Eq. mass} = \frac{\text{Mol. mass}}{2}$$

and $H_3PO_4 \rightleftharpoons 3H^+ + PO_4^{3-}$ Basicity = 3

$$\text{Eq. mass} = \frac{\text{Mol. mass}}{3}$$

Equivalent mass of a base *is the number of parts by mass of it which can neutralise one equivalent of an acid or it is the number of parts by mass of it which can furnish one equivalent or 17 parts by mass of OH^- ions* in an aqueous solution.

The number of OH^- ions that one molecule of a base can give in its aqueous solution is called its **acidity**. Therefore,

$$\text{Equivalent mass of a base} = \frac{\text{Molecular mass}}{\text{Acidity}}$$

For example,

$NaOH \rightleftharpoons Na^+ + OH^-$ Acidity = 1

$$\text{Eq. mass} = \frac{\text{Mol. mass}}{1}$$

$Ca(OH)_2 \rightleftharpoons Ca^{2+} + 2OH^-$ Acidity = 2

$$\text{Eq. mass} = \frac{\text{Mol. mass}}{2}$$

$Al(OH)_3 \rightleftharpoons Al^{3+} + 3OH^-$ Acidity = 3

$$\text{Eq. mass} = \frac{\text{Mol. mass}}{3}$$

Equivalent mass of a salt

The equivalent mass of a salt is that mass of it which contains one equivalent mass of metal.

Equivalent mass of salt

$$= \frac{\text{Mol. mass of salt}}{\text{Total valency of the metal in the salt}}$$

The total valency of a metal is equal to the valency of the metal ion multiplied by the number of atoms present in the formula of the salt. For example,

Salt	*Valency of metal ion*	*No. of metal atoms*	*Total valency*	*Mol. mass*	*Eq. mass*
NaCl	1	1	1 × 1 = 1	58.5	58.1/1 = 58.5
K_2CO_3	1	2	1 × 2 = 2	138	138/2 = 69
$NaHCO_3$	1	1	1 × 1 = 1	84	84/1 = 84
$CaCO_3$	2	1	2 × 1 = 2	100	100/2 = 50
$FeCl_3$	3	1	3 × 1 = 3	162.5	162.5/3 = 54.2
$Al_2(SO_4)_3$	3	2	3 × 2 = 6	342	342/6 = 57

Equivalent mass of an ion

Equivalent mass of an ion is obtained by dividing the formula mass of the ion with its valency (or charge)

$$\text{Equivalent mass of ion} = \frac{\text{Formula mass}}{\text{Charge on ion}}$$

For example,

$$\text{Equivalent mass of } SO_4^{2-} \text{ ion} = \frac{32 + 4 \times 16}{2} = 48$$

Equivalent mass of oxidising and reducing agent in a reaction

A given substance may have different equivalent masses in various reactions. In a simple way, *the equivalent mass of an oxidising agent is calculated from the number of oxygen atoms it produces and equivalent mass of a reducing agent is calculated from the number of oxygen atoms with which it combines. For example,*

(*i*) *Equivalent mass of $KMnO_4$ in the reaction :*

$$2\,KMnO_4 + 3\,H_2SO_4 \rightarrow K_2SO_4 + 2MnSO_4 + 3H_2O + 5(O)$$

2(39 + 55 + 64) = 316 ... 5 × 16

5 × 16 parts of oxygen are liberated from $KMnO_4$ = 316 parts

Competition File

8 parts of oxygen are liberated from

$$KMnO_4 = \frac{316}{5 \times 16} \times 8 = 31.6$$

$\therefore$ **Eq. mass of $KMnO_4$ = 31.6**

(*ii*) *Equivalent mass of $K_2Cr_2O_7$ in the reaction :*

$$\underset{\substack{2 \times 39 + 2 \times 52 + 7 \times 16 \\ = 294}}{K_2Cr_2O_7} + 4H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + 4H_2O + \underset{\substack{3 \times 16 \\ = 48}}{3(O)}$$

48 parts of oxygen are liberated from $K_2Cr_2O_3$ = 294

8 parts of oxygen are liberated from $K_2Cr_2O_7$

$$= \frac{294}{48} \times 8 = 49$$

$\therefore$ **Eq. mass of $K_2Cr_2O_7$ = 49**

(*iii*) *Equivalent mass of $FeSO_4$ in the reaction :*

$$\underset{\substack{2(56 + 32 + 64) \\ = 304}}{2FeSO_4} + H_2SO_4 + \underset{16}{O} \longrightarrow Fe_2(SO_4)_3 + H_2O$$

16 parts of oxygen react with $FeSO_4$ = 304 parts

8 parts of oxygen react with $FeSO_4 = \frac{304}{16} \times 8 = 152$

$\therefore$ **Eq. mass of $FeSO_4$ = 152**

(*iv*) *Equivalent mass of $Na_2S_2O_3$ in the reaction :*

$$\underset{\substack{2(46 + 64 + 48) \\ = 316}}{2Na_2S_2O_3} + H_2O + \underset{16}{O} \longrightarrow Na_2S_4O_6 + 2NaOH$$

16 parts of oxygen react with $Na_2S_2O_3$ = 316 parts

8 parts of oxygen react with $Na_2S_2O_3 = \frac{316}{16} \times 8 = 158$

$\therefore$ **Eq. mass of $Na_2S_2O_3$ = 158**

DULONG AND PETIT'S LAW

According to Dulong and Petit's law, the product of atomic mass and specific heat for a solid element is approximately equal to 6.4.

Atomic mass × Specific heat = 6.4

This relation can be used to calculate atomic mass.

First approximate atomic mass may be calculated as:

$$\text{Approximate atomic mass} = \frac{6.4}{\text{Specific heat}}$$

From this, valency can be calculated as

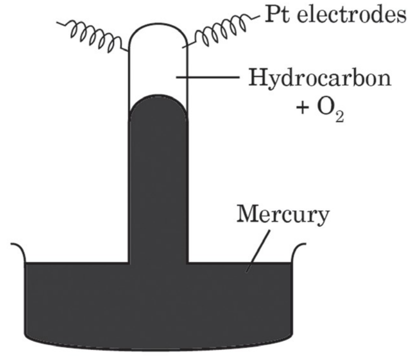
$$\text{Valency} = \frac{\text{Approx. atomic mass}}{\text{Equivalent mass}}$$

Take the valency to nearest whole number, then

Exact atomic mass = Equivalent mass × Valency

EUDIOMETRY

This method has been used for determining the molecular formula or percentage composition of a gaseous mixture of hydrocarbons. The apparatus consists of a closed graduated tube open at one end.



Eudiometer tube

The closed end is provided with platinum electrodes for the passage of electricity through the gas. Such tube is called **eudiometer tube**. It is filled with mercury and inverted over a trough containing mercury. The method involves the following steps:

(*i*) A known volume of the gaseous hydrocarbon or gaseous mixture mixed with excess of oxygen or air is introduced in the eudiometer tube over mercury.

(*ii*) The mixture is exploded by electric spark so that C and H are oxidised to CO_2 and H_2O respectively.

(*iii*) The mixture is then cooled so that water vapour condense to liquid water whose volume is negligible as compared to the volumes of other gases.

(*iv*) KOH is introduced which absorbs CO_2 and only unused O_2 is left. Thus,

decrease in volume on adding KOH = volume of CO_2 produced.

The volume of unused O_2 can be measured by absorbing in alkaline solution of pyrogallol or calculated by difference method.

Eudiometry is based on Avogadro's law which states that equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules or moles. This means that volume ratio among gases is same as mole ratio at same conditions of temperature and pressure.

The equation for combustion of hydrocarbon is

$$C_xH_y(g) + \left(x+\frac{x}{4}\right) O_2(g) \longrightarrow x\ CO_2(g) + \frac{y}{2} H_2O(l)$$

$$1 \text{ mol} \quad \left(x+\frac{y}{4}\right) \text{mol} \longrightarrow x \text{ mol} \quad + \quad \frac{y}{2} \text{ mol}$$

$$1 \text{ cc} \quad \left(x+\frac{y}{4}\right) \text{cc} \longrightarrow x \text{ cc} \quad -$$

Values of x and y can be calculated from the following relations:

Volume of O_2 used per cc of hydrocarbon

$$= \left(x+\frac{y}{4}\right) \text{cc}$$

Volume of CO_2 produced = x cc

Contraction in volume on explosion and cooling

$$= \left[1+\left(x+\frac{y}{4}\right)\right] - x = \left(1+\frac{y}{4}\right)\text{cc}$$

Illustration 1. *8.0 mL of a gaseous hydrocarbon was exploded with 45 mL of oxygen. The volume of gases on cooling was found to be 37 mL, 16 mL of which was absorbed by KOH and rest was absorbed in an alkaline solution of pyrogallol. If all volumes are measured under same conditions, deduce the formula of the hydrocarbon.*

Solution.

Volume of CO_2 produced = 16 mL (absorbed by KOH)

Contraction in volume on cooling

$$= 8.0 + 45 - 37 = 16 \text{ mL}$$

Now, volume of CO_2 produced for 1 mL of hydrocarbon

$$= x \text{ mL}$$

Volume of CO_2 produced for 8 mL of hydrocarbon

$$= 8\,x \text{ mL}$$

$$\therefore \quad 8x = 16 \quad \text{or} \quad x = 2$$

Contraction in volume on cooling for 8 mL of hydrocarbon = $8\left(1+\frac{y}{4}\right)$

$$\therefore \quad 8\left(1+\frac{y}{4}\right) = 16 \quad \text{or} \quad 8 + 2y = 16 \quad \text{or} \quad y = 4$$

Hence, formula of compound is C_2H_4.

Illustration 2. *10 mL of a gaseous hydrocarbon was burnt completely in 100 mL of O_2 at NTP. On cooling the gas occupied 80 mL at N.T.P. This volume became 50 mL on treatment with KOH solution. What is the formula of the hydrocarbon?*

Solution. Volume of CO_2 produced + Volume of unreacted O_2 = 80 mL

Volume of unreacted O_2 (CO_2 absorbed by KOH) = 50 mL

Volume of O_2 reacted with 10 mL of hydrocarbon = 100 – 50 = 50 mL

Volume of CO_2 produced by 10 mL of hydrocarbon and 50 mL of O_2 = 80 – 50 = 30 mL

$$C_x H_y(g) + \left(x+\frac{y}{4}\right) O_2(g) \longrightarrow x\ CO_2(g) + \frac{y}{2} H_2O\ (l)$$

Volume of CO_2 produced by 10 mL hydrocarbon,

$$10\,x = 30$$

$$\therefore \quad x = 3$$

Volume of O_2 reacted with 10 mL hydrocarbon,

$$10\left(x+\frac{y}{4}\right) = 50 \text{ mL}$$

$$x + \frac{y}{4} = 5$$

$$3 + \frac{y}{4} = 5 \text{ or } \frac{y}{4} = 2$$

$$\therefore \quad y = 8$$

Hence, formula of hydrocarbon is C_3H_8.

Competition File

OBJECTIVE TYPE QUESTIONS

A Topicwise MULTIPLE CHOICE QUESTIONS *with only one correct answer*

Select the Correct Answers :

Matter, Physical Measurements, Laws of Combination

A1. The number of significant figures in 0.0101 is :
(*a*) 3 (*b*) 2
(*c*) 4 (*d*) 5.

A2 Two elements A and B combine to form two compounds in which '*a*' gm of A combines with b_1 and b_2 gm of B respectively. According to law of multiple proportions,
(*a*) $b_1 = b_2$
(*b*) b_1 and b_2 bear a simple whole number ratio
(*c*) a_1 and b_1 bear whole number ratio
(*d*) no relation exists between b_1 and b_2.

A3. Two different oxides of a metal contain 20% and 27% oxygen by weight. This is in accordance with the law of :
(*a*) conservation of mass (*b*) constant composition
(*c*) multiple proportion (*d*) reciprocal proportion.

A4. Two elements X and Y have atomic masses 14 and 16 respectively. They can form a series of five compounds A, B, C, D and E in which for the same amount of X, Y is present in the ratio of 1: 2: 3: 4: 5. If the compound A has 28 parts by weight of X and 16 parts by weight of Y, then compound C will have 24 parts by weight of Y combined with
(*a*) 28 parts by weight of X
(*b*) 14 parts by weight of X
(*c*) 7 parts by weight of X
(*d*) 4.1 parts by weight of X

A5. Two oxides of an element contain 57.1% and 72.7% of oxygen. If the first oxide is MO, the second oxide is :
(*a*) MO (*b*) M_2O
(*c*) MO_3 (*d*) MO_2.

A6. Calculate to the correct number of significant figures: 4.26 – (15.635/5.0)
(*a*) 1.16 (*b*) 1.12
(*c*) 1.2 (*d*) 1.133.

Mole Concept, Atomic Mass and Molecular Mass

A7. Which of the following weighs least ?
(*a*) 2.0 gram mole of CO_2
(*b*) 0.1 mole of sucrose ($C_{12}H_{22}O_{11}$)
(*c*) 1 gram atom of calcium
(*d*) 1.5 mole of water.

A8. Which of the following has maximum number of atoms?
(*a*) 18 g of water (*b*) 16 g of O_2
(*c*) 4.4 g of CO_2 (*d*) 16 g of CH_4.

A9. Which of the following contains more molecules ?
(*a*) 1 g CO_2 (*b*) 1 g N_2.
(*c*) 1 g H_2 (*d*) 1 g CH_4.

A10. Which of the following weighs least :
(*a*) 2.24 litres of CO_2 at N.T.P.
(*b*) 6.02×10^{21} molecules of CO_2
(*c*) 1 g of CO_2
(*d*) 6.02×10^{22} atoms of carbon.

A11. The number of atoms present in 0.1 mole of P_4 are :
(*a*) 2.4×10^{23} atoms
(*b*) 6.02×10^{22} atoms
(*c*) same as in 0.2 mole of S_8
(*d*) same as in 3.1 g of phosphorus
(atomic mass of P = 31).

A12. The total number of atoms in 8.5 g of NH_3 is :
(*a*) 9.03×10^{23} (*b*) 3.01×10^{23}
(*c*) 1.204×10^{23} (*d*) 6.02×10^{23}

A1 The number of atoms of oxygen present in 11.2 L of ozone at N.T.P. are :
(*a*) 3.01×10^{22} (*b*) 6.02×10^{23}
(*c*) 9.03×10^{24} (*d*) 1.20×10^{24}.

A14. The number of atoms of He in 104 a.m.u. is :
(*a*) 3.1×10^{25} (*b*) 6.2×10^{25}
(*c*) 26 (*d*) 206.

A15. The number of water molecules present in a drop of water weighing 0.018 g is :
(*a*) 6.02×10^{26} (*b*) 6.02×10^{23}
(*c*) 6.02×10^{20} (*d*) 6.02×10^{19}.

A16. 34.2 g of sucrose ($C_{12}H_{22}O_{11}$) are dissolved in 90 g of water in a glass. The number of oxygen atoms in the solution are :
(*a*) 3.66×10^{26} (*b*) 6.6×10^{23}
(*c*) 3.66×10^{24} (*d*) 6.0×10^{22}.

A17. 1 mole of methane contains :
(*a*) 6.02×10^{23} atoms of H
(*b*) 4 gram atoms of hydrogen
(*c*) 1.81×10^{23} molecules of methane
(*d*) 3.0 g of carbon.

A18. The molar masses of oxygen and sulphur dioxide are 32 and 64 respectively. If 1 L of oxygen at 25°C and 750 mm Hg pressure contains N molecules, then the number of molecules in 2 L sulphur dioxide under same conditions of temperature and pressure is :
(*a*) N/2 (*b*) 3N/2 (*c*) 2N (*d*) 6 N.

Answers

A1. (*a*) **A2.** (*b*) **A3.** (*c*) **A4.** (*b*) **A5.** (*d*) **A6.** (*c*) **A7.** (*d*) **A8.** (*d*) **A9.** (*c*) **A10.** (*b*) **A11.** (*a*)
A12. (*a*) **A13.** (*c*) **A14.** (*c*) **A15.** (*c*) **A16.** (*c*) **A1** (*b*) **A18.** (*c*)

Competition File

A19. Which of the following contain highest number of atoms ?
(*a*) 1.0 g of water (*b*) 1.0 g of silver
(*c*) 1.0 g of nitrogen (*d*) 1.0 g of propane (C_3H_8).

A20. Which of the following has maximum mass ?
(*a*) 1.0 mole of H_2 gas
(*b*) 0.5 mole of sucrose ($C_{12}H_{22}O_{11}$)
(*c*) 1.2 mole of silver
(*d*) 22.4 L of N_2 at N.T.P.

A21. Number of molecules present in 1 ml of water is :
(*a*) 1 (*b*) 1000
(*c*) 2.69×10^{19} (*d*) 6.02×10^{20}.

A22 2 g of oxygen contain number of atoms equal to that contained by :
(*a*) 0.5 g hydrogen (*b*) 4 g sulphur
(*c*) 7 g nitrogen (*d*) 2.3 g sodium.

A23. 40 g of caustic soda contain :
(*a*) 6.02×10^{23} atoms of H
(*b*) 22.4 litres of N_2
(*c*) 6.02×10^{24} molecules of O_2
(*d*) 4 g of Na.

A24. 0.6 g of carbon was burnt in air to form CO_2. The number of molecules of CO_2 introduced into air will be :
(*a*) 6.02×10^{22} (*b*) 3.01×10^{22}
(*c*) 6.023×10^{23} (*d*) 3.01×10^{23}.

A25. The total number of electrons present in 3.2 g of methane are :
(*a*) $2 \times 10 \times 6.022 \times 10^{22}$ (*b*) $10 \times 6.022 \times 10^{23}$
(*c*) $10 \times 6.022 \times 10^{22}$ (*d*) 6.022×10^{23}.

A26. The number of atoms in 4.25 g of NH_3 is approximately
(*a*) 1×10^{23} (*b*) 2×10^{23}
(*c*) 4×10^{23} (*d*) 6×10^{23}

A27. Haemoglobin contains 0.33% of iron by weight. The molecular mass of haemoglobin is about 67200. The number of iron atoms (at mass of Fe = 56) present in one molecule of haemoglobin is
(*a*) 6 (*b*) 4 (*c*) 2 (*d*) 1

A28 How many moles of electron weigh one kilogram ?
(*a*) 6.022×10^{23} (*b*) $\frac{1}{9.108} \times 10^{31}$
(*c*) $\frac{6.022}{9.108} \times 10^{54}$ (*d*) $\frac{1}{9.108 \times 6.022} \times 10^{8}$

A29. An alkaloid contains 17.28% of nitrogen and its molecular mass is 162. The number of nitrogen atoms present in one molecule of alkaloid is
(*a*) five (*b*) four
(*c*) three (*d*) two

A30. Number of atoms in 558.6 g Fe (atomic mass of Fe = 55.86 g mol^{-1}) is
(*a*) twice that in 60 g carbon
(*b*) 6.023×10^{22}
(*c*) half that of 8 g He
(*d*) $558.6 \times 6.023 \times 10^{23}$

Percentage Composition, Molecular formula and Stoichiometry

A31. The empirical formula of sucrose is :
(*a*) CH_2O (*b*) CHO
(*c*) $C_{12}H_{22}O_{11}$ (*d*) $C(H_2O)_2$.

A32. One mole of calcium phosphide on reaction with excess of water gives
(*a*) One mole of phosphine
(*b*) Two moles of phosphoric acid
(*c*) Two moles of phosphine
(*d*) One mole of phosphorus pentoxide

A33. A sample of water contains *x* % of D_2O. Its molecular weight is 19. The value of '*x*' is
(*a*) 25 (*b*) 50 (*c*) 33.33 (*d*) 75

A34. A compound contains 8% sulphur. The minimum molecular weight of the compound is
(*a*) 100 (*b*) 200 (*c*) 350 (*d*) 400

A35. An aqueous solution of 6.3 g of oxalic acid dihydrate is made upto 250 mL. The volume of 0.1 N NaOH required to completely neutralize 10 mL of this solution is :
(*a*) 40 mL (*b*) 20 mL
(*c*) 10 mL (*d*) 5 mL

A36 Mixture X = 0.02 mol of $[Co(NH_3)_5SO_4]Br$ and 0.02 mol of $[Co(NH_3)_5Br]SO_4$ was prepared in 2 L of solution :
1 L of mixture X + excess of $AgNO_3 \longrightarrow$ Y
1 L of mixture X + excess of $BaCl_2 \longrightarrow$ Z
Number of moles of Y and Z are :
(*a*) 0.01, 0.01 (*b*) 0.02, 0.01
(*c*) 0.01, 0.02 (*d*) 0.02, 0.02

A37. In Haber process, 30 L of dihydrogen and 30 L of dinitrogen were taken for reaction which yielded only 50% of the expected product. What will be the composition of the gaseous mixture under the aforesaid conditions in the end ?
(*a*) 20 L NH_3, 25 L N_2 and 20 L H_2
(*b*) 10 L NH_3, 25 L N_2 and 15 L H_2
(*c*) 20 L NH_3, 10 L N_2 and 30 L H_2
(*d*) 20 L NH_3, 25 L N_2 and 15 L H_2

A38. A gaseous mixture contains 50% He and 50% CH_4 by volume. What is the percent by weight of CH_4 in the mixture ?
(*a*) 19.97% (*b*) 20.05%
(*c*) 50% (*d*) 80.03%

Answers

A19. (*d*) **A20.** (*b*) **A21.** (*c*) **A22.** (*b*) **A23.** (*a*) **A24.** (*b*) **A25.** (*a*) **A26.** (*d*) **A27.** (*b*) **A28.** (*d*) **A29.** (*d*)
A30. (*a*) **A3** (*c*) **A32.** (*c*) **A33.** (*b*) **A34.** (*d*) **A35.** (*a*) **A36.** (*a*) **A37.** (*b*) **A38.** (*d*)

Competition File

A39. The mass of carbon anode consumed (giving only carbon dioxide) in the production of 270 kg of aluminium metal from bauxite by Hall process is :

(*a*) 180 kg (*b*) 270 kg
(*c*) 540 kg (*d*) 90 kg

A40. The crystalline salt $Na_2SO_4.xH_2O$ on heating loses 55.9 % of its weight. The formula of crystalline salt is

(*a*) $Na_2SO_4.5H_2O$ (*b*) $Na_2SO_4.7H_2O$
(*c*) $Na_2SO_4.2H_2O$ (*d*) $Na_2SO_4 . 10H_2O$

A4 20 kg of $N_2(g)$ and 3.0 kg of $H_2(g)$ are mixed to produce $NH_3(g)$. The amount of $NH_3(g)$ formed is

(*a*) 17 kg (*b*) 51 kg
(*c*) 60 kg (*d*) 34

A42. A phosphorus oxide has 43.6% phosphorus (at. mass = 31). The empirical formula of the compound is

(*a*) P_2O_5 (*b*) P_4O_6
(*c*) P_2O_3 (*d*) P_4O_8

A43. Commercially available concentrated HCl contains 38·0% HCl by mass (density = 1·19 g mL^{-1}). The molarity of the solution is

(*a*) 10·40 M (*b*) 5·70 M
(*c*) 12·38 M (*d*) 13·46 M

A44. The molarity of a solution obtained by mixing 800 mL of 0.5 M HCl with 200 mL of 1 M HCl will be

(*a*) 0.8 M (*b*) 0.6 M
(*c*) 0.4 M (*d*) 0.2 M

A45. 4L of water is added to 2L of 6M HCl. The molarity of the final solution is

(*a*) 4 M (*b*) 2 M
(*c*) 1 M (*d*) 0·5 M

A4 The volume of 10.50 M solution required to prepare 1.0 L of 0.25 M solution of HNO_3 is :

(*a*) 250 mL (*b*) 500 mL
(*c*) 230 mL (*d*) 23.8 mL.

A47 The moles of sodium chloride in 250 cm^3 of 0.50 M NaCl are :

(*a*) 0.250 mol (*b*) 2 m
(*c*) 0.125 mol (*d*) 1.0 mol.

A48. 6 mL of a gaseous hydrocarbon was exploded with excess of oxygen and the product cooled. A contraction of 9 mL was observed. A further contraction of 12 mL was observed on treatment with aqueous KOH. The formula of hydrocarbon is

(*a*) CH_4 (*b*) C_2H_4
(*c*) C_2H_6 (*d*) C_2H_2

Answers

A39. (*d*) **A40.** (*d*) **A41.** (*a*) **A42.** (*a*) **A43.** (*c*) **A44.** (*b*) **A45** (*b*) **A46** (*d*) **A47.** (*c*) **A48** (*d*)

B MULTIPLE CHOICE QUESTIONS *from Competitive Examinations*

AIPMT & Other State Boards' Medical Entrance

B1. What volume of oxygen gas (O_2) measured at 0°C and 1 atm, is needed to burn completely 1 L of propane gas (C_3H_8) measured under the same conditions?

(*a*) 6 L (*b*) 5 L
(*c*) 10 L (*d*) 7 L

(C.B.S.E. PMT 2008)

B2. How many moles of lead (II) chloride will be formed from a reaction between 6.5 g of PbO and 3.2 g of HCl?

(*a*) 0.333 (*b*) 0.011
(*c*) 0.029 (*d*) 0.044

(C.B.S.E. PMT 2008)

B3. Volume occupied by one molecule of water (density = 1 g cm^{-3}) is :

(*a*) 6.023×10^{-23} cm^3 (*b*) 3.0×10^{-23} cm^3
(*c*) 5.5×10^{-23} cm^3 (*d*) 9.0×10^{-23} cm^3

(C.B.S.E. PMT 2008)

B Which of the following concentration terms is/are independent of temperature ?

(*a*) Molality only
(*b*) Molality and mole fraction
(*c*) Molarity and mole fraction
(*d*) Molality and normality
(*e*) Molarity only

(Kerala PMT 2009)

B5. 10 g of hydrogen and 64 g of oxygen were filled in a steel vessel and exploded. Amount of water produced in this reaction will be

(*a*) 3 mol (*b*) 4 mol
(*c*) 1 mol (*d*) 2 mol

(C.B.S.E. PMT. 2009)

B6. The number of molecules in 100 ml of 0.02 N H_2SO_4 is

(*a*) 6.02×10^{22} (*b*) 6.02×10^{21}
(*c*) 6.02×10^{20} (*d*) 6.02×10^{18}

(AMU Med. 2010)

B7. If 1.5 moles of oxygen combine with Al to form Al_2O_3, the mass of Al in g [Atomic mass of Al = 27] used in the reaction is

(*a*) 2.7 (*b*) 54
(*c*) 40.5 (*d*) 81
(*e*) 27

(Kerala PMT 2010)

B8. For a reaction A + 2B → C, the amount of C formed by starting the reaction with 5 moles of A and 8 moles of B is

(*a*) 5 moles (*b*) 8 moles
(*c*) 16 moles (*d*) 4 moles
(*e*) 1 mole

(Kerala PMT 2010)

B9. One kilogram of a sea water sample contains 6 mg of dissolved O_2. The concentration of O_2 in the sample in ppm is

(*a*) 0.6 (*b*) 6.0
(*c*) 60.0 (*d*) 16.0
(*e*) 32.0

(Kerala PMT 2010)

Answers

B1. (*b*) **B2.** (*c*) **B3.** (*b*) **B4.** (*b*) **B5.** (*b*) **B6.** (*c*) **B7.** (*b*) **B8.** (*d*) **B9.** (*b*)

Competition File

B10. 25.3 g of sodium carbonate, Na_2CO_3 is dissolved in enough water to make 250 mL of solution. If sodium carbonate dissociates completely, molar concentration of Na^+ and CO_3^{2-} ions are respectively (molar mass of Na_2CO_3 = 106 g mol^{-1})

(*a*) 0.477 M and 0.477 M

b) 0.955 M and 1.910 M

(*c*) 1.910 M and 0.955 M

(*d*) 1.90 M and 1.910 M *(C.B.S.E PMT 2010)*

B11. The number of atoms in 0.1 mol of triatomic gas is ($N_A = 6.02 \times 10^{23}$)

(*a*) 1.800×10^{22} (*b*) 6.026×10^{22}

(*c*) 1.806×10^{23} (*d*) 3.600×10^{22}

(C.B.S.E. PMT 2010)

B12. Which one of the following sets of compounds correctly illustrate the law of reciprocal proportions ?

(*a*) P_2O_3, PH_3, H_2O (*b*) P_2O_5, PH_3, H_2O

(*c*) N_2O_5, NH_3, H_2O (*d*) N_2O, NH_3, H_2O

(*e*) NO_2, NH_3, H_2O *(Kerala P.M.T. 2011)*

B13. 20.0 kg of N_2(g) and 3.0 kg of H_2(g) are mixed to produce NH_3(g). The amount of NH_3(g) formed is

(*a*) 17 kg (*b*) 34 kg

(*c*) 20 kg (*d*) 3 k

(*e*) 23 kg *(Kerala P.M.T. 2011)*

B14. Mole fraction of the solute in a 1.00 molal aqueous solution is

(*a*) 0.1770 (*b*) 0.0177

(*c*) 0.0344 (*d*) 1.7700

(A.I.P.M.T. 2011)

B15. What is the volume of CO_2 liberated (in litres) at 1 atmosphere and 0°C when 10 g of 100% pure calcium carbonate is treated with excess dilute sulphuric acid? (Atomic mass Ca = 40, C = 12, O = 16)

(*a*) 0.224 (*b*) 2.24

(*c*) 22.4 (*d*) 224

e) 11.2 *(Kerala P.M.T. 2012)*

B16. Which one of the following is the lightest ?

a) 0.2 mole of hydrogen gas.

(*b*) 6.023×10^{22} molecules of nitrogen.

c) 0.1 g of silver.

(*d*) 0.1 mole of oxygen gas.

(*e*) 1 g of water. *(Kerala P.M.T. 2012)*

B17 When 22.4 litres of H_2(*g*) is mixed with 11.2 litres of Cl_2(*g*), each at S.T.P, the moles of HCl(*g*) formed is equal to

(*a*) 1 mol of HCl(*g*) (*b*) 2 mol of HCl(*g*)

(*c*) 0.5 mol of HCl(*g*) (*d*) 1.5 mol of HCl(*g*)

(AIPMT 2014)

B18 1.0 g of magnesium is burnt with 0.56 g O_2 in a closed vessel. Which reactant is left in excess and how much? (At. wt. Mg = 24, O = 16)

(*a*) Mg, 0.16 g (*b*) O_2, 0.16 g

(*c*) Mg, 0.44 g (*d*) O_2, 0.28 g

(AIPMT 2014)

B19. The mass of $CaCO_3$ required to react completely with 20 mL of 1.0 M HCl as per the reaction:

$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$ is

(At. mass: Ca = 40, C = 12, O = 16)

(*a*) 1 g (*b*) 2 g

(*c*) 10 g (*d*) 20 g

(*e*) 200 g *(Kerala PMT 2015)*

B20. Which one of the following has maximum number of molecules?

(*a*) 16 g of O_2 (*b*) 16 g of NO_2

(*c*) 4 g of N_2 (*d*) 2 g of H_2

(*e*) 32 g of N_2 *(Kerala PMT 2015)*

B21. A mixture of gases contains H_2 and O_2 gases in the ratio of 1 : 4 (w/w). What is the molar ratio of the two gases in the mixture?

(*a*) 16 : 1 (*b*) 2 : 1

(*c*) 1 : 4 (*d*) 4 : 1

(A.I.P.M.T. 2015)

B22. If Avogadro number N_A is changed from 6.022×10^{23} mol^{-1} to 6.022×10^{20} mol^{-1}, this would change:

a) the ratio of chemical species to each other in a balanced equation.

b) the ratio of elements to each other in a compound.

(*c*) the definition of mass in units of grams.

(*d*) the mass of one mole of carbon.

(A.I.P.M.T. 2015)

B23. The number of water molecules is maximum in:

(*a*) 18 gram of water

(*b*) 18 moles of water

c) 18 molecules of water.

(*d*) 1.8 gram of water. *(A.I.P.M.T. 2015)*

B24. 20.0 g of a magnesium carbonate sample decomposes on heating to give carbon dioxide and 8.0 g magnesium oxide. What will be the percentage purity of magnesium carbonate in the sample?

(*a*) 60 (*b*) 84

(*c*) 75 (*d*) 96

(A.I.P.M.T. 2015)

B25. What is the mole fraction of the solute in a 1.00 m aqueous solution?

(*a*) 0.0354 (*b*) 0.0177

(*c*) 0.177 (*d*) 1.770

(A.I.P.M.T. 2015)

Answers

B10. (*c*) **B11.** (*c*) **B12.** (*a*) **B13.** (*a*) **B14.** (*b*) **B15.** (*b*) **B16.** (*c*) **B17.** (*a*) **B18.** (*a*) **B19.** (*a*)
B20. (*e*) **B21.** (*d*) **B22.** (*d*) **B23.** (*b*) **B24.** (*b*) **B25.** (*b*)

Competition File

B26. What is the mass of the precipitate formed when 50 mL of 16.9% solution of $AgNO_3$ is mixed with 50 mL of 5.8% NaCl solution?
(Ag = 107.8, N = 14, O = 16, Na = 23, Cl = 35.5)
(*a*) 7 g (*b*) 14 g
(*c*) 28 g (*d*) 3.5 g
(A.I.P.M.T. 2015)

B27. Suppose the elements X and Y combine to form two compounds XY_2 and X_3Y_2. When 0.1 mole of XY_2 weighs 10 g and 0.05 mole of X_3Y_2 weighs 9 g, the atomic weights of X and Y are
(*a*) 40, 30 (*b*) 60, 40
(*c*) 20, 30 (*d*) 30, 20 *(NEET 2016)*

B28. Which of the following is dependent on temperature?
(*a*) Molality (*b*) Molarity
(*c*) Mole fraction (*d*) Weight percentage
(NEET 2017)

B29. In which case is number of molecules of water maximum?
(*a*) 18 mL of water
(*b*) 0.18 g of water
(*c*) 0.00224 L of water vapours at 1 atm and 273 K
d) 10^{-3} mol of water *(NEET 2018)*

JEE (Main) & Other State Boards' Engineering Entrance

B30. If we consider that 1/6 in place of 1/12 mass of carbon atom is taken to be the relative atomic mass unit, the mass of one mole of a substance will
(*a*) decrease twice
(*b*) increase two fold
(*c*) remain unchanged
(*d*) be a fraction of molecular mass of the substance
(A.I.E.E.E. 2005)

B31. In the reaction :
$2Al(s) + 6HCl(aq) \longrightarrow 2Al^{3+}(aq) + 6Cl^-(aq) + 3H_2(g)$
(*a*) 33.6 L $H_2(g)$ is produced regardless of temperature and pressure for every mole of Al that reacts.
(*b*) 67.2 L $H_2(g)$ at STP is produced for every mole of Al that reacts.
(*c*) 11.2 L H_2 (*g*) at STP is produced for every mole of HCl (*aq*) consumed.
(*d*) 6 L HCl (*aq*) is consumed for every 3 L of $H_2(g)$ is produced. *(A.I.E.E.E. 2007)*

B32. 80 g of oxygen contain as many atoms as in
(*a*) 10 g of hydrogen
(*b*) 5 g of hydrogen
(*c*) 80 g of hydrogen
(*d*) 1 g of hydrogen *(Karnataka C.E.T. 2008)*

B33. What is the molality of pure water?
(*a*) 1 (*b*) 18
(*c*) 55.5 (*d*) None of these
(Orissa JEE 2008)

B34. The volume of 10N and 4 N HCl required to make 1 L of 7 N HCl are
(*a*) 0.50 L of 10 N HCl and 0.50 L of 4 N HCl
(*b*) 0.60 L of 10 N HCl and 0.40 L of 4 N HCl
(*c*) 0.80 L of 10 N HCl and 0.20 L of 4 N HCl
(*d*) 0.75 L of 10 N HCl and 0.25 L of 4 N HCl
(Orissa J.E.E. 2009)

B35. The molarity of 98% H_2SO_4 (*d* = 1.8 g/mL) by weight is
(*a*) 6 M (*b*) 18 M
(*c*) 10 M (*d*) 4 M *(MP-PET 2009)*

B36. 20 mL of 10 N HCl are mixed with 10 mL of 36 N H_2SO_4 and the mixture is made 1L. Normality of the mixture will be
(*a*) 0.56 N (*b*) 0.50 N
(*c*) 0.40 N (*d*) 0.35 N
(MP-PET 2009)

B37. Excess of carbon dioxide is passed through 50 mL of 0.5 M calcium hydroxide solution. After the completion of the reaction, the solution was evaporated to dryness. The solid calcium carbonate was completely neutralised with 0.1 N hydrochloric acid. The volume of hydrochloric acid required is
(*a*) 200 cm^3 (*b*) 500 cm^3
(*c*) 400 cm^3 (*d*) 300 cm^3
(Karnataka C.E.T. 2009)

B38. How much time (in hours) would it take to distribute one Avogadro number of wheat grains if 10^{20} grains are distributed each second ?
(*a*) 0.1673 (*b*) 1.673
(*c*) 16.73 (*d*) 167.3
(*e*) 1673 *(Kerala PET 2010)*

B39. Two oxides of a metal contain 36.4% and 53.4% of oxygen by mass respectively. If the formula of the first oxide is M_2O, then that of the second is
(*a*) M_2O_3 (*b*) MO
(*c*) MO_2 (*d*) M_2O_5
(J.K. C.E.T. 2011)

B40. Arrange the following in the order of increasing mass (atomic mass : O = 16, Cu = 63, N = 14)
I. one atom of oxygen
II. one atom of nitrogen
III. 1×10^{-10} mole of oxygen
IV. 1×10^{-10} mole of copper
(*a*) II < I < III < IV (*b*) I < II < III < IV
(*c*) III < II < IV < I (*d*) IV < II < III < I
(*e*) II < IV < I < III
(Kerala P.E.T. 2011)

Answers

B26. (*a*) **B27.** (*a*) **B28.** (*b*) **B29.** (*a*) **B30.** (*a*) **B31.** (*c*) **B32.** (*b*) **B33.** (*c*) **B34.** (*a*) **B35.** (*b*)
B36. (*a*) **B37.** (*b*) **B38.** (*b*) **B39.** (*b*) **B40.** (*a*)

Competition File

B41. A mixture of ethane and ethene occupies 41 L at 1 atm and 500 K. The mixture reacts completely with $\frac{10}{3}$ mole of O_2 to produce CO_2 and H_2O. The mole fractions of ethane and ethene in the mixture are (R = 0.082 L atm K^{-1} mol^{-1}) respectively

(*a*) 0.50, 0.50 (*b*) 0.75, 0.25
(*c*) 0.67, 0.33 (*d*) 0.25, 0.75
(*e*) 0.33, 0.67 *(Kerala P.E.T. 2011)*

B42. A mixture of $CaCl_2$ and NaCl weighing 4.44 g is treated with sodium carbonate solution to precipitate all the calcium ions as calcium carbonate. The calcium carbonate so obtained is heated strongly to get 0.56 g of CaO.
The percentage of NaCl in the mixture is [Atomic mass of Ca = 40]

(*a*) 31.5 (*b*) 75
(*c*) 25 (*d*) 40.2
(Karnataka C.E.T. 2011)

B43. 50 cm^3 of 0.2 N HCl is titrated against 0.1 N NaOH solution. The titration was discontinued after adding 50 cm^3 of NaOH. The remaining titration is completed by adding 0.5 N KOH. The volume of KOH required for completing the titration is

(*a*) 10 cm^3 (*b*) 12 cm^3
(*c*) 16.2 cm^3 (*d*) 21.0 cm^3
(Karnataka C.E.T. 2011)

B44. A 100% pure sample of a divalent metal carbonate weighing 2 g on complete thermal decomposition releases 448 cc of carbon dioxide at STP. The equivalent mass of the metal is

(*a*) 40 (*b*) 20
(*c*) 28 (*d*) 12
(*e*) 56 *(Kerala P.E.T. 2012)*

B45. The total number of electrons in 18 mL of water (density = 1 g mL^{-1}) is

(*a*) 6.02×10^{23} (*b*) 6.02×10^{25}
(*c*) 6.02×10^{24} (*d*) $6.02 \times 18 \times 10^{23}$
(Karnataka C.E.T. 2012)

B46. Two solutions of HCl, A and B, have concentrations of 0.5 M and 0.1 M respectively. The volume of solutions A and B required to make 2 litres of 0.2 M HCl are

a) 0.5 L of A + 1.5 L of B
b) 1.5 L of A + 0.5 L of B
(*c*) 1.0 L of A + 1.0 L of B
d) 0.75 L of A + 1.25 L of B *(A.M.U. Engg. 2012)*

B47. The number of water molecules present in a drop of water weighing 0.018 g is

(*a*) 6.022×10^{26} (*b*) 6.022×10^{23}
(*c*) 6.022×10^{19} (*d*) 6.022×10^{20}
(Karnataka CET 2013)

B48. How many grams of concentrated nitric acid solution should be used to prepare 250 mL of 2.0 $MHNO_3$? The concentrated acid is 70% HNO_3.

(*a*) 70.0 g conc. HNO_3
b) 54.0 g conc. HNO_3
(*c*) 45.0 g conc. HNO_3
(*d*) 90.0 g conc. HNO_3 *(J.K. CET 2013)*

B49. The molarity of a solution obtained by mixing 750mL of 0.5 M HCl with 250 mL of 2 M HCl will be

(*a*) 0.975 M (*b*) 0.875 M
(*c*) 1.00 M (*d*) 1.75 M
(JEE Main 2013)

B50. A gaseous hydrocarbon gives upon combustion 0.72 g of water and 3.08 g of CO_2. The empirical formula of the hydrocarbon is

(*a*) C_7H_8 (*b*) C_2H_4
(*c*) C_3H_4 (*d*) C_6H_6 *(JEEMain2013)*

B51. 10 g of a mixture of BaO and CaO requires 100 cm^3 of 2.5M HCl to react completely. The percentage of calcium oxide in the mixture is approximately (Given molar mass of BaO = 153, CaO = 56).

(*a*) 52.6 (*b*) 55.1
(*c*) 44.9 (*d*) 47.4
(KarnatakaCET2014)

B52. 25 cm^3 of oxalic acid completely neutralised 0.064 g of sodium hydroxide. Molarity of oxalic acid solution is

(*a*) 0.064 (*b*) 0.045
(*c*) 0.015 (*d*) 0.032
(KarnatakaCET2014)

B53. A 5.82 g silver coin is dissolved in nitric acid. When sodium chloride is added to the solution, all the silver is precipitated as AgCl. The AgCl precipitate weighs 7.20 g. The percentage of silver in the coin is

(*a*) 60.3% (*b*) 80%
(*c*) 93.1% (*d*) 70%
(A.M.U. Egg. 2014)

B54 25 mL of 3.0 M HCl are mixed with 75 mL of 4.0 M HCl. If the volumes are additive, the molarity of the final mixture will be

(*a*) 4.0 M (*b*) 3.75 M
(*c*) 4.25 M (*d*) 3.50 M
(A.M.U. Egg. 2014)

B55. The ratio of masses of oxygen and nitrogen in a particular gaseous mixture is 1 : 4. The ratio of number of their molecules is

(*a*) 3 : 16 (*b*) 1 : 4
(*c*) 7 : 32 (*d*) 1 : 8 *(JEEMain2014)*

B56. The number of Cl^- ions in 100 mL of 0.001 M HCl solution is

(*a*) 6.022×10^{23} (*b*) 6.022×10^{20}
(*c*) 6.022×10^{19} (*d*) 6.022×10^{24}
(A.M.U. Engg 2015)

Answers

B41. (*c*) **B42.** (*b*) **B43.** (*a*) **B44.** (*b*) **B45.** (*c*) **B46.** (*a*) **B47.** (*d*) **B48.** (*c*) **B49.** (*b*) **B50.** (*a*)
B51. (*a*) **B52.** (*d*) **B53.** (*c*) **B54.** (*b*) **B55.** (*c*) **B56.** (*c*)

Competition File

B57 0.30 g of an organic compound containing C, H and O on combustion yields 0.44 g CO_2 and 0.18 g H_2O. If one mol of compound weighs 60, then molecular formula of the compound is
(*a*) C_3H_8O (*b*) $C_2H_4O_2$
(*c*) CH_2O (*d*) C_4H_6O
(Karnataka CET 2015)

B58. What amount of dioxygen (in gram) contains 1.8×10^{22} molecules?
(*a*) 0.960 (*b*) 96.0
(*c*) 0.0960 (*d*) 9.60
(Karnataka CET 2015)

B59. If 27g of water is formed during complete combustion of pure propene (C_3H_6), the mass of propene burnt is
(*a*) 42 g (*b*) 21 g
(*c*) 14 g (*d*) 56 g
(*e*) 40 g *(Kerala PET 2016)*

B60. When 2.46 g of a hydrated salt ($MSO_4 . xH_2O$) is completely dehydrated, 1.20 g of anhydrous salt is obtained. If the molecular weight of anhydrous salt is 120 g mol^{-1}, what is the value of x?
(*a*) 2 (*b*) 4
(*c*) 5 (*d*) 6
(*e*) 7 *(Kerala PET 2016)*

B61. Calculate the molality of a solution that contains 51.2g of naphthalene, ($C_{10}H_8$) in 500 mL of carbon tetrachloride. The density of CCl_4 is 1.60 g/mL.
(*a*) 0.250 m (*b*) 0.500 m
(*c*) 0.750 m (*d*) 0.840 m
(*e*) 1.69 m *(Kerala PET 2016)*

B62. The number of oxygen atoms in 4.4 g of CO_2 is
(*a*) 1.2×10^{23} (*b*) 6×10^{22}
(*c*) 6×10^{23} (*d*) 12×10^{23}
(Karnataka CET 2016)

B63. An organic compound contains C = 40%, H = 13.33% and N = 46.67%. Its empirical formula is
(*a*) C_2H_2N (*b*) C_3H_7N
(*c*) CH_4N (*d*) CHN
(Karnataka CET 2016)

B64. At 300 K and 1 atm, 15 mL of a gaseous hydrocarbon requires 375 mL air containing 20% O_2 by volume for complete combustion. After combustion the gases occupy 330 mL. Assuming that the water formed is in liquid form and the volumes were measured at the same temperature and pressure the formula of the hydrocarbon is
(*a*) C_3H_6 (*b*) C_3H_8
(*c*) C_4H_8 (*d*) C_4H_{10}
(JEE Main 2016)

B65. You are supplied with 500 mL each of 2 N HCl and 5N HCl. What is the maximum volume of 3 M HCl that you can prepare using only these two solutions?
(*a*) 250 mL (*b*) 500 mL
(*c*) 750 mL (*d*) 1000 mL
(WB-JEE 2017)

B66. 0.126 g of an acid is needed to completely neutralise 20 mL of 0.1 NaOH solution. The equivalent weight of the acid is
(*a*) 53 (*b*) 40
(*c*) 45 (*d*) 63 *(WB-JEE 2017)*

B67. In a flask, the weight ratio of CH_4(g) and SO_2(g) at 298 K and 1 bar is 1 : 2. The ratio of the number of molecules of $SO_2(g)$ and $CH_4(g)$ is
(*a*) 1 : 4 (*b*) 4 : 1
(*c*) 1 : 2 (*d*) 2 : 1 *(WB-JEE 2017)*

B68. What will be the normality of the salt solution obtained by neutralizing x mL of y (N) HCl with y mL of x (N) NaOH and finally adding $(x + y)$ mL distilled water ?
(*a*) $\frac{2(x+y)}{xy}$ N (*b*) $\frac{xy}{2(x+y)}$ N
(*c*) $\left(\frac{2xy}{x+y}\right)$ N (*d*) $\left(\frac{x+y}{xy}\right)$ N
(WB-JEE 2017)

B69. If 3.01×10^{20} molecules are removed from 98 mg of H_2SO_4, then number of moles of H_2SO_4 left are
(*a*) 0.5×10^{-3} mol (*b*) 0.1×10^{-3} mol
(*c*) 9.95×10^{-2} mol (*d*) 1.66×10^{-3} mol
(Karnataka CET 2017)

B70. 10 g of $MgCO_3$ decomposes on heating to 0.1 mole CO_2 and 4 g MgO. The percent purity of $MgCO_3$ is
(*a*) 24% (*b*) 44%
(*c*) 54% (*d*) 74%
(*e*) 84% *(Kerala PET 2017)*

B71. The compound $Na_2CO_3 . x H_2O$ has 50% H_2O by mass. The value of 'x' is
(*a*) 4 (*b*) 5
(*c*) 6 (*d*) 7
(*c*) 8 *(Kerala PET 2017)*

B72. 1 gram of a carbonate (M_2CO_3) on treatment with excess HCl produces 0.01186 mole of CO_2. The molar mass of M_2CO_3 in g mol^{-1} is
(*a*) 1186 (*b*) 84.3
(*c*) 118.6 (*d*) 11.86
(Kerala PET 2017)

B73. Calculate the molarity of a solution of 30 g of $Co(NO_3)_2.6H_2O$ in 4.3 L of solution? Consider atomic mass of Co = 59u, N = 14u, O = 16u, H = 1u.
(*a*) 0.023 M (*b*) 0.23 M
(*c*) 0.046 M (*d*) 0.46 M
(J.K. CET 2018)

Answers

B57. (*b*) **B58.** (*a*) **B59** (*b*) **B60** (*e*) **B61.** (*b*) **B62.** (*a*) **B63.** (*c*) **B64.** (*b*) **B6** (*c*) **B66.** (*d*)
B67 (*c*) **B68.** (*b*) **B69.** (*a*) **B70.** (*e*) **B71.** (*c*) **B72.** (*b*) **B73.** (*a*)

Competition File

B7 How many moles of electrons will weigh one kilogram?

(*a*) 6.023×10^{23} (*b*) $\frac{1}{9.108} \times 10^{21}$

(*c*) $\frac{6.023}{9.108} \times 10^{54}$ (*d*) $\frac{1}{9.108 \times 6.023} \times 10^{8}$

(WB JEE 2018)

B75. A metal M (specific heat 0.16) forms a metal chloride with a 65% chlorine present in it. The formula of the metal chloride will be

(*a*) MCl (*b*) MCl_2

(*c*) MCl_3 (*d*) MCl_4

(WB JEE 2018)

B76 1.0 g of Mg is burnt with 0.28 g of O_2 in a closed vessel. Which reactant is left in excess and how much?

(*a*) Mg, 5.8 g (*b*) Mg, 0.58 g

(*c*) O_2, 0.24 g (*d*) O_2, 2.4 g

(Karnataka CET 2018)

B77. 1 mol of $FeSO_4$ (atomic weight of Fe is 55.84 g mol^{-1}) is oxidised to $Fe_2(SO_4)_3$. Calculate the equivalent weight of ferrous ion.

(*a*) 55.84 (*b*) 27.92

(*c*) 18.61 (*d*) 111.68

(*e*) 83.76 *(Kerala PET 2018)*

B78. Mass % of carbon in ethanol is

(*a*) 52 *b*) 13

(*c*) 34 *d*) 90

e) 80 *(Kerala PET 2018)*

B79. The ratio of mass percent of C and H of an organic compound ($C_xH_yO_z$) is 6 : 1. If one molecule of the above compound ($C_xH_yO_z$) contains half as much oxygen as required to burn one molecule of compound C_xH_y completely to CO_2 and H_2O, the empirical formula of compound $C_xH_yO_z$ is :

(*a*) $C_3H_6O_3$ (*b*) C_2H_4O

(*c*) $C_3H_4O_2$ (*d*) $C_2H_4O_3$

(JEE Main 2018)

B80. 8 g of NaOH is dissolved in 18 g of H_2O. Mole fraction of NaOH in solution and molality (in mol kg^{-1}) of the solution respectively are

(*a*) 0.167, 11.11 (*b*) 0.2, 22.20

(*c*) 0.2, 11.11 (*d*) 0.167, 22.20

(JEE Main 2019)

B81. The amount of sugar ($C_{12}H_{22}O_{11}$) required to prepare 22 of its 0.1 M aqueous solution is

(*a*) 68.4 g (*b*) 17.1 g

(*c*) 34.2 g (*d*) 136.8 g

(JEE Main 2019)

B82. A solution of sodium sulphate contains 92 g of Na^+ ions per kilogram of water. The molality of Na^+ ions in that solution in mol kg^{-1} is

(*a*) 16 (*b*) 8

(*c*) 4 (*d*) 12

(JEE Main 2019)

JEE (Advance) for IIT Entrance

B83. Given that the abundances of isotopes ^{54}Fe, ^{56}Fe and ^{57}Fe are 5%, 90% and 5% respectively, the atomic mass of Fe is

(*a*) 55.85 (*b*) 55.95

(*c*) 55.75 (*d*) 56.05 ***(IIT-JEE 2009)***

Answers

B74. (*d*) **B75.** (*b*) **B76.** (*b*) **B77.** (*a*) **B78.** (*a*) **B79.** (*d*) **B80.** (*a*) **B81.** (*a*) **B82.** (*c*) **B83.** (*b*)

MULTIPLE CHOICE QUESTIONS
with more than one correct answers

C1. Which of the following method/methods of expressing concentrations is/are independent of temperature ?

(*a*) Mole fraction (*b*) Molarity

(*c*) Normality (*d*) Molality

C2. A solution contains 25% water, 25% ethanol (C_2H_5OH) and 50% acetic acid (CH_3COOH) by mass. The mole fraction of

a) Water = 0.502

(*b*) Ethanol = 0.302

(*c*) Acetic acid = 0.196

d) Ethanol + acetic acid = 0.498

C3. 8 g of O_2 has the same number of oxygen atoms as

(*a*) 11 g CO_2 (*b*) 14 g CO

(*c*) 32 g SO_2 (*d*) 8 g O_3

C4. The mass of $\frac{1}{12}$ th of ^{12}C is same as that of

(*a*) $\frac{1}{28}$ th of N_2 (*b*) 1u

(*c*) $\frac{1}{8}$ th of O (*d*) $\frac{1}{12}$ th of He

C5. In $MgSO_4$ (at. mass : Mg = 24, S = 32, O = 16), the mass percentage of

(*a*) Mg = 80% (*b*) Mg = 20%

(*c*) S = 26.7% (*d*) S = 53.3%

Answers

C1. (*a,d*) **C2.** (*a,d*) **C3.** (*a,b,d*) **C4.** (*a,b*) **C5.** (*b,c*)

Competition File

C6. The following substances are present in different containers

(*i*) One gram atom of nitrogen

(*ii*) One mole of calcium

(*iii*) One atom of silver

(*iv*) One mole of oxygen molecules

(*v*) 10^{23} atoms of carbon and

(*vi*) One gram of iron.

The correct order of increasing masses (in grams) is/are :

(*a*) (*iii*) < (*vi*) < (*i*) < (*v*)

(*b*) (*iii*) < (*vi*) < (*iv*) < (*ii*)

(*c*) (*vi*) < (*v*) < (*i*) < (*iv*)

(*d*) (*iii*) < (*ii*) < (*v*) < (*iv*)

C7. Which of the following methods of expressing concentration varies with temperature?

(*a*) Molality (*b*) Weight per cent

(*c*) Normality (*d*) Molarity

C8. Which of the following units are not correct for the physical quantity ?

(*a*) Acceleration : ms^{-2} (*b*) Pressure : $kg\ m^{-2}s^{-2}$

(*c*) Power : Js (*d*) Frequency : s^{-1}

Answers

C6. (*b*, *c*) **C7.** (*c*, *d*) **C8.** (*b*, *c*)

D MULTIPLE CHOICE QUESTIONS
based on the given passage/comprehension

Passage-I

A mole is a collection of 6.022×10^{23} particles and the number 6.022×10^{23} is called Avogadro number. The mass of this number of atoms in an element is equal to its gram atomic mass and mass of this number of molecules in a compound is equal to its gram molecular mass. The volume occupied by this number of molecules of a gas at N.T.P is 22.4 L. When 6.022×10^{23} molecules of a substance are dissolved in 1 L of solution, the solution is known as 1 molar volume.

Answer the following questions :

D1. The mass of 10 molecules of naphthalene ($C_{10}H_8$) is

(*a*) 2.12×10^{-22} g (*b*) 2.12×10^{-21} g

(*c*) 2.12×10^{-23} g (*d*) 1280 g

D2. Suppose the chemists would have choosen 10^{20} as the number of particles in a mole, the mass of 1 mole of oxygen gas would be

(*a*) 5.32×10^{-43} g (*b*) 5.32×10^{-3} g

(*c*) 5.32×10^{-23} g (*d*) 5.32×10^{3} g

D3. One million atoms of silver (at. mass = 107.81) atoms weigh

(*a*) 1.79×10^{-16} g (*b*) 3.58×10^{-16} g

(*c*) 3.58×10^{-6} g (*d*) 1.79×10^{-16} g

Passage-II

The earlier method for determining the molecular weight of proteins was based on chemical analysis. The following composition of proteins were found :

Haemoglobin : 0.335% Fe

Cytochrome protein : 0.376% Fe

Peroxidase enzyme : 0.29% Se

Answer the following questions :

D4. If haemoglobin contains 4 atoms of iron, then approximate molecular mass of haemoglobin is (at. mass of Fe = 55.85)

(*a*) 16700 (*b*) 33400 (*c*) 66800 (*d*) 1670

D5. The mole % of Se in the enzyme peroxidase is (at. mass of Se = 78.96)

(*a*) 2.16×10^{-3} (*b*) 2.7×10^{5}

(*c*) 3.67×10^{-3} (*d*) 1.83×10^{-3}

D6. If the cytochrome protein contains one atom per molecule then the molecular mass of protein is

(*a*) 14850 u (*b*) 29600 u

(*c*) 32960 u (*d*) 12840 u

D7. How many atoms of Se are present in 1 μg of peroxidase enzyme assuming one molecule of enzyme contains 1 atom of Se (at. mass of Se = 78.96) ?

(*a*) 2.23×10^{19} (*b*) 4.52×10^{14}

(*c*) 3.82×10^{21} (*d*) 2.23×10^{6}

D8. How many moles of iron are present in 1 mg of haemoglobin (assuming a molecule of haemoglobin contains 4 Fe atoms)?

(*a*) 1.50×10^{-8} (*b*) 6.0×10^{-8}

(*c*) 3.0×10^{-8} (*d*) 1.875×10^{-9}

Answers

Passage I. **D1.** (*b*) **D2.** (*b*) **D3.** (*a*) **Passage II.** **D4.** (*c*) **D5.** (*c*) **D6** (*a*) **D7.** (*b*) **D8.** (*b*)

Competition File

Passage III

Oleum or fuming sulphuric acid contains SO_3 dissolved in sulphuric acid and has the molecular formula $H_2S_2O_7$. It is formed by passing SO_3 in H_2SO_4. When water is added to oleum, SO_3 reacts with water to form H_2SO_4.

$$SO_3\,(g) + H_2O\,(l) \longrightarrow H_2SO_4\,(aq)$$

As a result, mass of H_2SO_4 increases. When 100 g sample of oleum is diluted with desired amount of water (in gram) then the total mass of pure H_2SO_4 obtained after dilution is known as percentage labelling of oleum.

% Labelling of oleum = Total mass of H_2SO_4 present in oleum after dilution

or = Mass of H_2SO_4 initially present + Mass of H_2SO_4 produced after dilution

From this, the percentage composition of H_2SO_4 and SO_3 (free) and SO_3 (combined) can be calculated.

Answer the following questions :

D9. The percentage of SO_3 in 109% H_2SO_4 is

(*a*) 9% (*b*) 36%

(*c*) 40% (*d*) 60%

D10. The percentage of free SO_3 and H_2SO_4 in 112% H_2SO_4 is

(*a*) 53.6, 46.4 (*b*) 12.0, 88.0

(*c*) 88.0, 12.0 (*d*) 26.8, 73.2

Answers

Passage III. D9. (*c*) **D10.** (*a*)

Assertion Reason Type Questions

The questions given below consist of an Assertion and the Reason. Use the following key to choose the appropriate answer.

(*a*) If both assertion and reason are CORRECT and reason is the CORRECT explanation of the assertion.

(*b*) If both assertion and reason are CORRECT, but reason is NOT THE CORRECT explanation of the assertion.

(*c*) If assertion is CORRECT, but reason is INCORRECT.

(*d*) If assertion is INCORRECT but reason is CORRECT.

(*e*) If both assertion and reason are INCORRECT.

1. ssertion : 22 carat gold is a compound.

Reason : A compound has fixed composition of the elements present in it.

2. Assertion : Both 32 g of SO_2 and 8 g of CH_4 contain same number of molecules.

Reason : Equal moles of two compounds contain same number of molecules.

3. Assertion : The standard unit for expressing the mass of atoms is a.m.u.

Reason : a.m.u. stands for mass of 1 atom of carbon.

4. Assertion : The sum of 154.2 + 6.1 + 23 is 183.

Reason : The result of addition is reported to the same number of decimal places as that of the term with least number of decimal places.

5. Assertion : 1 mol of O and 1 mol of O_2 contain equal number of particles.

Reason : 1 mol of molecules is always double than 1 mol of atoms in all diatomic molecules.

6. Assertion : Graphite is an element.

Reason : Element is the pure form of a substance containing same kind of atoms.

7. Assertion : Steam is a mixture.

Reason : In a compound, the composition of the elements must be fixed.

8. Assertion : Empirical and molecular formula of Na_2CO_3 is same.

Reason : Na_2CO_3 does not form hydrate.

9. Assertion : The empirical mass of ethene is half of its molecular mass.

Reason : The empirical formula represents the simplest whole number ratio of various atoms present in a compound.

10. Assertion : Significant figures for 0.200 is 3 whereas for 200 it is 1.

Reason : Zero at the end or right of a number are significant provided they are not on the right side of the decimal point.

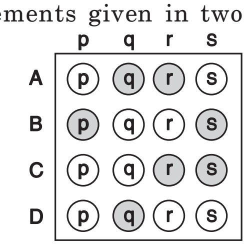
Answers

1. (*d*) **2.** (*a*) **3.** (*c*) **4.** (*a*) **5.** (*c*) **6.** (*a*) **7.** (*d*) **8.** (*c*) **9.** (*a*) **10.** (*c*).

Competition File

Matrix Match Type Questions

Each question contains statements given in two columns, which have to be matched. Statements in Column I are labelled as A, B, C and D whereas statements in Column II are labelled as p, q, r and s. Match the entries of Column I with appropriate entries of Column II. Each entry in Column I may have one or more than one correct option from Column II. The answers to these questions have to be appropriately bubbled as illustrated in the following example.



If the correct matches are A-q, A-r, B-p, B-s, C-r, C-s and D-q, then the correctly bubbled matrix will look like the following:

1. Match the physical quantity given in Column I with the symbol or definition in Column II

Column I	Column II
(A) Force	(p) J
(B) Energy	(q) s^{-1}
(C) Frequency	(r) kg $\,^{2}s^{-1}$
(D) Work	(s) kg ms^{-2}

2. Match the term given in Column I with the description given in Column II.

Column I	Column II
(A) Molality	(p) independent of temperature
(B) Molarity	(q) mol L^{-1}
(C) Mole fraction	(r) g equiv L^{-1}
(D) Normality	(s) mol kg^{-1}

3. Match the following :

Column I	Column II
(A) 88 g of CO_2	(p) 0.25 mol
(B) 6.022×10^{23} molecules of H_2O	(q) 2 mol
(C) 5.6 litres of O_2 at STP	(r) 1 mol
(D) 96 g of O_2	(s) 6.022×10^{23} molecules
(E) 1 mol of any gas	(t) 3 mol

Answers

(1) : (A) – (s) (B) – (p), (r) (C) – (q) (D) – (p), (s)
(2) : (A) – (p), (s) (B) – (q) (C) – (p) (D) – (r)
(3) : (A) – (q) (B) – (r) (C) – (p) (D) – (t) (E) – (s)

Integer Type and Numerical Value Type Questions

Integer Type: The answer to each of the following question is a **single digit integer** ranging from 0 to 9.

1. 1.420 g of anhydrous $ZnSO_4$ was left in moist air. After a few days its weight was found to be 2.528 g. How many water molecules are present in its hydrated salt formula (molar mass of $ZnSO_4$ = 161.5)?

2. Moles of iron which can be made from $Fe_2O_3$by the use of 294 g of carbon monoxide in the reaction :

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + CO_2 \text{ are :}$$

3. The ratio of molecular mass/empirical formula mass of a hydrocarbon having 92.4% C and vapour density of 39 is

4. 428 mL of 10 M HCl and 572 mL of 3 M HCl are mixed. The molarity of the resulting solution is

5. Silver (atomic mass = 108 g mol^{-1}) has a density of 10.5 cm^{-3}. The number of silver atoms on a surface of area $10^{-12}m^2$ can be expressed in scientific notation as $y \times 10^x$. The value of x is

(I.I.T. JEE 2010)

6. The value of n in the formula $Be_n\ Al_2\ Si_6\ O_{18}$ is

(I.I.T. 2010)

7. Reaction of Br_2 with Na_2CO_3 in aqueous solution gives sodium bromide and sodium bromate with evolution of CO_2 gas. The number of sodium bromide molecules involved in the balanced chemical equation is

(I.I.T. 2011)

8. 29.2% (w/w) HCl stock solution has a density of 1.25 g mL^{-1}. The molecular weight of HCl is 36.5 g mol^{-1}. The volume (mL) of stock solution required to prepare a 200 mL solution of 0.4 M HCl is

(I.I.T. J.E.E. 2012)

Answers

1. (7) **2.** (7) **3.** (6) **4.** (6) **5.** (7) **6.** (3) **7.** (5) **8.** (8)

Competition File

9. If the value of Avogadro number is 6.023×10^{23} and the value of Boltzmann constant is 1.380×10^{-23} JK^{-1}, then number of significant digits in the calculated value of the universal constant is *(JEE Advance 2014)*

Numerical Value Type: Give the correct numerical value (in decimal notation truncated/rounded off to the second decimal place.

10. To measure the quantity of $MnCl_2$ dissolved in an aqueous solution, it was completely converted to $KMnO_4$ using the reaction,

$$MnCl_2 + K_2S_2O_8 + H_2O \rightarrow KMnO_4 + H_2SO_4 + HCl$$

(equation not balanced)

Few drops of concentrated HCl were added to this solution and gently warmed. Further, oxalic acid (225 mg) was added in portions till the colour of the permanganate ion disappeared. The quantity of $MnCl_2$ (in mg) present in the initial solution is ________. (Atomic weights in g mol^{-1}: Mn = 55, Cl = 35.5) *(JEE Advanced 2018)*

Answers

9. (4) **10.** 126.00

NCERT Exemp ar Problems

Multiple Choice Questions (Type-I)

Objective Questions

1. Two students performed the same experiment separately and each one of them recorded two readings of mass which are given below. Correct reading of mass is 3.0 g. On the basis of given data mark the correct option out of the following statements.

Student	Readings (*i*)	(*ii*)
A	3.01	2.99
B	3.05	2.95

(*a*) Results of both the students are neither accurate nor precise.

(*b*) Results of student A are both precise and accurate.

(*c*) Results of student B are neither precise nor accurate.

(*d*) Results of student B are both precise and accurate.

2. A measured temperature on Fahrenheit scale is 200°F. What will this reading be on Celsius scale ?

(*a*) 40°C (*b*) 94°C
(*c*) 93.3°C (*d*) 30°C

3. What will be the molarity of a solution, which contains 5.85g of NaCl(*s*) per 500 mL ?

(*a*) 4 mol L^{-1} (*b*) 20 mol L^{-1}
c) 0.2 mol L^{-1} (*d*) 2 mol L^{-1}

4. If 500 mL of a 5M solution is diluted to 1500 mL, what will be the molarity of the solution obtained ?

(*a*) 1.5 M (*b*) 1.66 M
(*c*) 0.017 M (*d*) 1.59 M

5. The number of atoms present in one mole of an element is equal to Avogadro number. Which of the following element contains the greatest number of atoms ?

(*a*) 4 g He (*b*) 46 g Na
(*c*) 0.40 g Ca (*d*) 12 g He

6. If the concentration of glucose ($C_6H_{12}O_6$) in blood is 0.9 g L^{-1}, what will be the molarity of glucose in blood ?

(*a*) 5 M (*b*) 50 M
(*c*) 0.005 M (*d*) 0.5 M

7. What will be the molality of the solution containing 18.25 g of HCl gas in 500 g of water ?

(*a*) 0.1 m (*b*) 1 M
(*c*) 0.5 m (*d*) 1 m

8. One mole of any substance contains 6.022×10^{23} atoms/molecules. Number of molecules of H_2SO_4 present in 100 mL of 0.02M H_2SO_4 solution is _______.

(*a*) 12.044×10^{20} molecules
(*b*) 6.022×10^{23} molecules
(*c*) 1×10^{23} molecules
(*d*) 12.044×10^{23} molecules

9. What is the mass percent of carbon in carbon dioxide ?

(*a*) 0.034 % (*b*) 27.27 %
c) 3.4 % (*d*) 28.7%

Answers

1 (*b*) **2.** (*c*) **3.** (*c*) **4.** (*b*) **5.** (*d*) **6.** (*c*) **7.** (*d*) **8.** (*a*) **9.** (*b*)

Competition File

10. The empirical formula and molecular mass of a compound are CH_2O and 180 g respectively. What will be the molecular formula of the compound ?

(*a*) $C_9H_{18}O_9$ (*b*) CH_2O
(*c*) $C_6H_{12}O_6$ (*d*) $C_2H_4O_2$

11. If the density of a solution is 3.12 g mL^{-1}, the mass of 1.5 mL solution in significant figures is _________ .

(*a*) 4.7 g (*b*) 4680×10^{-3}g
(*c*) 4.680 g (*d*) 46.80 g.

12. Which of the following statements about a compound is incorrect ?

(*a*) A molecule of a compound has atoms of different elements.
(*b*) A compound cannot be separated into its constituent elements by physical methods of separation.
(*c*) A compound retains the physical properties of its constituent elements.
(*d*) The ratio of atoms of different elements in a compound is fixed.

13. Which of the following statements is correct about the reaction given below?

$$4Fe(s) + 3O_2(g) \longrightarrow 2Fe_2O_3(g)$$

(*a*) Total mass of iron and oxygen in reactants = total mass of iron and oxygen in product, therefore it follows law of conservation of mass.
(*b*) Total mass of reactants = total mass of product; therefore, law of multiple proportions is followed.
(*c*) Amount of Fe_2O_3 can be increased by taking any one of the reactants (iron or oxygen) in excess.
(*d*) Amount of Fe_2O_3 produced will decrease if the amount of any one of the reactants (iron or oxygen) is taken in excess.

14. Which of the following reactions is not correct according to the law of conservation of mass?

(*a*) $2Mg\ (s) + O_2\ (g) \longrightarrow 2MgO\ (s)$
(*b*) $C_3H_8\ (g) + O_2\ (g) \longrightarrow CO_2(g) + H_2O(g)$
(*c*) $P_4(s) + 5O_2(g) \longrightarrow P_4O_{10}\ (s)$
(*d*) $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$

15. Which of the following statements indicates that law of multiple proportions is being followed?

(*a*) Sample of carbon dioxide taken from any source will always have carbon and oxygen in the ratio 1 : 2.
(*b*) Carbon forms two oxides namely CO_2 and CO, where masses of oxygen which combine with fixed mass of carbon are in the simple ratio 2 : 1.
(*c*) When magnesium burns in oxygen, the amount of magnesium taken for the reaction is equal to the amount of magnesium in magnesium oxide formed.
(*d*) At constant temperature and pressure 200 mL of hydrogen will combine with 100 mL oxygen to produce 200 mL of water vapour.

Answers

10. (*c*) **11.** (*a*) **12.** (*c*) **13.** (*a*) **14.** (*b*) **15.** (*b*)

Multiple Choice Questions (Type-II)

16. One mole of oxygen gas at STP is equal to

(*a*) 6.022×10^{23} molecules of oxygen
(*b*) 6.022×10^{23} atoms of oxygen
(*c*) 16 g of oxygen
(*d*) 32 g of oxygen

17. Sulphuric acid reacts with sodium hydroxide as follows :

$$H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$$

When 1L of 0.1M sulphuric acid solution is allowed to react with 1L of 0.1M sodium hydroxide solution, the amount of sodium sulphate formed and its molarity in the solution obtained is

(*a*) 0.1 mol L^{-1} (*b*) 7.10 g
(*c*) 0.025 mol L^{-1} (*d*) 3.55 g

18. Which of the following pairs have the same number of atoms ?

(*a*) 16 g of $O_2(g)$ and 4 g of H_2(g)
(*b*) 16 g of O_2 and 44 g of CO_2
(*c*) 28 g of N_2 and 32 g of O_2
(*d*) 12 g of C(*s*) and 23 g of Na (*s*)

19. Which of the following solutions have the same concentration ?

(*a*) 20 g of NaOH in 200 mL of solution
(*b*) 0.5 mol of KCl in 200 mL of solution
(*c*) 40 g of NaOH in 100 mL of solution
(*d*) 20 g of KOH in 200 mL of solution

Answers

16. (*a, d*) **17.** (*b, c*) **18.** (*c, d*) **19.** (*a, b*)

Competition File

20. 16 g of oxygen has same number of molecules as in

a) 16 g of CO (b) 28 g of N_2

(c) 14 g of N_2 d) 1.0 g of H_2

21. Which of the following terms are unitless ?

(a) Molality (b) Molarity

(c) Mole fraction (d) Mass percentage

22. One of the statements of Dalton's atomic theory is given below :

"Compounds are formed when atoms of different elements combine in a fixed ratio."

Which of the following laws is not related to this statement ?

(a) Law of conservation of mass

b) Law of definite proportions

(c) Law of multiple proportions

(d) Avogadro law

Answers

20. (*c, d*) **21.** (*c, d*) **22.** (*a, d*)

Matching Type Questions

23. Match the following :

(*i*) 88 g of CO_2	(*a*) 0.25 mol
(*ii*) 6.022×10^{23} molecules of H_2O	(*b*) 2 m
(*iii*) 5.6 litres of O_2 at STP	(*c*) 1 mol
(*iv*) 96 g of O_2	(*d*) 6.022×10^{23} molecules
(*v*) 1 mol of any gas	(*e*) 3 mol

24. Match the following physical quantities with units :

Physical quantity	Unit
(*i*) Molarity	(*a*) g mL^{-1}
(*ii*) Mole fraction	(*b*) mol
(*iii*) Mole	(*c*) Pascal
(*iv*) Molality	(*d*) Unitless
(*v*) Pressure	(*e*) mol L^{-1}
(*vi*) Luminous intensity	(*f*) Candela
(*vii*) Density	(*g*) mol kg^{-1}
(*viii*) Mass	(*h*) Nm^{-1}
	(*i*) kg

Answers

23. (*i*) – (*b*); (*ii*) – (*c*); (*iii*) – (*a*); (*iv*) – (*e*); (*v*) – (*d*).

24. (*i*) – (*e*); (*ii*) – (*d*); (*iii*) – (*b*); (*iv*) – (*g*); (*v*) – (*c*); (*vi*) – (*f*); (*vii*) – (*a*); (*viii*) – (*i*).

Assertion and Reason Type Questions

In the following questions a statement of Assertion (A) followed by a statement of Reason (R) is given. Choose the correct option out of the choices given below each question.

25. **ssertion (A) :** The empirical mass of ethene is half of its molecular mass.

Reason (R) : The empirical formula represents the simplest whole number ratio of various atoms present in a compound.

(*a*) Both A and R are true and R is the correct explanation of A.

(*b*) A is true but R is false.

(*c*) A is false but R is true.

(*d*) Both A and R are false.

26. **Assertion (A) :** One atomic mass unit is defined as one twelfth of the mass of one carbon-12 atom.

Reason (R) : Carbon-12 isotope is the most abundant isotope of carbon and has been chosen as standard.

(*a*) Both A and R are true and R is the correct explanation of A.

(*b*) Both A and R are true but R is not the correct explanation of A.

(*c*) A is true but R is false.

(*d*) Both A and R are false.

Answers

25. (*a*) **26.** (*b*)

Competition File

27. Assertion (A) : Significant figures for 0.200 is 3 where as for 200 it is 1.

Reason (R) : Zero at the end or right of a number are significant provided they are not on the right side of the decimal point.

(*a*) Both A and R are true and R is the correct explanation of A.

b) Both A and R are true but R is not a correct explanation of A.

(*c*) A is true but R is false.

(*d*) Both A and R are false.

28. Assertion (A) : Combustion of 16 g of methane gives 18 g of water.

Reason (R) : In the combustion of methane, water is one of the products.

(*a*) Both A and R are true but R is not the correct explanation of A.

b) A is true but R is false.

(*c*) A is false but R is true.

(*d*) Both A and R are false.

Answers

27. (*c*) **28.** (*c*)

Hints & Explanations for Difficult Objective Type Questions

A. mcq with only one correct answer

A4. (*b*) : Ratio of X and Y in compound A is 1:1

Ratio of X and Y in compound C is 1:3

Now, if in compound A, 28 parts by wt.of X combine with 16 parts by wt. of Y, then in compound C, 28 parts of X will combine with $16 \times 3 = 48$ parts of Y.

$\therefore$ 24 parts by wt. of Y will combine with 14 parts of X in compound C.

A5. (*d*) : In first oxide : Metal = 42.9 parts and oxygen = 57.1 parts

Since the formula is MO.

$\therefore$ 42.9 parts of metal = 1 atom

57.1 parts of oxygen = 1 atom

In second oxide, metal = 27.3 parts and oxygen = 72.7 parts

Now, 42.9 parts of metal = 1 atom

27.3 parts of metal $= \frac{1}{42.9} \times 27.3$

= 0.636 atoms and

57.1 parts of oxygen = 1 atom

72.7 parts of oxygen $= \frac{1}{57.1} \times 72.7$

= 1.267 atoms

Ratio of M : O = 0.636 : 1.267

or = 1 : 2

Formula = MO_2.

A6. (*c*) : 4.26 – (15.635/5.0)

4.26 – 3.1 (upto two signicicant figures)

= 1.2 (upto first decimal place after rounding off).

A7. (*d*) : Different weights are (*a*) 88g (*b*) 34.2g. (*c*) 40g (*d*) 27g. 27g (1.5 mol) of water weigh least.

A8. (*d*) : (*a*) $6.022 \times 10^{23} \times 3$ (*b*) $0.5 \times 6.022 \times 10^{23} \times 2$ (*c*) $0.1 \times 6.022 \times 10^{23} \times 3$ (*d*) $1 \times 6.022 \times 10^{23} \times 5$.

A11. (*a*) : No. of atoms of P_4 $= 0.1 \times 6.02 \times 10^{23} \times 4$

$= 2.4 \times 10^{23}$ atoms.

A12. (*a*) : Moles of NH_3 $= \frac{8.5}{17} = 0.5$

No. of H atoms $= 0.5 \times 6.022 \times 10^{23} \times 3$

$= 9.03 \times 10^{23}$ atoms.

A13. (*c*) : No. of molecules of O_3 in 22.4 L $= 6.022 \times 10^{23}$

No. of molecules of O_3 in 11.2 L

$= \frac{1}{2} \times 6.022 \times 10^{23}$

No. of O atoms $= \frac{1}{2} \times 3 \times 6.022 \times 10^{23}$

$= 9.03 \times 10^{23}$.

A14. (*c*) : Mass of 1 He atom = 4 a.m.u.

No. of He atoms $= \frac{104}{4} = 26$

A15. (*c*) : Moles of water $= \frac{0.018}{18} = 1 \times 10^{-3}$

Molecules of water $= 6.022 \times 10^{23} \times 10^{-3}$

$= 6.022 \times 10^{20}$

A16. (c) : No. of O atoms in 34.2 g of $C_{12}H_{22}O_{11}$.

$= \frac{6.02 \times 10^{23} \times 34.2}{342} \times 11$

$= 6.62 \times 10^{23}$

No. of O atoms in 90 g of H_2O

$= \frac{6.02 \times 10^{23} \times 90}{18} \times 1$

$= 3.01 \times 10^{24}$

Competition File

Total number of O atoms

$= 0.662 \times 10^{24} + 3.01 \times 10^{24}$

$= 3.67 \times 10^{24}$.

A21. (*c*) : 22400 mL of water contain molecules

$= 6.022 \times 10^{23}$

1 mL of water contains molecules

$$= \frac{6.022 \times 10^{23}}{22400} = 2.69 \times 10^{19}.$$

A22. (*b*) :

$$2\text{ g O} = \frac{2}{16} = \frac{1}{8}\text{ mol}$$

$$4\text{ g S} = \frac{4}{32} = \frac{1}{8}\text{ mol}$$

These will contain same number of atoms.

A23. (*a*) : Caustic soda is NaOH

40 g of NaOH contain 6.02×10^{23} molecules

$= 6.02 \times 10^{23}$ atoms of H.

A24. (*b*) : $C + O_2 \longrightarrow CO_2$

12 g of C gives CO_2 = 44 g

$$0.6\text{ g of C will give } CO_2 = \frac{44}{12} \times 0.6 = 2.2\text{ g}$$

$$\text{Moles of } CO_2 = \frac{2.2}{44} = 0.05$$

No. of molecules $= 0.05 \times 6.02 \times 10^{23}$

$= 3.01 \times 10^{22}$

A25. (*a*) : 1 molecule of CH_4 contains electrons = 6 + 4 = 10

16 g of CH_4 contain 6.022×10^{23} molecules

$= 6.022 \times 10^{23} \times 10$ electrons

$$\therefore 3.2\text{ g of } CH_4 \text{ contain} = \frac{6.022 \times 10^{23} \times 10 \times 3.2}{16}$$

$= 2 \times 10 \times 6.022 \times 10^{22}$ electrons.

A26. (*d*) :

$$\text{Moles of } NH_3 = \frac{4.25}{17}$$

$$\text{No. of atoms} = \frac{4.25}{17} \times 6.02 \times 10^{23} \times 4$$

$= 6.02 \times 10^{23}$.

A27. (*b*) : Fe present in 67200 u $= \frac{0.33}{100} \times 67200$

$= 221.8$ u

$$\text{No. of iron atoms present in 67200 u} = \frac{221.8}{56} = 4$$

A28. (*d*) : Mass of 1 mol of electrons

$= 9.108 \times 10^{-31} \times 6.02 \times 10^{23}$ kg

$$1\text{ kg of electrons} = \frac{1}{9.108 \times 6.02} \times 10^{8}\text{ mol.}$$

A29. (*d*) : Suppose 1 molecule of alkaloid contains *x* atoms of N

$$\therefore \quad \%\text{ of N} = \frac{14 \times x}{162} \times 100 = 17.28$$

$$\therefore \quad = \frac{17.28 \times 162}{14 \times 100} = 2$$

No. of N atoms in alkaloid = 2

A30. (*a*) : No. of atoms of Fe $= \frac{558.6}{55.86} \times 6.02 \times 10^{23}$

$= 10 \times 6.02 \times 10^{23}$

$$\text{No. of atoms in 60 g of C} = \frac{60}{12} \times 6.02 \times 10^{23}$$

$= 5 \times 6.02 \times 10^{23}$

$\therefore$ No. of atoms in 558.6 g of Fe = twice the no. of atoms in 60 g of C.

A31. (*c*) : Molecular formula of sucrose = $C_{12}H_{22}O_{11}$

Empirical formula = $C_{12}H_{22}O_{11}$.

A32. (*c*) : $Ca_3P_2 + 6H_2O \longrightarrow 2PH_3 + 3Ca(OH)_2$

1 mol of Ca_3P_2 gives 2 moles of PH_3.

A33. (*b*) : Out of 1 mole of water, number of moles of D_2O

$$= \frac{x \times 1}{100} = 0.01\,x$$

Number of moles of D_2O × mass of D_2O + Number of moles of H_2O × mass of H_2O = 19

$0.01x \times 20 + (1 - 0.01x) \times 18 = 19$

$0.01x \times 20 + 18 - 0.01x \times 18 = 19$

$0.2x - 0.18x = 1$

$0.02x = 1$

$x = 1/0.02 = 50$

A34. (*d*) : Let '*a*' atoms of sulphur be present in *x* g of the compound.

8% of weight of compound

= weight of sulphur in compound

$$\frac{8}{100}x = 32 \times a$$

$$x = \frac{3200}{8}a = 400\,a$$

The minimum molecular weight is the one in which *a* = 1 (*i.e.* 1 atom of sulphur is present).

Hence, minimum molecular weight = 400

A35. (*a*) : Normality of oxalic acid $= \frac{6.3}{63} \times \frac{1000}{250} = 0.4$ N

$N_1V_1 \equiv N_2V_2$

$0.4 \times 10 = 0.1 \times V_2$

$$V_2 = \frac{0.4 \times 10}{0.1} = 40\text{ mL.}$$

Competition File

A36. (*a*) : Mixture X contains 0.02 mol Br^- and 0.02 mol SO_4^{2-} in 2 L.

$\therefore$ 1 L of X will contain 0.01 mol Br^- and 0.01 mol SO_4^{2-}

With excess of $AgNO_3$, 0.01 mol of AgBr *i.e.*, Y is formed.

With excess of $BaCl_2$, 0.01 mol of $BaSO_4$ *i.e.*, Z is formed.

A37. (*b*) : $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$

1 L of N_2 reacts with 3 L of H_2 to form 2 L of NH_3. Thus, H_2 is limiting reagent.

10 L of N_2 will react with 30 L of H_2 to form 20 L of NH_3.

Since actual yield is 50% of the expected value,

NH_3 formed = 10 L

N_2 reacted = 5 L

N_2 Unreacted = 30 – 5 = 25 L

H_2 reacted = 15 L

H_2 Unreacted = 30 – 15 = 15 L

$\therefore$ Mixture will contain 10 L NH_3, 25 L N_2, 15 L H_2.

A38. (*d*): Equal volumes contain equal no. of moles

$\therefore$ Molar ratio = 1 : 1

Ratio by weight = 4 : 16 = 1 : 4

$\therefore$ CH_4 present by weight

$$\frac{4}{5} \times 100 = 80\%$$

A39. (*d*): $2\,Al_2O_3 + 3C \longrightarrow 4\,Al + 3CO_2$

4 × 27 kg of Al consumes C of anode

= 3× 12 = 36 kg

Production of 270 kg of Al consumes C of anode

$$= \frac{36}{4 \times 27} \times 270 = 90 \text{ kg}$$

A40. (*d*): Molecular mass of Na_2SO_4 = 2 × 23 + 32 + 4 × 16 = 142

A crystalline salt on becoming anhydrous loses 55.9% by mass.

44.1 g of anhydrous salt contain H_2O = 55.9

142 g of anhydrous salt contain H_2O

$$= \frac{55.9}{44.1} \times 142 = 180 \text{ g}$$

No. of water molecules = $\frac{180}{18}$ = 10 molecules

A41. (*a*) : $N_2 + 3H_2 \longrightarrow 2NH_3$

1 mol of N_2 (28g) combine with 3 mol (6g) of H_2

$\therefore$ 3 kg of N_2 will react with $\frac{28}{3} \times 3$ = 14 kg of N_2

H_2 is the limiting reagent.

6g of H_2 produce NH_3 = 34g

3.0 kg of H_2 will produce NH_3 = $\frac{34}{6} \times 3$ = 17 kg

A42. (*a*):

Element	Percentage	Atomic ratio	Simplest ratio	Whole no. ratio
P	43.6%	$\frac{43.6}{31} = 1.41$	$\frac{1.41}{1.41} = 1$	2
O	56.4%	$\frac{56.4}{16} = 3.52$	$\frac{3.52}{1.41} = 2.5$	5

Formula = P_2O_5

A43. (*c*): 38 g of HCl is present in 100 g of solution

$$\text{Vol. of solution} = \frac{100}{1.19}$$

$$\text{Molarity} = \frac{\frac{38}{36.5} \times 1000}{\frac{100}{1.19}}$$

$$= 12.38 \text{ M}$$

A44. (*b*): $M_1V_1 + M_2V_2 = M(V_1 + V_2)$

0.5 × 800 + 1 × 200 = M (800 + 200)

$$M = \frac{400 + 200}{1000} = 0.6\,M$$

A45. (*b*): Final volume = 2 L + 4 L = 6 L

$$\text{Molarity} = \frac{6 \times 2}{6} = 2 \text{ M}$$

A46. (*d*):

$$M_1V_1 = M_2V_2$$

$$10.50 \times V_1 = 0.25 \times 1.0$$

$$V_1 = \frac{0.25 \times 1.0}{10.50}$$

$$= 0.0238 \text{ L} = 23.8 \text{ mL}$$

A47. (*c*) :

$$\text{Molarity} = \frac{\text{Moles}}{\text{Vol (in mL)}} \times 1000$$

$$0.50 = \frac{\text{Moles}}{250} \times 1000$$

$$\text{Moles of NaCl} = \frac{0.50 \times 250}{1000}$$

$$= 0.125 \text{ mol.}$$

A48. (*d*) :

$C_xH_y(g)$	+	$\left(x+\frac{y}{4}\right)O_2$	$\longrightarrow$	$xCO_2(g)$	+	$\frac{y}{2}H_2O(l)$
1 mL		$\left(x+\frac{y}{4}\right)$mL		x mL		negligible
6 mL		$6\left(x+\frac{y}{4}\right)$mL		$6x$ mL		

Equating theoretical and experimental value of CO_2

$6x = 12$ or $x = 2$

Theoretical value of contraction after explosion and cooling for 1mL of hydrocarbon

$$1+\left(x+\frac{y}{4}\right)-x=1+\frac{y}{4}$$

For 6 mL of hydrocarbon contraction on explosion and cooling = 9 mL

For 1 mL of hydrocarbon contraction on explosion and cooling $=\frac{9}{6}=1.5$ mL

Equating

$$1+\frac{y}{4}=1.5 \quad \text{or} \quad \frac{y}{4}=0.5 \quad \text{or} \quad y=2$$

∴ Formula of hydrocarbon = C_2H_2

B. mcq from competitive examinations

B1. (*b*) : $C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(l)$

22.4 L at N.T.P.   5 × 22.4 L at N.T.P.

22.4 L of C_3H_8 at N.T.P. require $O_2 = 5 \times 22.4$ L

1 L of C_3H_8 at N.T.P. will require O_2

$$=\frac{5\times 22.4}{22.4}\text{ L at N.T.P.}$$

$= 5$ L

B2. (*c*) : $PbO + 2HCl \longrightarrow PbCl_2 + H_2O$

Molar mass of PbO = 207 + 16 = 223 g mol^{-1}

Moles of PbO $=\frac{6.5}{223}=0.029$ mol

Moles of HCl $=\frac{3.2}{36.5}=0.088$ mol

1 mol of PbO reacts with 2 mol of HCl

∴ PbO is a limiting reagent

0.029 mol of PbO will produce $PbCl_2$ = 0.029 mol.

B3. *b*) : Mass of 1 molecule of water $=\frac{18}{6.022\times 10^{23}}$ g

$$\text{Volume}=\frac{18\text{ g}}{6.022\times 10^{23}\times 1\text{g/cm}^3}=3.0\times 10^{-23}\text{ cm}^3$$

B4. (*b*) : Molality and mole fraction are independent of temperature.

B5. (*b*) : $2H_2 + O_2 \longrightarrow 2\,H_2O$

10 g = 5 mol   64 g = 2 mol

Limiting reagent is O_2

1 mole of O_2 gives 2 moles of H_2O

∴ 2 moles of O_2 will give 4 moles of H_2O.

B6. (*c*) Normality = n × Molarity

For acids, n = Basicity

∴ For H_2SO_4, n = 2

0.02 N = 2 × Molarity

or Molarity of 0.02 N H_2SO_4

$$=\frac{0.02}{2}=0.01\text{M}$$

0.01 M solution of H_2SO_4 contains 0.01 moles of H_2SO_4 in 1000 mL solution.

∴ Number of moles in 100 mL solution

$$=\frac{0.01}{1000}\times 100$$

= 0.001 moles

Number of molecules in 0.001 moles

$= 6.022 \times 10^{23} \times 0.001$

$= 6.022 \times 10^{20}$ molecules

B7. (*b*) $4Al + 3O_2 \longrightarrow 2Al_2O_3$

4 × 27 = 54 g   3 mol

Amount of Al that combines with 3 moles of O_2

= 54

∴ Amount of Al that combines with 1.5 moles of O_2 = 27 g

B8. (*d*) $A + 2B \longrightarrow C$

1 Mol   2 Mol   1 mol

8 moles of B require 4 moles of A. Therefore, limiting reagent is B.

2 moles of B give 1 mole of C

∴ 8 moles of B give $=\frac{1}{2}\times 8=4$ moles of C.

B9. (*b*) 1 kg of sea water = 10^6 mg sea water

10^6 mg sea water has 6 mg dissolved O_2

The concentration of $O_2=\frac{6}{10^6}\times 10^6$ ppm

= 6 ppm

B10. (*c*) $$\text{Molarity}=\frac{25.3/106}{250}\times 1000$$

= 0.955 M

$Na_2CO_3 \longrightarrow 2Na^+ + CO_3^{2-}$

∴ $[Na^+] = 2 \times 0.955 = 1.910$ M

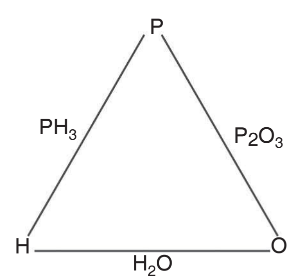
$[CO_3^{2-}] = 0.955$ M

B11. (*c*) No. of atoms in 0.1 mol

$= 6.022 \times 10^{23} \times 0.1 \times 3$

$= 1.806 \times 10^{23}$

B12. (*a*)



Ratio of number of hydrogen and oxygen combining with one P is

3 : 1.5 *i.e.*, 2 : 1

It is same as in water (2 : 1)

Competition File

B13. (*a*) $\underset{28}{N_2} + \underset{2\times 3}{3H_2} \longrightarrow \underset{34}{2NH_3}$

1 mole of N_2 (28 g) combine with 3 moles of H_2 (6 g)

$\therefore$ 3 kg of H_2 can react with $\frac{28}{6} \times 3 = 14$ kg of N_2

H_2 is limiting reagent.

Moles of NH_3 formed $= \frac{34}{6} \times 3 = 17$ kg.

B14. (*b*) Moles of solute = 1 m

Moles of solvent $= \frac{1000}{18} = 55.56$

Mole fraction of solute $= \frac{1}{1+55.56} = \mathbf{0.0177}$.

B15. (*b*) $\underset{\substack{1\text{ mol}\\100g}}{CaCO_3} + H_2SO_4 \longrightarrow CaSO_4 + H_2O + \underset{\substack{1\text{ mol}\\22.4L}}{CO_2}$

100g of $CaCO_3$ liberate CO_2 = 22.4L

10g of $CaCO_3$ will liberate $CO_2 = \frac{22.4}{100} \times 10 = 2.24L$

B16. (*c*) (*a*) 0.2 mol of H_2 gas = 0.2 × 2 = 0.4 g

(*b*) 6.023×10^{23} molecules of N_2 = 28g

6.023×10^{22} molecules of N_2

$$= \frac{28}{6.023\times10^{23}} \times 6.023\times10^{22} = 2.8g$$

(*c*) Weight of silver = 0. 1g

(*d*) 0.1 mol of O_2 gas = 0.1 × 32 = 3.2 g

(*e*) Weight of water = 1g

$\therefore$ 0.1 g of silver (*c*) is the lightest.

B17. (*a*) 22.4 L of H_2 at S.T.P. = 1 mol

11.2 L of Cl_2 at S.T.P. = 0.5 mol

So, Cl_2 is limiting reagent.

$H_2 + Cl_2 \longrightarrow 2HCl$

1 mol of Cl_2 reacts with 1 mol of H_2 to give 2 mol of HCl

$\therefore$ 0.5 mol of Cl_2 will react with 0.5 mol of Cl_2 to give 1 mol of HCl.

B18. (*a*) Moles of Mg $= \frac{1}{24} = 0.0417$ mol

Moles of $O_2 = \frac{0.56}{32} = 0.0175$ mol

$\underset{1\text{ mol}}{Mg} + \underset{0.5\text{ mol}}{\frac{1}{2}O_2} \longrightarrow \underset{1\text{ mol}}{MgO}$

O_2 is limiting reagent

0.5 mol of O_2 react with Mg to form 1 mol of MgO

0.0175 mol of O_2 react with Mg $= \frac{1\times0.0175}{0.5}$

$= 0.035$

Moles of Mg left unreacted = 0.0417 – 0.035

$= 6.7 \times 10^{-3}$ mol

Amount of Mg left unreacted $= 6.7 \times 10^{-3} \times 24$

$= 0.1608$ g

B19. (*a*) $CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$

According to equation, 1 mol of $CaCO_3$ reacts with 2 mol of HCl

Moles of HCl $= \frac{20}{1000} \times 1.0 = 0.02$ mol

Moles of $CaCO_3$ required $= \frac{0.02}{2} = 0.01$ mol

Mass of $CaCO_3$ required = 0.01 × 100 = 1 g

B20. (*e*) (*a*) 16 g $O_2 = \frac{16}{32} = 0.5$ mol $= 0.5 \times N_A$ molecules

(*b*) 16 g $NO_2 = \frac{16}{46} = 0.35$ mol $= 0.35 \times N_A$ molecules

(*c*) 4 g $N_2 = \frac{4}{28} = 0.14$ mol $= 0.14 \times N_A$ molecules

(*d*) 2 g $H_2 = \frac{2}{2} = 1$ mol $= N_A$ molecules

(*e*) 32 g $N_2 = \frac{32}{28} = 1.14$ mol $= 1.14 \times N_A$ molecules

$\therefore$ (*e*) 32 of N_2 has maximum number of molecules

B21. (*d*) Let masses of H_2 and O_2 be x and $4x$ g

Moles of $H_2 = \frac{x}{2}$

Mole of $O_2 = \frac{4x}{32}$

Mole ratio $= \frac{x/2}{4x/32} = \frac{4}{1}$

B22. (*d*) Mass of 1 mol (6.022×10^{23} atoms of carbon = 12 g)

If Avogadro number is changed to 6.022×10^{20} atoms, then mass of 1 mol of carbon

$$= \frac{12 \times 6.022\times10^{20}}{6.022\times10^{23}}$$

$= 12 \times 10^{-3}$ g

B23. (*b*) (*a*) 18 g of water $= \frac{18}{18} = 1$ mol

$= N_A$ molecules

(*b*) 18 mol of water $= 18 \times N_A$ molecules

(*c*) 18 molecules of water

(*d*) 1.8 g of water $= \frac{1.8}{18} = 0.1 \times N_A$ molecules

$\therefore$ 18 mol of water has maximum number of water molecules.

B24. (*b*) $\underset{\substack{1\text{ mol}\\84\text{ g}}}{MgCO_3\ (s)} \longrightarrow \underset{\substack{1\text{ mol}\\40\text{ g}}}{MgO\ (s)} + CO_2\ (g)$

40 g of MgO are produced from $MgCO_3$ = 84 g

8 g of MgO are produced from $MgCO_3 = \frac{84}{40} \times 8$

$= 16.8$ g

Competition File

$$\% \text{ Purity of } MgCO_3 = \frac{16.8}{20} \times 100$$
$$= 84\ \%$$

B25. (*b*) 1 m aqueous solution means 1 mol of solute is present in 1000 g of water

$\therefore$ Moles of solute = 1 mole

$$\text{Moles of water} = \frac{1000}{18} = 55.55 \text{ mol}$$

$$\text{Mole fraction of solute} = \frac{1}{1+55.55} = 0.0177$$

B26. (*a*) 16.9 % $AgNO_3$ solution means that 16.9 g of $AgNO_3$ is present in 100 mL of solution.

$$50 \text{ mL of solution contains } AgNO_3 = \frac{16.9}{100} \times 50$$
$$= 8.45 \text{ g}$$

or $$\frac{8.45}{170} = 0.049 \text{ mol}$$

5.8 % NaCl solution means that 5.8 g of NaCl is present in 100 mL of solution.

$$50 \text{ mL of solution contains NaCl} = \frac{5.8}{100} \times 50 = 2.9 \text{ g}$$

or $$\frac{2.9}{58.5} = 0.049 \text{ mol}$$

$$AgNO_3 + NaCl \longrightarrow AgCl + NaNO_3$$

1 mol of $AgNO_3$ combines with 1 mol of NaCl to give 1 mol of AgCl

0.049 mol of $AgNO_3$ combines with 0.049 mol of NaCl to give AgCl

$= 0.049$ mol
$= 0.049 \times 143.5$
$= 7$ g

B27. (*a*) : Let atomic weight of element X is x and that of element Y is y.

For XY_2

0.1 mol of XY_2 = 10 g

$$1.0 \text{ mol of } XY_2 = \frac{10}{0.1} \times 1.0 = 100 \text{ g}$$

or $x + 2y = 100$ g mol^{-1} .....(*i*)

For X_3Y_2

0.05 mol of X_3Y_2 = 9 g

$$1 \text{ mol } X_3Y_2 = \frac{9}{0.05} = 180 \text{ g}$$

or $3x + 2y = 180$ g mol^{-1} .....(*ii*)

Subtracting eq. (*i*) from eq. (*ii*)

$2x = 80$ g mol^{-1} or $x = 40$ g mol^{-1}

and $40 + 2y = 100$

or $$y = \frac{100-40}{2} = 30 \text{ g mol}^{-1}$$

B28. (*b*) : Molarity depends upon temperature.

B29. (*a*) : 18 mL of water has maximum number of molecules.

(*a*) : 18 mL of water = 18 × 1 g/mL = 18 g

$$\text{Molecules of water} = \frac{18}{18} \times N_A = N_A$$

(*b*) 0.18 g of water

$$\text{Molecules of water} = \frac{0.18}{18} \times N_A = 0.01 \times N_A$$

(*c*) $$\text{Moles of water} = \frac{0.00224 \text{ L}}{22.4} = 0.0001$$

Molecules of water = 0.0001 × N_A = 0.0001 × N_A

(*d*) 10^{-3} mol of water

Molecules of water = $10^{-3} \times N_A = 0.001\ N_A$

B30. (*a*) : If atomic mass unit on the scale of $\frac{1}{6}$ of C – 12 is 2 a.m.u. on the scale of $\frac{1}{12}$ of C–12, the mass of one mole of substance will become half of the normal value.

B31. (*c*) : For each mole of Al reacted 1.5 mole of H_2 (g) (or 33.6 L at STP) is formed and for each mole of HCl consumed 0.5 mole of H_2(g) (or 11.2 L at STP) is formed.

B32. (*b*) :

$$\text{Moles of oxygen} = \frac{80}{16} = 5 \text{ mol}$$

No. of atoms of oxygen = $5 \times 6.022 \times 10^{23}$

$$\text{Moles of hydrogen in 5g} = \frac{5}{1} = 5 \text{ mol}$$

No. of atoms of hydrogen = $5 \times 6.022 \times 10^{23}$

B34. (*a*) :

$$N_1V_1 + N_2V_2 = NV$$

where $_1 + V_2 = 1L$

or $V_2 = (1 - V_1)L$

$10 \times V_1 + 4\,(1 - V_1) = 1 \times 7$

$10V_1 + 4 - 4V_1 = 7$

$6V_1 = 3$

$$V_1 = \frac{3}{6} = 0.5\,L$$

$V_2 = 1 - 0.5 = 0.5$ L

B35. (*b*) : 98% H_2SO_4 means 98 g H_2SO_4 is present in 100 g of solution.

Volume of 100 g solution

$$= \frac{100}{1.8} = 55.55 \text{ mL}$$

$$\text{Molarity} = \frac{\text{Moles of } H_2SO_4 \times 1000}{\text{Volume of solution}}$$

$$= \frac{98}{98 \times 55.55} \times 1000 = 18M$$

B36. (*a*) :

$$N_1V_1 + N_2V_2 = NV$$

where $V_1 + V_2 = 1L = 1000$ mL

or $V_2 = (1 - V_1)L$

$10 \times 20 + 36 \times 10 = N \times 1000$

$200 + 360 = 1000$ N

or $$N = \frac{560}{1000} = 0.56\,N$$

Competition File

B37. (*b*) : $Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$

Moles of $Ca(OH)_2$ present in 50 mL of 0.5 M solution

$$= \frac{0.5}{1000} \times 50 = 2.5 \times 10^{-2} \text{ mol}$$

Moles of $CaCO_3$ formed = 2.5×10^{-2} mol

$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$

Moles of HCl required

$$= 2 \times 2.5 \times 10^{-2} \text{ mol}$$
$$= 5.0 \times 10^{-2} \text{ mol}$$

Volume of 0.1 N HCl required

$$= \frac{0.1}{1000} \times x = 5.0 \times 10^{-2}$$

$$x = \frac{5.0 \times 10^{-2} \times 1000}{0.1} = 500 \text{ mL}$$

B38. (*b*) Time taken to distribute 10^{20} grains

$$= 1 \text{ second}$$

Time taken to distribute 6.022×10^{23} grains

$$= \frac{1}{10^{20}} \times 6.022 \times 10^{23} \text{ seconds}$$

$$= \frac{6.022 \times 10^{23}}{10^{20} \times 60 \times 60} \text{ hours}$$

$$= 1.673 \text{ hours}$$

B39. (*b*) Let atomic mass of metal = x, then

$$\% \text{ of oxygen in } M_2O = \frac{16}{2x + 16} \times 100 = 36.4$$

or $\quad 1600 = 72.8x + 582.4$

$$x = \frac{1600 - 582.4}{72.8} = 13.978$$

Now, in second oxide, oxygen and metal are 53.4% and 46.6% respectively.

∴ Atomic ratio = M : O

$$\frac{46.6}{13.978} : \frac{53.4}{16}$$

$$3.3 : 3.3$$

or $\quad 1 : 1$

∴ Formula of metal oxide = MO

B40. (*a*) I. Mass of 1 atom of oxygen $= \dfrac{16}{6.022 \times 10^{23}}$

$$= 2.66 \times 10^{-23} \text{ g.}$$

II. Mass of 1 atom of nitrogen $= \dfrac{14}{6.022 \times 10^{23}}$

$$= 2.32 \times 10^{-23} \text{ g.}$$

III. Mass of 1×10^{-10} mol of oxygen = 16×10^{-10} g

IV. Mass of 1×10^{-10} mol of copper = 63×10^{-10} g

∴ Order of increasing atomic mass :

II < I < III < IV.

B41. (*c*) $C_2H_6 + \frac{7}{2}O_2 \longrightarrow 2CO_2 + 3H_2O$

$C_2H_4 + 3O_2 \longrightarrow 2CO_2 + 2H_2O$

Moles of mixture, $n = \dfrac{pV}{RT}$

$$= \frac{1 \text{ atm} \times 41 \text{ L}}{(0.082 \text{ L atm K}^{-1}\text{mol}^{-1}) \times 500\text{K}}$$

$$= 1 \text{ mol}$$

Let moles of $C_2H_6 = x$, moles of $C_2H_4 = 1 - x$

Oxygen required

$$\frac{7}{2}x + (1 - x)3 = \frac{10}{3}$$

$$21x + 18 - 18x = 20$$

$$3x = 2 \quad \text{or} \quad x = 0.67 \text{ mol}$$

Moles of ethane = 0.67,

Moles of ethene = 1 – 0.67 = 0.33

Mole fraction of ethane $= \dfrac{0.67}{1} = 0.67$,

Mole fraction of ethene $= \dfrac{0.33}{1} = 0.33$

B42. (*b*) $NaCl + CaCl_2 \xrightarrow{Na_2CO_3} CaCO_3$

Let wt. of $CaCl_2$ in mixture = x

and wt. of NaCl in mixture = $4.44 - x$

$$CaCO_3 \longrightarrow CaO + CO_2$$

Moles of Ca in $CaCl_2$ = Moles of Ca in $CaCO_3$

Moles of Ca in $CaCO_3$ = Moles of Ca in CaO

$$\frac{x}{111} = \frac{0.56}{56}$$

$$\therefore \quad x = \frac{0.56}{56} \times 111 = 1.11 \text{ g}$$

Wt. of NaCl in mixture = 4.44 – 1.11 = 3.33 g

$$\% \text{ of NaCl} = \frac{3.33}{4.44} \times 100 = 75\%$$

B43. (*a*) Equivalents of HCl $= \dfrac{0.2 \times 50}{1000} = 10 \times 10^{-3}$

Equivalents of NaOH $= \dfrac{0.1 \times 50}{1000} = 5.0 \times 10^{-3}$

5.0×10^{-3} equivalents of NaOH will neutralise 5.0×10^{-3} equivalents of HCl

Equivalent of HCl left = $(10 - 5) \times 10^{-3} = 5 \times 10^{-3}$

Volume of solution = 50 + 50 = 100 cm^3

$$\text{Normality} = \frac{5 \times 10^{-3}}{100} \times 1000 = 0.05 \text{ N}$$

100 cm^3 of 0.05 N HCl is titrated against 0.5 N KOH.

$$\underbrace{N_1 \times V_1}_{HCl} = \underbrace{N_2 \times V_2}_{KOH}$$

$$0.05 \times 100 = 0.5 \times V_2$$

$$\therefore \quad V_2 = \frac{0.05 \times 100}{0.5} = \mathbf{10\ cm^3}.$$

Competition File

B44. (*b*) $MCO_3 \xrightarrow{\Delta} MO + CO_2$

1 mol — 22400 cc

2g — 448 cc

448 cc of CO_2 is evolved from MCO_3 = 2g

22400 cc of CO_2 will be evolved from

$$MCO_3 = \frac{2}{448} \times 22400 = 100 \text{ g}$$

$\therefore$ 1 mol of MCO_3 = 100 g

or Molecular mass of MCO_3 = 100 g mol^{-1}

If M is the atomic mass of metal,

$$M + 12 + 3 \times 16 = 100$$

$$\therefore \quad M = 100 - 12 - 48 = 40$$

$$\text{Equivalent wt.} = \frac{\text{Atomic mass}}{\text{Valency}} = \frac{40}{2} = 20$$

B45. (*c*) Mass of water = 18 mL × 1g mL^{-1} = 18 g

or = 1 mol

No. of molecules in 1 mol of water

$= 6.02 \times 10^{23}$

No. of electrons in 1 molecule of water

$= 2 + 8 = 10$

$\therefore$ No. of electrons in 1 mol of water

$= 6.02 \times 10^{23} \times 10$

$= 6.02 \times 10^{24}$

B46. (*a*) Let volume of solution A = x L

Volume of solution B = $(2 - x)$ L

$$M_1 V_1 + M_2V_2 = M(V_1 + V_2)$$

$$0.5 \times x + 0.1\,(2 - x) = 0.2\,(2)$$

$$0.5x + 0.2 - 0.1x = 0.4$$

$$0.4x = 0.2$$

$$x = \frac{0.2}{0.4} = 0.5\text{L}$$

Vol. of A = 0.5 L, Vol. of B = 2.0 – 0.5 = 1.5 L

B47. (*d*) Moles of water in a drop of water = $\frac{0.018}{18} = 1 \times 10^{-3}$

1 mole of water = 6.022×10^{23} molecules

1×10^{-3} mol of water = $6.022 \times 10^{23} \times 1 \times 10^{-3}$

$= 6.022 \times 10^{20}$ molecules.

B48. (*c*) Molarity = $\frac{\text{Moles of } HNO_3}{\text{Volume}} \times 1000$

or Moles of HNO_3 = $\frac{\text{Molarity} \times \text{Volume}}{1000}$

$$= \frac{2.0 \times 250}{1000} = 0.5$$

Mass of HNO_3 = $0.5 \times 63 = 31.5$ g.

Mass of 70% HNO_3 = $\frac{31.5 \times 100}{70} = 45.0$g

B49. (*b*) $M_1V_1 + M_2V_2 = M_{mix}(V_1 + V_2)$

$$0.5 \times 750 + 2 \times 250 = M_{mix}(1000)$$

$$M_{mix} = \frac{0.5 \times 750 + 2 \times 250}{1000}$$

$$= \frac{375 + 500}{1000} = 0.875 \text{ M}$$

B50. (*a*) Moles of CO_2 produced = $\frac{3.08}{44} = 0.07$

Moles of H_2O produced = $\frac{0.72}{18} = 0.04$

Combustion of hydrocarbon C_xH_y may be represented as :

$$C_x H_y + \left(x + \frac{y}{4}\right) O_2 \longrightarrow x\,CO_2 + \frac{y}{2} H_2O$$

$$x = 0.07, \; \frac{y}{2} = 0.04$$

$$\therefore \quad y = 0.08$$

$$\frac{x}{y} = \frac{0.07}{0.08} = \frac{7}{8}$$

$\therefore$ Empirical formula of hydrocarbon C_7H_8.

B51. (*a*) $BaO + 2HCl \longrightarrow BaCl_2 + H_2O$

$CaO + 2HCl \longrightarrow CaCl_2 + H_2O$

The reaction shows that 2 moles of mixture of BaO and CaO require 4 moles of HCl to react completely.

Moles of HCl in 100 cm^3 = $\frac{2.5 \times 100}{1000}$ = 0.25 mol

4 moles of HCl react with 2 moles of mixture

0.25 mol of HCl reacts with mixture = $\frac{2 \times 0.25}{4}$

= 0.125 mol

Let mass of CaO be x g so that mass of BaO

$= (10 - x)$ g

$$\frac{x}{56} + \frac{10 - x}{153} = 0.125$$

$$153x + 560 - 56x = 0.125 \times 56 \times 153 = 1071$$

or $\quad 97x = 1071 - 560 = 511$

$$\therefore \quad x = \frac{511}{97} = 5.268$$

$$\therefore \quad \%\ CaO = \frac{5.268}{10} \times 100$$

$= 52.68\%$

B52. (*d*)

$$\begin{array}{c} COOH \\ | \\ COOH \end{array} + 2NaOH \longrightarrow \begin{array}{c} COONa \\ | \\ COONa \end{array} + 2H_2O$$

Moles of 0.064 g of NaOH = $\frac{0.064}{40} = 0.0016$

Competition File

Since 2 moles of NaOH react with 1 mol of oxalic acid,

Moles of oxalic acid = $\frac{0.0016}{2} = 8 \times 10^{-4}$

Volume of solution = 25 mL

Molarity of oxalic acid solution = $\frac{8\times10^{-4}}{25} \times 1000$

= 0.032 M

B53. (*c*) 143.5 g of AgCl contain Ag = 108 g

Amount of Ag in 7.20 g of AgCl = $\frac{108}{143.5}\times 7.20$

= 5.42 g

% of Ag = $\frac{5.42}{5.82} \times 100 = 93.1\%$

B54. (*b*)

$$M_1V_1 + M_2V_2 \equiv M_3V_3$$
$$3.0 \times 25 + 4.0 \times 75 \equiv M_3 \times (25 + 75)$$
$$\therefore \quad 75 + 300 = M_3 \times 100$$
$$\therefore \quad M_3 = \frac{375}{100} = 3.75 \text{ M}$$

B55. (*c*) Ratio of masses of O_2 and N_2 =1 : 4

$\therefore$ Ratio of moles of O_2 and N_2 = $\frac{1}{32} : \frac{4}{28}$

= 7 : 32

$\therefore$ Ratio of molecules of O_2 and N_2 = 7 : 32

B56. (*c*) 1000 mL of 0.001 M HCl contain Cl^- ions

= 0.001 mol

100 mL of 0.001 M HCl will contain Cl^- ions

= $\frac{0.001}{1000}\times 100$

= 10^{-4} mol

No. of Cl^- ions = $10^{-4} \times 6.022 \times 10^{23} = 6.022 \times 10^{19}$

B57. (*b*) Percentage of C = $\frac{12}{44}\times\frac{0.44}{0.30}\times 100 = 40\%$

Percentage of H = $\frac{2}{18}\times\frac{0.18}{0.30}\times 100 = 6.67\%$

Percentage of O = 100 – (40 + 6.67) = 53.33%

Calculation of empirical formula

Element	Percentage composition	At. mass	Moles of atoms	Mole ratio or atomic ratio
C	40	12	40/12 = 3.33	1
H	6.67	1	6.67/1 = 6.67	2
O	53.33	16	53.33/16 = 3.33	1

Empirical formula = CH_2O

Empirical formula mass = 12 + 2 × 1 + 16 = 30

Molecular mass = 60

$\therefore \quad n = \frac{60}{30} = 2$

$\therefore$ Molecular formula = $(CH_2O)_2$ = $\mathbf{C_2H_4O_2}$

B58. (*a*) : 6.022×10^{23} molecules of oxygen = 32 g

1.8×10^{22} molecules of oxygen = $\frac{32}{6.022\times10^{23}} \times 1.8 \times 10^{22}$

= 0.956 g ≈ **0.96 g**

B59. (*b*) : $\underset{\text{Propene}}{CH_3—CH{=}CH_2} + \frac{9}{2} O_2 \longrightarrow 3CO_2 + 3H_2O$

3 mol (or 54g) of water are obtained from 1 mol (42g) of propene

or 54g of water is obtained from propene = 42 g

27g of water is obtained from propene

= $\frac{42}{54} \times 27 = 21$ g

B60. (*e*) : 2.46g of hydrated salt ($MSO_4.xH_2O$) gives anhydrous salt = 1.20 g

$\therefore$ Mass of water = 2.46 – 1.20 = 1.26 g

Percentage of water = $\frac{1.26}{2.46} \times 100 = 51.2\%$

Percentage of anhydrous salt = 100 – 51.2 = 48.8%

48.8 g of anhydrous salt contain water = 51.2 g

120 g (1 mol) of anhydrous salt contain water

= $\frac{51.2}{48.8} \times 120 = 126$ g

No. of water molecules = $\frac{126}{18}$ = 7 molecules

B61. (*b*) : Mass of CCl_4 = Volume × Density

= 500 × 1.60 = 800 g

Molar mass of naphthalene ($C_{10}H_8$)

= 12 × 10 + 1 × 8 = 128g mol^{-1}

Moles of naphthalene = $\frac{51.2}{128}$

Molality = $\frac{51.2 \times 1000}{128 \times 800}$ = 0.500 m

B62. (*a*) : Moles of CO_2 = $\frac{4.4}{44}$ = 0.1 mol

1 mol of CO_2 = 6.022×10^{23} molecules of CO_2

or = $2 \times 6.022 \times 10^{23}$ atoms of oxygen

$\therefore$ 0.1 mol of CO_2 = $2 \times 6.022 \times 10^{23} \times 0.1$

= 1.2044×10^{23}

Competition File

B63. (*c*) :

Element	Percentage	Atomic mass	Moles	Simplest ratio
C	40.0	12	$\frac{40}{12} = 3.33$	3.33/3.33=1
H	13.33	1	$\frac{13.3}{1} = 13.33$	13.33/3.33=4
N	46.67	14	$\frac{46.67}{14} = 3.33$	3.33/3.33=1

Empirical formula = CH_4N

B64. (*b*) : $C_xH_y + \left(x+\frac{y}{4}\right)O_2 \longrightarrow xCO_2 + \frac{y}{2}H_2O$

15mL $\quad 15\left(x+\frac{y}{4}\right)$

0 $\quad$ 0 $\quad$ 15*x* mL

Volume of oxygen used,

$$Vo_2 = \frac{20\times 375}{100} = 75\text{mL}$$

$$15\left(x+\frac{y}{4}\right) = 75$$

$$\left(x+\frac{y}{4}\right) = 5$$

This corresponds to formula C_3H_8.
It may be noted that in this case the further information i.e, 330 mL volume is neglected. If we use that information then none of the answer is correct.

B65. (*c*) : Maximum volume of 3M HCl solution can be prepared by taking 500 mL of 2N HCl and *x* mL of 5 N HCl so that total volume becomes (500 + *x*) mL

For HCl, molarity = normality

Applying $N_1V_1 + N_2V_2 \equiv N_3V_3$

$$500 \times 2 + x \times 5 \equiv (500 + x) \times 3$$

$$1000 + 5x = 1500 + 3x$$

$$x = 500 \text{ or } x = 250 \text{ mL}$$

Thus, maximum volume of 3 M solution formed

$$= 500 + 250 = 750 \text{ mL}$$

B66. (*d*) : Equivalents of acid = Equivalents of base

1000 mL of base = 0.1 equivalent

$$20 \text{ mL of base} = \frac{0.1}{1000} \times 20 = 2 \times 10^{-3}$$

Equivalents of acid = 2×10^{-3}

$$\text{Gram equivalents} = \frac{\text{Mass}}{\text{Equivalent wt}}$$

$$2 \times 10^{-3} = \frac{0.126}{\text{Equivalent wt}}$$

$$\therefore \quad \text{Equivalent wt} = \frac{0.126}{2\times 10^{-3}} = 63$$

B67. (*c*) : Let mass of $CH_4(g)$ = 1 g

$$\text{Moles of } CH_4(g) = \frac{1}{16}$$

$$\text{Molecules of } CH_4(g) = \frac{1}{16} \times 6.022 \times 10^{23}$$

Mass of SO_2 (g) = 2

$$\text{Moles of } SO_2(g) \quad \frac{2}{64}$$

$$\text{Molecules of } SO_2 = \frac{2}{64} \times 6.022 \times 10^{23}$$

Ratio of number of molecules of $SO_2(g)$ and $CH_4(g)$

$$\frac{2}{64} \times 6.022 \times 10^{23} : \frac{1}{16} \times 6.022 \times 10^{23}$$

$$\frac{1}{32} : \frac{1}{16} \text{ or } 1 : 2$$

B68. (*b*) : Equivalents of HCl $= \frac{y}{1000} \times x = xy \times 10^{-3}$

Equivalents of NaOH $= \frac{x}{1000} \times y = xy \times 10^{-3}$

	HCl	+ NaOH	⟶ NaCl	+ H_2O
Equivalents added	$xy \times 10^{-3}$	$xy \times 10^{-3}$	0	0
Equivalents left	0	0	$xy \times 10^{-3}$	$xy \times 10^{-3}$

Equivalents of NaCl present in $(x + y)$ mL solution $= xy \times 10^{-3}$

Volume of water added = $(x + y)$ mL

Total volume of solution = $(x + y) + (x + y)$ mL

$$\text{Normality} = \frac{xy\times 10^{-3}}{(x+y)+(x+y)} \times 1000$$

$$= \frac{xy}{2(x+y)} \text{N}$$

B69. (*a*) : No. moles of H_2SO_4 in 98 mg $= \frac{98\times 10^{-3}}{98} = 1 \times 10^{-3}$

No. of moles of 3.01×10^{20} molecules of

$$H_2SO_4 = \frac{3.01\times 10^{20}}{6.022\times 10^{23}}$$

$$= 0.5 \times 10^{-3}$$

Moles of H_2SO_4 left $= 1 \times 10^{-3} - 0.5 \times 10^{-3} = 0.5 \times 10^{-3}$

B70. (*e*) : $MgCO_3 \xrightarrow{\text{Heat}} MgO + CO_2$

10g $\quad$ 4g $\quad$ 0.1 mol

Molar mass of $MgCO_3$ = 24 + 12 + 3 × 16

= 84 g mol^{-1}

Molar mass of MgO = 24 + 16 = 40 g mol^{-1}

40 g of MgO is obtained from 84 g of $MgCO_3$

$$\text{4 g of MgO will be obtained from} = \frac{84}{40} \times 4$$

$$= 8.4 \text{ g of } MgCO_3$$

$$\therefore \text{ Percentage purity} = \frac{8.4}{10} \times 100 = 84\%$$

B71. (*c*) : Molar mass of Na_2CO_3 = 2 × 23 + 12 + 3 × 16

= 106 g mol^{-1}

$Na_2CO_3 . x\, H_2O$ has 50% H_2O by mass

$$(106 + 18x) \times \frac{50}{100} = 18x$$

$$53 + 9x = 18x$$

$$9x = 53 \text{ or } x = 5.88 \approx 6$$

Competition File

B72. (*b*) : $M_2CO_3 + 2HCl \longrightarrow 2MCl + H_2O + CO_2$

1g   0.01186 mol

No. of moles of M_2CO_3 = No. of moles of CO_2

$$\frac{1}{M} = 0.01186$$

$$M = \frac{1}{0.01186} = 84.3$$

Molar mass of M_2CO_3 = 84.3 $gmol^{-1}$

B73. (*a*) : Mass of $Co(NO_3)_2{\cdot}6H_2O$ = 30g

Molar mass = 59 + 2 × 14 + 12 × 16 + 12 × 1 = 291

$$\text{Moles of } Co(NO_3)_2{\cdot}6H_2O = \frac{30}{291}$$

$$\text{Molarity} = \frac{30}{291 \times 4.3} = 0.024 \text{ M}$$

B74. (*d*) : Mass of an electron = 9.108×10^{-31} kg

Mass of 1 mol of electrons = $9.108 \times 10^{-31} \times 6.023 \times 10^{23} = 9.108 \times 6.023 \times 10^{-8}$ kg

$$\text{No. of moles in 1 kg} = \frac{1}{9.108 \times 10^{-8} \times 6.023 \times 10^{23}}$$

$$= \frac{1}{9.108 \times 6.023 \times 10^{23}} \times 10^{8}$$

B75. (*b*) : According to Dulong-Petit's law

Approx. at. mass × sp. heat = 6.4

$$\therefore \text{ At. mass (approx)} = \frac{6.4}{0.16} = 40$$

Let formula of metal chloride = MCl_x

$$\% \text{ Chlorine} = \frac{35.5 \times x \times 100}{40 + 35.5x}$$

or $$\frac{35.5x \times 100}{40 + 35.5x} = 65$$

$$\frac{35.5x}{40 + 35.5x} = \frac{65}{100}$$

or $$\frac{40}{35.5x} + 1 = \frac{100}{65}$$

$$\frac{40}{35.5x} = \frac{100}{65} - 1 = \frac{35}{65}$$

$$\frac{40}{35.5x} = \frac{7}{13}$$

$$35.5x = \frac{40 \times 13}{7}$$

$$x = \frac{40 \times 13}{7 \times 35.5} = 2.09 \simeq 2$$

Formula = MCl_2

B76. (*b*) : $2Mg + O_2 \longrightarrow 2\,MgO$

$$\text{Moles of Mg} = \frac{1.0}{24} = 0.042$$

$$\text{Moles of } O_2 = \frac{0.28}{32} = 0.00875$$

1 mol of O_2 requires Mg = 2 mol

0.00875 mol of O_2 requires Mg = 2 × 0.00875 = 0.0175

∴ Mg is in excess.

Moles of Mg left in excess = 0.042 – 0.0175 = 0.0245

Mass of Mg left in excess = 0.0245 × 24 = 0.58 g

B77. (*a*) : $2\overset{+2}{Fe}SO_4 \longrightarrow 2\overset{+3}{Fe_2}(SO_4)_3$

Change in oxidation state per Fe atom is 1.

∴ Equivalent wt. = Atomic wt. = 55.84 g mol^{-1}

B78. (*a*) : Ethanol : C_2H_5OH

Molecular mass = 2 × 12 + 6 × 1 + 16 × 1 = 46

$$\text{Mass \% C} = \frac{\text{Mass of C}}{\text{Molecular mass}} \times 100$$

$$= \frac{24}{46} \times 100 = 52\%$$

B79. (*d*) :

Element	Relative mass	Relative mole	Simplest whole no. ratio
C	6	6/12 = 0.5	1
H	1	1/1 = 1	2

∴ $x = 1$ and $y = 2$

Combustion of C_xH_y

$$C_xH_y + \left(x + \frac{y}{4}\right)O_2 \longrightarrow xCO_2 + \left(\frac{y}{2}\right)H_2O$$

Oxygen atoms required for complete combustion of C_xH_y

$$= 2\left(x + \frac{y}{4}\right)$$

$$z = \frac{1}{2}\left[2\left(x + \frac{y}{4}\right)\right] = x + \frac{y}{4}$$

$$z = 1 + \frac{2}{4} = \frac{3}{2}$$

$$\text{Ratio of } x : y : z = 1 : 2 : \frac{3}{2}$$

$$= 2 : 4 : 3$$

Formula : $C_2H_4O_3$

B80. (*a*) :

$$\text{Moles of NaOH} = \frac{8}{40} = 0.2 \text{ mol}$$

$$\text{Moles of water} = \frac{18}{18} = 1.0 \text{ mol}$$

$$\text{Mole fraction of NaOH} = \frac{0.2}{0.2 + 1.0} = 0.167$$

$$\text{Molality} = \frac{0.2 \times 1000}{18} = 11.11 \text{ m}$$

B81. (*a*) :

$$\text{Molarity} = \frac{\text{Moles of sugar}}{\text{Vol. of solution in L}}$$

$$0.1 = \frac{\text{Moles of sugar}}{2 \text{ L}}$$

Moles of sugar = 0.1 × 2 = 0.2 mole

Mass of sugar = 0.2 × 342 = 68.4 g

B82. (*c*) :

$$\text{Moles of } Na^+ = \frac{92}{23} = 4$$

$$\text{Molality} = \frac{4}{1} = 4 \text{ m}$$

B83. (*b*) : Atomic mass of iron

$$= \frac{54 \times 5 + 56 \times 90 + 57 \times 5}{100}$$

$$= \frac{270 + 5040 + 285}{100} = 55.95$$

Competition File

C. mcq with more than one correct answers

C1. (*a, d*) : Mole fraction and molality are independent of temperature.

C2. (*a, d*) : Moles of water = $\frac{25}{18}$ = 1.388

Moles of ethanol = $\frac{25}{46}$ = 0.543

Moles of acetic acid = $\frac{50}{60}$ = 0.833

Total moles = 1.388 + 0.543 + 0.833 = 2.764

x (water) = $\frac{1.388}{2.764}$ = 0.502

x (ethanol) = 0.196

x (acetic acid) = 0.302

x(ethanol) + x(acetic acid) = 0.498.

C3. (*a, b, d*) : 8g of O_2 = $\frac{8}{32}$ = 0.25 mol

No. of O atom = 0.25 × 2 × N_A = 0.5 N_A

(*a*) 11 g of CO_2 = $\frac{11}{44}$ = 0.25 mol

No. of O atoms = 2 × 0.25 × N_A = 0.5 N_A

(*b*) 14 g of CO = $\frac{14}{28}$ = 0.5 mol

No. of O atoms = 0.5 × N_A = 0.5 N_A

(*c*) 32 g of SO_2 = $\frac{32}{64}$ = 0.5 mol

No. of O atoms = 0.5 × 2 × N_A = 1.0 N_A

(*d*) 8 g of O_3 = $\frac{8}{48}$ = 0.166 mol

No. of O atoms = 0.166 × 3 × N_A
= 0.5 N_A

So, *a*, *b* and *d* have same number of O atoms.

C5. (*b, c*) : Molar mass of $MgSO_4$ = 24 + 32 + 4 × 16 = 120

% mass of Mg = $\frac{24}{120}$ × 100 = 20%

% mass of S = $\frac{32}{120}$ × 100 = 26.7%

C6. (*b, c*) : If N_A is Avogadro's number

(*i*) 14g (*ii*) 40 g

(*iii*) $\frac{108}{6.022\times10^{23}}$ = 1.79 × 10^{-22}g

iv) 32 g (*v*) 1.99 g (*vi*) 1g

So (*b*) and (*c*) are correct.

C7. (*c, d*) : Molarity and normality depends on temperature.

D. mcq based on comprehension

D1. (*b*) : Molar mass of $C_{10}H_8$ = 10 × 12 + 8 × 1 = 128

Mass of 10 molecules of $C_{10}H_8$ = $\frac{128}{6.02\times10^{23}}\times10$

= 2.12 × 10^{-21} g

D2. (*b*) : $\frac{32\times10^{20}}{6.02\times10^{23}}$ = 5.32 × 10^{-3} g

D3. (*a*) : $\frac{107.81}{6.02\times10^{23}}\times10^{6}$ = 1.79× 10^{-16} g

D4. (*c*) : % Fe = $\frac{4\times55.85}{\text{Mol. mass of haemoglobin}}$

∴ Mol. mass = $\frac{4\times55.85\times100}{0.335}$ = 66687

or = 66800

D5. (*c*) : $\frac{0.29}{78.96}$ = 3.672 × 10^{-3}

D6. (*a*) : Mol. mass = $\frac{1\times55.85\times100}{0.376}$

= 14853 ≃ 14850

D7. (*b*) : Mol. mass of peroxidase = $\frac{1\times78.96\times100}{0.29}$

= 27228

No. of atoms in 1 μg = $\frac{27228}{6.022\times10^{23}}\times10^{-6}$

= 4.52 × 10^{14}

D8. (*b*) : Mol. mass of haemoglobin = 66800 (Q.C4)

Moles of haemoglobin = $\frac{1\times10^{-3}}{66800}$

= 1.497 × 10^{-8}

Moles of iron = 1.497 × 10^{-8} × 4

= 5.99 × 10^{-8} ≈ 6.0 × 10^{-8}

D9. (*c*): 100 g of 109% oleum on dilution gives 109 g H_2SO_4 solution, and the SO_3 present in 100 g oleum will combine with 9 g water as

$$H_2O + SO_3 \longrightarrow H_2SO_4$$

Moles of water added = $\frac{9}{18}$ = Moles of free SO_3 present

∴ Mass of SO_3 = $\frac{9}{18}$ × 80 = 40 g

% of SO_3 = $\frac{40}{100}$ × 100 = 40%

D10. (*a*): 100 g of oleum on dilution gives 112 g H_2SO_4

Wt. of H_2SO_4 after dilution = 112 g

Wt. of water added = 12 g

Moles of H_2O = $\frac{12}{18}$ = 0.67 = Moles of free SO_3

Wt. of SO_3 = 0.67 × 80 = 53.6

% of SO_3 = 53.6 %, % of H_2SO_4 = 46.4

Integer Type and Numerical Value Type Questions

● **1.** (7) :1.420 g of anhydrous $ZnSO_4$ combine with water

= 2.528 – 1.420

= 1.108 g

Competition File

∴ 161.5 g of anhydrous $ZnSO_4$ combines with

water $= \frac{1.108}{1.420} \times 161.5$

$= 126$

No. of moles of water $= \frac{126}{18} = 7$

No. of water molecules in the salt $= 7$

● **2.** (7) : $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

Moles of CO used $= \frac{294}{28} = 10.5$

3 mol of CO are used to make = 2 mol of Fe

10.5 mol of CO are used to make $= \frac{2}{3} \times 10.5 = 7$

● **3.** (6) : Empirical formula comes out to be CH

Empirical formula mass = 12 + 1 = 13

Molecular mass = 2 × 39 = 78

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

$$= \frac{78}{13} = 6.$$

● **4.** (6) : $M_1V_1 + M_2V_2 \equiv M_3V_3$

$10 \times 428 + 3 \times 572 = M_3 \times 1000$

∴ $M_3 = 6$

● **5.** (7) : Density $= \frac{\text{Mass}}{\text{Volume}}$

10.5 g/cm^3 means 10.5 g of Ag is present in 1 cm^3

No. of Ag atoms in 1 cm^3 $= \frac{10.5}{108} \times 6.02 \times 10^{23}$

$= 5.85 \times 10^{22}$

No. of atoms in 1 cm $= (58.5 \times 10^{21})^{1/3}$

$= 3.882 \times 10^7$

No. of atoms in 1 cm^2 $= (3.882 \times 10^7)^2$

No. of atoms in 10^{-12}m^2 or 10^{-8} cm^2

$= (3.882 \times 10^7)^2 \times 10^{-8}$

$= 1.506 \times 10^7$

∴ Value of x = 7

● **6.** (3) : According to charge balance

$n(+2) + 2(+3) + 6(+4) + 18(-2) = 0$

$2n = 6$

or $n = 3$

● **7.** (5) : $3Br_2 + 3Na_2CO_3 \rightarrow 5NaBr + NaBrO_3 + 3CO_2$

● **8.** (8) : 29.2% (*w/w*) HCl means that 29.2g of HCl is present in 100g of solution.

Volume of HCl solution $= \frac{100}{1.25} = 80$g

Molarity of solution $= \frac{29.2/36.5}{80} \times 1000$

$= 10$M

Applying $M_1V_1 = M_2V_2$

$10 \times V_1 = 200 \times 0.4$

$V_1 = \frac{200 \times 0.4}{10} = 8$

● **9.** (4) : Boltzmann constant,

$$k = \frac{\text{Gas constant, R}}{\text{Avogadro number, } N_A}$$

or $R = k \times N_A$

$= 1.380 \times 10^{-23} \times 6.023 \times 10^{23}$

Since both the terms have 4 significant figures. The number of significant figures in the term having least number of significant figures is 4. No. of significant figures in calculated value will be 4.

●**10.** (126.00) :

$2MnCl_2 + 5K_2S_2O_8 + 8H_2O \longrightarrow 2KMnO_4 + 4K_2SO_4 + 6H_2SO_4 + 4\,HCl$ ...(*i*)

$2KMnO_4 + 5H_2C_2O_4 + 3H_2SO_4 \longrightarrow K_2SO_4 + 2MnSO_4 + 8H_2O + 10\,CO_2$ ...(*ii*)

Mass of oxalic acid added = 225 mg

Moles of oxalic acid $= \frac{225 \times 10^{-3}}{90}$

$= 2.5 \times 10^{-3}$

From equation (*ii*)

Moles of $KMnO_4$ react with 5 mol of $H_2C_2O_4$ = 2 mol

Moles of $KMnO_4$ react with 2.5×10^{-3} mol of $H_2C_2O_4 = \frac{2}{5} \times 2.5 \times 10^{-3} = 1 \times 10^{-3}$ mol

Moles of $MnCl_2$ required initially $= 1 \times 10^{-3}$ mol

Mass of $MnCl_2$ required initially $= 1 \times 10^{-3} \times 126$

$= 126$ mg

NCERT Exemplar Problems: MCQs Type I

1. (*b*) : Average reading of student A $= \frac{3.01 + 2.99}{2} = 3.00$

Average reading of student B $= \frac{3.05 + 2.95}{2} = 3.00$

Correct reading = 3.0

For both the students average value is same as the correct value. Hence readings of both the students are correct. But the readings of student A are close to each other (differ only by 0.02) and hence readings of student A are precise. Thus, results of student A are both precise and accurate.

2. (*c*) : $°C = \frac{5}{9}[°F - 32]$

$= \frac{5}{9} \times (200 - 32)$

$= 93.3°C$

3. (*c*) : Molarity $= \frac{5.85/58.5}{500} \times 1000$

$= 0.2$ mol L^{-1}

4. (*b*) $M_1V_1 = M_2V_2$

$5\text{ M} \times 500\text{ mL} = M_2 \times 1500\text{ mL}$

∴ $M_2 = \frac{5 \times 500}{1500} = 1.66$ M

Competition File

5. (*d*) : 4 g He = $\frac{4}{4}$ = 1 mol = N_A atoms.

46 g Na = $\frac{46}{23}$ = 2 mol = $2N_A$ atoms.

0.40 g Ca = $\frac{0.40}{40}$ = 0.01 mol = 10^{-2} N_A atoms.

12 g He = $\frac{12}{4}$ = 3 mol = 3 N_A atoms.

Thus, 12 g of He contain maximum number of atoms.

6. (*c*) Molar mass of glucose ($C_6H_{12}O_6$)

= 6 × 12 + 1 × 12 + 6 × 16 = 180

$$\text{Molarity} = \frac{0.9}{180 \times 1} = 0.005 \text{ M.}$$

7. (*d*)

$$\text{Molality} = \frac{\text{Moles of HCl}}{\text{Mass of water (in g)}} \times 1000$$

$$= \frac{18.25/36.5}{500} \times 1000 = 1 \text{ m.}$$

8. (*a*) : $\text{Moles of } H_2SO_4 = \frac{0.02}{1000} \times 100 = 0.002 \text{ mole}$

Molecules = 0.002 × 6.022 × 10^{23}

= 12.044 × 10^{20}

9. (*b*) : Molar mass of CO_2 = 12 + 2 × 16 = 44

$$\% \text{ carbon} = \frac{12}{44} \times 100 = 27.27\%$$

10. (*c*) : Empirical formula mass

= 12 + 2 + 16 = 30

$$n = \frac{\text{Molecular formula mass}}{\text{Empirical formula mass}}$$

$$= \frac{180}{30} = 6$$

Molecular formula = $(CH_2O)_6$

= $C_6H_{12}O_6$

11. (*a*) : Mass = 3.12 g mL^{-1} × 1.5 mL

∴ = 4.68 g

= 4.7

(upto 2 significant figures)

12. (*c*) A compound does not retain the physical properties of its constituent elements.

14. (*b*) Equation is not balanced.

15. (*b*) Statement (*b*) is correct according to law of conservation of mass.

NCERT Exemplar Problems: MCQs Type II

16. (*a*, *d*) 1 mole of O_2 at S.T.P.

= 6.022 × 10^{23} molecules of O_2

= 32 g of O_2.

17. (*b*, *c*) 1 L of 0.1 M H_2SO_4 contains = 0.1 mol H_2SO_4

1 L of 0.1 M NaOH contains = 0.1 mol NaOH

According to the given equation, 1 mol of H_2SO_4 reacts with 2 mol of NaOH. Therefore, 0.1 mol of NaOH will react with 0.1/2 = 0.05 mol of H_2SO_4 and 0.05 mol of H_2SO_4 will remain unreacted. Therefore, NaOH is the limiting reagent.

∴ 0.01 mol of NaOH will produce 0.05 mol Na_2SO_4

Molar mass of Na_2SO_4 = 2 × 23 + 1 × 32 + 4 × 16

= 142

Mass of Na_2SO_4 formed = 0.05 × 142 = 7.10 g

Volume of solution after mixing = 2 L

H_2SO_4 left unreacted = 0.05 mol

$$\text{Molarity of solution} = \frac{0.05}{2} = 0.025 \text{ mol L}^{-1}$$

18. (*c*, *d*) 16 g O_2 = $\frac{16}{32}$ = 0.5 mol

= 2 × 0.5 × N_A atoms = N_A atoms

4 g H_2 = $\frac{4}{2}$ = 2 mol

= 2 × 2 × N_A atoms = 4 N_A atoms

44 g CO_2 = $\frac{44}{44}$ = 1 mol

= 3 × 1 × N_A atoms = 3 N_A atoms

28 g N_2 = $\frac{28}{28}$ = 1 mol

= 2 × 1 × N_A atoms = 2 N_A atoms

32 g O_2 = $\frac{32}{32}$ = 1 mol

= 1 × 2 × N_A atoms = 2 N_A atoms

12 g C = $\frac{12}{12}$ = 1 mol = N_A atoms

23 g Na = $\frac{23}{23}$ = 1 mol = N_A atoms

∴ (*c*) and (*d*) are correct

19. (*a*, *b*) Molar conc. of NaOH solution = $\frac{20/40}{200} \times 1000$

= 2.5 M.

Molar conc. of KCl solution = $\frac{0.5}{200} \times 1000 = 2.5$ M

Molar conc. of NaOH solution = $\frac{40/40}{100} \times 1000$

= 10 M

Molar conc. of KOH solution = $\frac{20/56}{200} \times 1000$

= 1.78 M

Thus, (*a*) and (*b*) have the same concentration.

20. (*c*, *d*) 16 g CO = $\frac{16}{28}$ = 0.57 mol = 0.57 N_A molecules

28 g N_2 = $\frac{28}{28}$ = 1.0 mol = N_A molecules

14 g N_2 = $\frac{14}{28}$ = 0.5 mol = 0.5 N_A molecules

1.0 g H_2 = $\frac{1.0}{2}$ = 0.5 mol = 0.5 N_A molecules.

NCERT Exemplar Problems: Assertion-Reason Type

25. (*a*) Both A and R are true and A is the correct explanation of A.

26. (*b*) Both A and R are true but R is not the answer of A.

27. (*c*) A is true but R is false.

28. (*c*) A is false but R is true.

Unit Practice Test

for Board Examination

Time Allowed: 1 hr. ***Maximum Marks: 25***

1. Express decimal equivalent of 2/3 to three significant places. (1)
2. What is the mass of one ^{12}C atom in grams? (1)
3. Express the result of 12.26 + 16.0 + 1.089 to correct number of significant figures. (1)
4. Define limiting reagent. (1)
5. The human body temperature is 98.6°F. What is its value in °C and K ? (1)
6. One atom of an element weighs 9.75×10^{-23} g. Calculate its atomic mass. (2)
7. Round up the following to three significant figures :

   (*i*) 10.4207

   (*ii*) 0.04597

   (*iii*) 5.000

   (*iv*) 2808 (2)
8. The volume of a drop of rain was found to be 0.448 ml at N.T.P. How many molecules of water and number of atoms of hydrogen are present in this drop? (2)
9. (*i*) Calculate the amount of carbon dioxide that could be produced when 2 moles of carbon are burnt with 16 g of dioxygen.

   (*ii*) How many atoms of He are present in 52 u of He? (3)
10. Calculate the amount of $KClO_3$ needed to supply sufficient oxygen for burning 112 L of CO gas at N.T.P. (3)
11. Commercially available sulphuric acid contains 93% acid by mass and has density of 1.84 g mL^{-1}. Calculate

    (*i*) the molarity of the solution

    (*ii*) volume of concentrated acid required to prepare 2.5 L of 0.50 M H_2SO_4. (3)
12. What is meant by empirical formula and molecular formula? How are they related to each other? Explain with an example. A compound on analysis gave the following percentage composition: Na = 14.31%, S = 9.97%, H = 6.22% and O = 69.50%. Calculate the molecular formula of the compound assuming that all the hydrogens in the compound is present in combination with oxygen as water of crystallisation. The molecular mass of the compound is 322. (5)

To check your performance, see HINTS and SOLUTIONS to some questions at the end of Part I of the book.