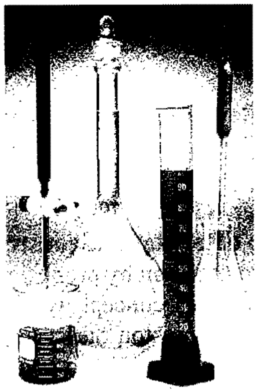


CHAPTER 1

BASIC PRINCIPLES

1.1 CHEMISTRY AND ITS SCOPE

Chemistry is a branch of physical science which deals with the study of matter, its physical and chemical properties, its chemical composition, the physical and chemical changes which it undergoes and the energy changes that accompany these processes.

All objects in this universe are composed of matter. Most of these objects are visible (solids and liquids) but some are invisible. Chemistry is termed as a material science because it is concerned with all material substances such as air, water, rocks, minerals, plants, animals including man, the earth on which we all live, and other planets. According to one of the famous scientists of twentieth century, **Linus Pauling,** *Chemistry is the science of substances, their properties, their structure and their transformations*.

Chemistry is a very interesting subject which touches almost every aspect of our lives, our culture and our environment. It has changed our civilization to a great extent. The present day chemistry has provided man with more comforts for a healthier and happier life. A large number of materials which we use these days were unknown at the turn of the century. A few decades back, our clothes and footwears were exclusively of natural origin such as vegetable fibres, wool, hair, skin of animals, etc., but, now the synthetic fibres produced in chemical factories have largely replaced them. Modern chemistry has given man new plastics, fuels, metal alloys, fertilizers, building materials, drugs, energy sources, etc.

During the last few decades, the expansion of chemistry has been tremendous. The field has become wide and complex. For convenience and better understanding of the subject, it has been divided into various branches. The four main branches of chemistry are:

(1) Organic chemistry; (2) Inorganic chemistry;

(3) Physical chemistry; (4) Analytical chemistry.

(1) Organic chemistry: It is concerned with the study of compounds of carbon except carbonates, bicarbonates, cyanides, isocyanides, carbides and oxides of carbon. It is actually the study of hydrocarbons and their derivatives.

(2) Inorganic chemistry: It deals with the study of all known elements and their compounds except organic compounds. It is concerned with the materials obtained from minerals, air, sea and soil.

(3) Physical chemistry: It is concerned with the physical properties and constitution of matter, the laws of chemical combination and theories governing reactions. The effect of temperature, pressure, light, concentration, etc., on reactions come under the scope of physical chemistry.

(4) Analytical chemistry: It deals with various methods of analysis of chemical substances both qualitative and quantitative. It includes chemical and physical methods of analysis.

A number of specialised branches have been introduced as to cope with the extraordinary expansion in the subject of chemistry. Some of the specialised branches are:

(i) Biochemistry: It comprises the studies of the substances related to living organisms and life processes.

(ii) Medicinal chemistry: It deals with the application of chemical substances for the prevention and cure of various diseases in living beings.

(iii) Soil and agriculture chemistry: It deals with the analysis and treatment of soils so as to increase its fertility for the better yields of crops. It is concerned with the chemicals used as fertilizers, insecticides, germicides, herbicides, etc.

(iv) Geochemistry: It includes the study of natural substances like ores and minerals, coal, petroleum, etc.

(v) Industrial chemistry: It deals with the study of chemical processes for the production of useful chemicals on a large scale at relatively low costs.

(vi) Nuclear chemistry: It is the most recent branch. It includes the study of nuclear reactions, the production of radioactive isotopes and their applications in various fields.

(vii) Structural chemistry: It deals with various techniques used for elucidation of the structure of chemical substances. It is concerned with the properties of substances in terms of their structure.

(viii) Polymer chemistry: It includes the study of chemical substances of very high molecular masses of the order of 100,000 or greater, called polymers—natural or artificial. This branch is gaining popularity as the use of plastics, rubber, synthetic fibres, silicones, etc., is on the increase these days.

(ix) Limnochemistry: It deals with the study of chemistry involved in the river water or water reservoirs.

(x) Phytochemistry: It includes the study of chemistry of plants.

Thus, it can be said that there is no other branch of science which is so wide in its scope as chemistry.

1.2 BRIEF HISTORY OF CHEMISTRY

It is difficult to specify the date when science of chemistry came into existence; however, its growth must have gone side by side with the growth of civilization. Broadly, the history of chemistry can be studied under five periods of its development.

(i) Ancient period up to 350 A.D.: In ancient times, many chemical operations such as souring of milk, conversion of sweet juices into wines, the conversion of wines into vinegar, etc., were known. Around 3000 B.C., techniques of making glass, pottery, pigments, dyes, perfumes and extraction of metals especially gold* and silver were known in China, India, Egypt and Greece. The beginning of chemistry as a science could probably be set about 400 B.C., when the theory was proposed that everything is composed of four elements: earth, air, fire and water. The first book of chemistry was written in Egypt around 300 A.D. The term chemistry meant the Egyptian art.

(ii) The alchemical period (350–1500): During this period, scientists called alchemists tried to discover two things: an elixir of life which could make man eternally young and a philospher's stone which could transmute base metals like zinc, copper, iron, etc., into gold. The alchemists failed in their efforts because no philospher's stone and elixir of life actually existed but we are indebted to them for designing new types of apparatus and for discovering new chemical operations such as distillation, sublimation, extraction of gold by amalgamation process and preparation of caustic alkalies from ashes of plants.

(iii) †Iatro chemistry period (1500–1650): During this era, chemists paid their attention towards medico-chemical problems. They believed that the primary object of chemistry was to prepare medicines and not to make gold from base metals. During this period, the study of gases was begun and quantitative experiments were undertaken for the first time. **Robert Boyle** (1627–1691) found that when a metal is heated in air, the mass increases. He also established the relationship between volume and pressure of a gas. In 1661, Boyle wrote the book 'The Skeptical Chymist' in which he criticised the basic ideas of alchemy.

(iv) The phlogiston period (1650–1774): The phlogiston theory was proposed by **Ernst Stahl** (1660–1734). Phlogiston was described as a substance in a combustible material which is given off when the material burns. This theory persisted for about 100 years and was a centre of much controversy. During the end of the eighteenth century, much work was done with gases, especially by **Joseph Black, Henry Cavendish, Joseph Priestley** and **Carl Scheele. Priestley** was a very conservative scientist. Even after his discovery of oxygen, he still believed in phlogiston theory.

(v) Modern period: Lavoisier (1743–1793), a French chemist, is regarded as the **father of modern chemistry.** He presented the exact explanation of combustion by proposing that oxygen is necessary for combustion. This concept was largely responsible for the overthrow of the phlogiston theory. Among his other contributions, he showed that water is composed of hydrogen and oxygen, proposed the theory of indestructibility of matter, presented a clear definition of an element and proposed a system of chemical nomenclature.

Another major step towards modern chemistry was taken in the first decade of the nineteenth century when the English chemist, **John Dalton,** postulated that all elements are made up of atoms. He pictured atoms as tiny, indestructible units that could combine to form **compound atoms** or **molecules.** Dalton proposed that each element has its own kind of atoms and the atoms of different elements differ in essentially nothing but their masses. He determined the relative masses of atoms of many elements. Thus, a new era had begun. The other important chemists of this period are:

(a) **Richter**—Law of Reciprocal Proportions (1794)
(b) **Proust**—Law of Definite Proportions (1799)
(c) **Gay-Lussac**—Law of Combining Volumes of Gases (1808)
(d) **Avogadro**—Avogadro Hypothesis (1811)
(e) **Berzelius**—Introduced the Modern Symbols for Elements (1813)
(f) **Faraday**—Laws of Electrolysis (1833)
(g) **Thomas Graham**—Law of Gaseous Diffusion (1861)
(h) **Mendeleev**—Periodic Law and Periodic Table (1869)
(i) **Arrhenius**—Theory of Ionization (1887)
(j) **Henry Becquerel**—Discovery of Radioactivity (1896)
(k) **Madam Curie**—Discovered Radium and Polonium (1898)

The twentieth century is regarded as an active era of chemistry. During this period, chemistry has made many contributions to human knowledge and civilization. Now, we live in a world of synthetic materials. Chemistry of today is actually helping in solving major problems of our present day civilization such as population explosion, food and diseases, depletion of sources of energy, depletion of natural sources and environmental pollution.

1.3 MATTER AND ENERGY

Besides life, matter and energy are regarded the two fundamental entities with which whole of the universe is composed of. **Matter is anything that has mass and occupies space.** All bodies in the universe conform to this definition. Mass is the quantity of matter in a particular sample of matter. Mass of a body is constant and does not change regardless of where it is measured. The mass of a

*Gold was probably the first metal to be used because it occurred as a free metal in the earth.
†*Iatro* is a Greek word meaning a physician.

body would be the same on the moon as it is on the earth. Our senses of sight and touch usually tell us that an object occupies space except in the case of colourless, odourless and tasteless gases where some other evidence is required to satisfy the definition of matter.

The term **weight** should not be used in place of mass as it has a different meaning. The term **weight** refers to the force with which an object is attracted towards earth. An object resting on earth experiences a force called its weight, W, that is equal to its mass m, multiplied by the acceleration due to gravity g, that is,

$$W = mg$$

The weight of an object thus depends on the value of 'g' which varies from place to place. However, the mass of an object is determined by comparing the weights of two objects, one of known mass, the other of unknown mass in the same location on earth as both experience the same gravitational acceleration.

Matter is indestructible, *i.e.*, it can neither be created, nor destroyed, but it can change its form; thus, the total quantity of matter of the universe is constant.

Energy is defined as the capacity of doing work. Anything which has the capacity to push the matter from one place to another possesses energy. There are various forms of energy such as heat, light, etc. Energy is neither created, nor destroyed, but can only be transformed from one form of energy to another.

The world became aware of the fact that matter can be converted into energy with the discovery of nuclear reactions, especially nuclear fission and nuclear fusion. The relationship between mass and energy was given by **Einstein.** The famous relation is:

$$E = mc^2$$

where, E = energy, m = mass and c = velocity of light.

On account of this equation, the above two laws are amalgamated into a single statement:

"The total amount of matter and energy available in the universe is fixed."

Example 1. *Calculate the amount of energy released in ergs, calories and in joules when* $0.001\,kg$ *of mass disappears.*

[*Given, Velocity of light* $= 3 \times 10^8\ ms^{-1}$]

Solution: According to Einstein equation, $E = mc^2$

$$m = 0.001\,\text{kg} = 1 \times 10^{-3}\ \text{kg};\ c = 3 \times 10^8\ \text{m s}^{-1}$$

$$E = (1 \times 10^{-3})(3 \times 10^8)^2 = 9 \times 10^{13}\ \text{J}$$

$$1\,\text{J} = 10^7\ \text{erg} = 0.24\ \text{cal}$$

$$9 \times 10^{13}\ \text{J} = 9 \times 10^{13} \times 10^7\ \text{erg} = 9 \times 10^{13} \times 0.24\ \text{cal}$$

$$= 9 \times 10^{20}\ \text{erg} = 2.16 \times 10^{13}\ \text{cal}$$

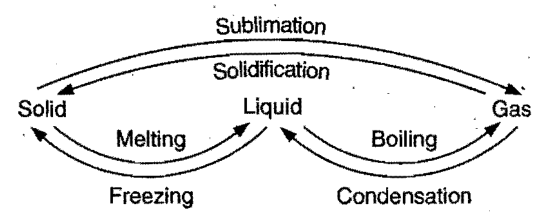
Classification of Matter

(i) Physical classification: Matter can exist in any one of three forms, (a) solid, (b) liquid and (c) gas.

In the solid state, substances are rigid. They have a definite shape and fixed volume. There is negligible effect of changes in pressure and temperature on their volumes. The individual particles that make up a solid occupy definite positions in the structure and are very near to one another. This form of matter is associated with minimum amount of energy.

In liquid state, substances have no definite shape but possess a fixed volume. There is slight effect of pressure and temperature on their volumes. They have the property of flowing. The particles are nearer to one another than in a gas and this form of matter is associated with energy more than solids.

In a gaseous state, substances have no definite shape and volume. Gases fill completely any vessel in which they are confined and thus occupy the whole space available to them. There is a large effect of pressure and temperature on their volumes. The particles are far apart from one another and move with very high speeds in all possible directions. This form of matter is associated with maximum amount of energy.

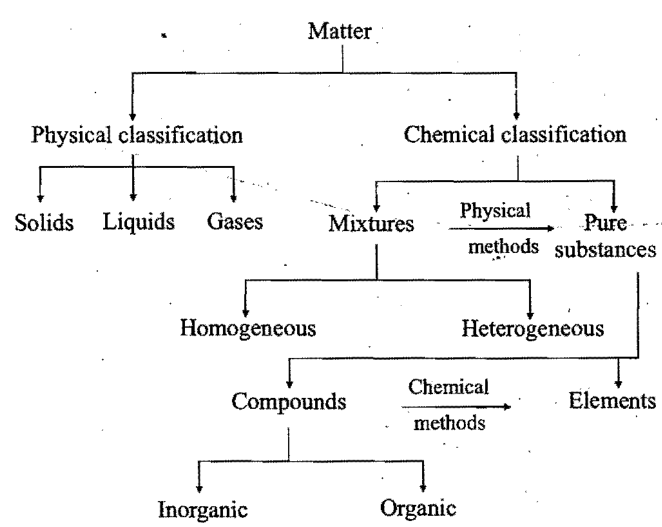


$$\text{Solid} \underset{-\,\text{energy}}{\overset{+\,\text{energy}}{\rightleftharpoons}} \text{Liquid} \underset{-\,\text{energy}}{\overset{+\,\text{energy}}{\rightleftharpoons}} \text{Gas}$$

$$\underset{\text{Release of energy}}{\overset{\text{Absorption of energy}}{\rightleftharpoons}}$$

Depending on temperature and pressure, a substance can exist in any one of the three forms of matter.

(ii) Chemical classification: Matter exists in nature in the form of chemical substances. **A pure substance is defined as a variety of matter, all samples of which have same composition and properties.** Pure substances are divided into elements and compounds. Most of the materials found in nature are in the form of mixtures consisting of two or more substances. There are two types of mixtures—Homogeneous and Heterogeneous. Both types of mixtures can be separated into their components (pure substances) by mechanical and physical methods. The classification can be summarized in the following way:



Properties of Matter: Properties are the characteristic qualities with the help of which different kinds of matter can be commonly recognised. In chemistry, substances are distinguished by two types of properties, *viz* (i) Chemical properties and (ii) Physical properties.

The chemical properties of substances are those in which they undergo change in composition either alone or by interactions with other substances, *i.e.*, to form new substances having different compositions from the substances which undergo change.

The properties of substances which are observed in absence of any change in composition under specific physical state are termed physical properties. Colour, density, melting point, boiling point, hardness, refractive index, thermal conductivity, electrical conductivity, malleability, ductility, etc., are some examples of the physical properties. The properties of matter can be further classified into two: (i) Intensive properties and (ii) Extensive properties. The **intensive properties** are those which do not depend upon the quantity of matter, *e.g.*, colour, density, melting point, boiling point, refractive index, etc. These properties are same irrespective of the quantity of the substance. Chemical properties are also intensive properties. The **extensive properties** of matter depend on the quantity of matter. Volume, mass, weight, energy, etc., are the extensive properties.

1.4 ELEMENTS AND COMPOUNDS

Elements are pure substances that cannot be decomposed into simpler substances by chemical changes. The smallest particles of an element possess the same properties as the bigger particles. An element can also be defined as **a pure substance which consists of only one type of atoms.** Due to discovery of isotopes, this definition does not seem to be correct. The modern definition of an element is that it is a simple individual which has a definite atomic number (see atomic structure) and has a definite position in the periodic table. It cannot be decomposed in a chemical change. In chemistry, the elements are the chemical alphabet and compounds are the words, *i.e.*, combinations of elements.

There are presently 117 different elements known. Every element has been given a definite name and for convenience a nick name which in chemical language is called a **symbol.** *Symbol is a small abbreviation to represent a full and lengthy name of the element.* Symbols have been derived:

(i) either by taking the first letter of the name of the element which is capitalized:

O—Oxygen	N—Nitrogen	F— Fluorine
C—Carbon	H—Hydrogen	U— Uranium
P—Phosphorus	S—Sulphur	I— Iodine

(ii) or by taking the first letter and one more letter from the name of the element. The first letter is always capitalized.

Ca—Calcium	Ni—Nickel	Al—Aluminium
Mg—Magnesium	Co—Cobalt	Bi—Bismuth
Cl—Chlorine	Br—Bromine	Ba—Barium

(iii) or from names of the elements in other languages such as Latin, German, etc.

Na—Sodium (Latin name Natrium)
Cu—Copper (Latin name Cuprum)
Fe—Iron (Latin name Ferrum)
Ag—Silver (Latin name Argentum)
Pb—Lead (Latin name Plumbum)
Au—Gold (Latin name Aurum)
K—Potassium (Latin name Kalium)
Hg—Mercury (Latin name Hydragyrum)
W—Tungsten (German name Wolfram)

Out of 117 elements known, 88 have been isolated from natural sources and the remaining have been prepared by artificial means. The man made elements are:

S.No.	Name	Symbol	S.No.	Name	Symbol
1.	Neptunium	Np	16.	Hassium or Unniloctium	Hs or Uno
2.	Plutonium	Pu	17.	Meitnerium or Unnilennium	Mt or Une
3.	Americium	Am	18.	Ununnilium	Uun
4.	Curium	Cm	19.	Unununium	Uuu
5.	Berkelium	Bk	20.	Ununbium	Uub
6.	Californium	Cf	21.	Ununtrium	Uut
7.	Einsteinium	Es	22.	Ununquadium	Uuq
8.	Fermium	Fm	23.	Ununpentium	Uup
9.	Mendelevium	Md	24.	Ununhexium	Uuh
10.	Nobelium	No	25.	Ununoctium	Uuo
11.	Lawrencium	Lr	26.	Technetium	Tc
12.	Kurchatovium	Ku	27.	Promethium	Pm
13.	Hahnium	Ha	28.	Astatine	At
14.	Seaborgium or Unnilhexium	Sg or Unh	29.	Francium	Fr
15.	Nielsbohrium or Unnilseptium	Bh or Uns			

The elements from S. No. 1 to 25 are called transuranic elements. The credit for the discovery of most of the transuranic elements goes to the scientist **G.T. Seaborg.** The first artificially produced element was technetium. It was synthesised in 1937 by scientists at the University of California at Berkley.

Most of the earth's crust is made up of a small number of elements. Only ten of the naturally occurring elements make up 99% mass of the earth's crust, oceans and atmosphere. The following table shows the abundance of highly abundant elements in nature:

Note : Among the naturally occurring elements, ${}^{1}_{1}H$ is lightest and ${}^{238}_{92}U$ is the heaviest atom.

Abundance of Elements (Earth's Crust, Oceans and Atmosphere)

Oxygen	49.5%	Chlorine	0.19%
Silicon	25.7%	Phosphorus	0.12%
Aluminium	7.5%	Manganese	0.09%
Iron	4.7%	Carbon	0.08%
Calcium	3.4%	Sulpur	0.06%
Sodium	2.6%	Barium	0.04%
Potassium	2.4%	Chromium	0.033%
Magnesium	1.9%	Nitrogen	0.030%
Hydrogen	0.87%	Fluorine	0.027%
Titanium	0.58%	Zirconium	0.023%
	99.2%		0.7%

If the entire universe is considered, then 90% of matter is hydrogen. Helium is the second most abundant element amounting to 9% and the remaining elements make up only 1% of the universe with oxygen, neon, carbon and nitrogen next in order of decreasing abundance.

The commercial use of an element depends not only upon its abundance but also upon its accessibility. Some of the common elements such as copper, zinc, tin and lead are not abundant but are found in nature in rich deposits from which these can be easily extracted. On the other hand, the elements such as titanium and zirconium which are found in abundance in nature are not widely used because their ores are not rich and their extraction is difficult and expensive.

Metals, Non-metals and Metalloids

All the elements may be classified into two groups, **metals** and **non-metals.** The division is based on both physical and chemical properties.

Metals are regarded as those elements which possess the following properties:

(i) They are generally solids at ordinary conditions. Mercury is an exception which is in liquid state.
(ii) They are lustrous in nature.
(iii) They possess high density.
(iv) They are good conductors of electricity and heat.
(v) They are malleable and ductile.
(vi) They possess generally high melting and boiling points.
(vii) They react with mineral acids liberating hydrogen.
(viii) They form basic oxides.
(ix) They form non-volatile hydrides if combine with hydrogen.
(x) They have molecules usually mono-atomic in the vapour state.

Sodium, calcium, aluminium, copper, silver, zinc, iron, nickel, gold, mercury, etc., are the examples of metals.

The non-metals do not show the above properties. Six of the non-metals, carbon, boron, phosphorus, sulphur, selenium and iodine, are solids. Bromine is the only liquid non-metal at room temperature and normal pressure. The remaining non-metals; nitrogen, oxygen, fluorine, chlorine, hydrogen, helium, argon, neon, krypton, xenon and radon are gases. Non-metals are generally (i) brittle, (ii) non-lustrous, (iii) having low melting and boiling points, (iv) non-conductors of heat, (v) capable of forming acidic oxides or neutral oxides, (vi) not capable of evolving hydrogen from acids, and (vii) capable of forming volatile hydrides.

There are some elements which do not fit completely into either the metal or non-metal class. Elements which have some properties of both metals and non-metals are called semi-metals or metalloids. The semi-metals are silicon, germanium, arsenic, antimony and tellurium.

The above classification of elements is a rough one as certain metals like lithium, sodium, potassium possess low density; certain non-metals like hydrogen and graphite (a form of carbon) are good conductors of electricity. Metals rarely combine with one another while non-metals combine with one another to form compounds. Metals and non-metals commonly combine with each other to form compounds.

Compounds

Compounds are also pure substances that are composed of two or more different elements in a fixed proportion by mass. Compounds containing more than four elements are rare. The properties of a compound are altogether different from the properties of the elements from which it has been constituted. The compound **water** has a definite composition, *i.e.*, 11.2% hydrogen and 88.8% oxygen, Thus, the two are present in the ratio of 1 : 8 by mass. The properties of water are totally different from the properties of hydrogen and oxygen both. Hydrogen and oxygen are in gaseous state while water is in liquid state under ordinary atmospheric conditions. Oxygen supports combustion while hydrogen is combustible but water is normally used for extinguishing fire. Component elements in compounds can be separated only by chemical means and not by physical methods.

Compounds are classified into two types:

(i) Organic compounds: The compounds obtained from living sources are termed organic compounds. The term organic is now applied to hydrocarbons (compounds of carbon and hydrogen) and their derivatives.

(ii) Inorganic compounds: The compounds obtained from non-living sources such as rocks and minerals are termed inorganic compounds. The compounds of all elements except hydrocarbons and their derivatives are included in this category. The number of organic compounds is very large in comparison to inorganic compounds.

Some Specific Properties of Substances: Some specific properties of substances are given below:

(i) Deliquescence: The property of certain compounds of taking up the moisture present in atmosphere and becoming wet when exposed, is known as deliquescence. These compounds are known as deliquescent. Sodium hydroxide, potassium hydroxide, anhydrous calcium chloride, anhydrous magnesium chloride, anhydrous ferric chloride, etc., are the examples of deliquescent compounds. Sodium chloride is not deliquescent but when common salt is placed in atmosphere it becomes wet due to presence of an impurity of magnesium chloride.

(ii) Hygroscopicity: Certain compounds combine with the moisture of atmosphere and are converted into hydroxides or hydrates. Such substances are called hygroscopic. Anhydrous copper sulphate, quick lime (CaO), anhydrous sodium carbonate, etc., are of hygroscopic nature.

(iii) Efflorescence: The property of some crystalline substances of losing their water of crystallisation on exposure and becoming powdery on the surface is called efflorescence. Such salts are known as efflorescent. The examples are:

Ferrous sulphate ($FeSO_4 \cdot 7H_2O$), sodium carbonate ($Na_2CO_3 \cdot 10H_2O$), sodium sulphate ($Na_2SO_4 \cdot 10H_2O$), potash alum $[K_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O]$, etc.

(iv) Malleability: This property is shown by metals. When the solid is beaten and does not break but is converted into a thin sheet, it is said to possess the property of malleability. Copper, gold, silver, aluminium, lead, etc., can be easily hammered into sheets. Gold is the most malleable metal.

(v) Ductility: The property of a metal to be drawn into wires is termed ductility. Copper, silver, gold, aluminium, iron, etc., are ductile in nature. Platinum is the most ductile metal.

(vi) Elasticity: When the stress is small, the solid completely regains its original shape, size or volume after the deforming force is removed. The solid is then said to be elastic. Steel, glass, ivory, etc., are elastic bodies.

(vii) Plasticity: When stress is increased on a metal, a limit is reached beyond which, if the stress is removed, the solid does not come back to its original shape or size. It acquires a permanent deformation. Such materials can be given any shape without any difficulty.

(viii) Brittleness: The solid materials which break into small pieces on hammering are called brittle. The solids of non-metals are generally brittle in nature.

(ix) Hardness: A material is said to be harder than the other if it can scratch it. The hardness is measured on Mho's scale. For this purpose, ten minerals have been selected which have been assigned hardness from 1 to 10.

Hardness	Mineral	Hardness	Mineral
1	Talc	6	Orthoclase
2	Gypsum	7	Quartz
3	Calcite	8	Topaz
4	Fluorite	9	Corundum
5	Apatite	10	Diamond

On Mho's scale, hardness of diamond is maximum and that of talc is minimum. If a material can scratch topaz but cannot scratch corundum it possesses hardness equal to 8.

1.5 MIXTURES

A **mixture** is a material containing two or more substances either elements or compounds or both in any proportion. Substances which form a mixture are called **components.** Components are present in the mixture without loss of their identity. There are two types of mixtures—**homogeneous** and **heterogeneous.** In a homogeneous mixture, the components are mixed uniformly to microscopic level. The components cannot be seen by naked eye or with the help of a microscope. The mixture is uniform throughout having a single **phase*.** The homogeneous mixture is **isotropic** in nature, *i.e.*, every portion of it has the same composition and properties.

Alloys such as brass, steel, 22-carat gold; solutions such as common salt dissolved in water, sugar dissolved in water, iodine dissolved in carbon tetrachloride, benzene in toluene, methyl alcohol in water; gasoline (a mixture of hydrocarbons), air, etc., are some of the examples of homogeneous mixtures.

A heterogeneous mixture is not uniform. It can have two or more phases. The components can be seen by naked eye or with the help of a microscope. It has **anisotropic** properties, *i.e.*, properties are not uniform throughout the mixture. Soil, a mixture of sulphur and sand, a mixture of iron filings and sand, smoke, etc., are the examples of heterogeneous mixtures.

The components of a mixture differ in many of their physical and chemical properties. The advantage of this difference is taken in the separation of a mixture. The method of separation employed should not bring about the destruction of any one of the components. Some preliminary techniques based on physical properties are described here in brief.

(i) Filtration: This method is useful when one of the components is an insoluble solid in a solvent. The insoluble solid is obtained by filtration of the suspension through filter-paper. For example, common salt containing sand is separated by filtration. The mixture is mixed with water. It is shaken so as to dissolve common salt. The sand remains insoluble. The suspension is put to filtration. The sand collects on the filter-paper. It is taken in a basin and dried by heating. The filtrate is taken in evaporating dish and heated till whole of the water is evaporated. Solid common salt is thus obtained in the dish.

Sugar containing charcoal, potassium nitrate containing saw dust or mixtures having insoluble components can be separated by filtration.

(ii) Sublimation: It is a process in which a solid substance is directly converted into its vapours by application of heat and vapour is reconverted into solid by subsequent cooling. The method is used when one of the components undergoes sublimation and other components are not decomposed by heating. For example, naphthalene can be separated from common salt by sublimation. Similarly, a mixture of ammonium chloride and potassium chloride can be separated by sublimation as ammonium chloride sublimes on heating.

(iii) Distillation: It is a process of converting a liquid into its vapour by heating and then condensing the vapours again into the same liquid by cooling. Thus, distillation involves vaporisation and condensation both.

Distillation = Vaporisation + Condensation

This method is employed to separate liquids which have different boiling points or a liquid from non-volatile solid or solids either in solution or suspension. The mixture of copper sulphate and water or mixture of water (b.p. 100°C) and methyl alcohol (b.p. 45°C) can be separated by this method.

* Phase is defined as part of a system which has uniform properties and composition. A solution or mixture of sugar and water is a one phase system. Every drop of the solution has same properties and same composition.

(iv) Magnetic separation: If one of the components of a mixture has magnetic properties, it can be separated by using a magnet. Iron is separated from a mixture of iron filings and sulphur by moving a magnet through the mixture.

(v) Solvent extraction: This method is based on the preferential solubility of one of the components of the mixture in a particular solvent (usually a low-boiling organic solvent) which forms a distinctly separate layer with the other liquid if present in the mixture. For example, iodine present in water can be recovered with the help of ether or carbon disulphide. For this method, a separating funnel is utilized. The aqueous solution of iodine is taken in separating funnel to which ether is added. The funnel is shaken. Two layers are formed. The upper layer which is dark brown consists of ether and iodine and the colourless lower layer consists of only water. The lower layer is taken out. The coloured layer is then poured out and ether is removed cautiously by distillation when iodine is left behind.

Two immiscible liquids such as water and oil can also be separated by the use of a separating funnel.

Example 2. *How will you separate the following mixtures?*

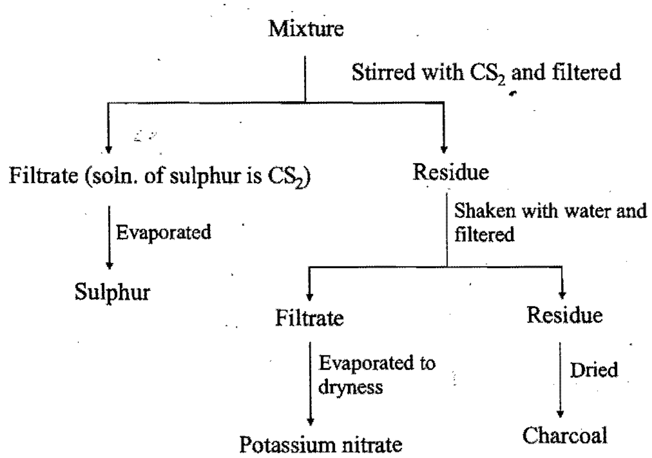
(*a*) *Sulphur, potassium nitrate and charcoal,*

(*b*) *Sand, common salt, iron filings and naphthalene,*

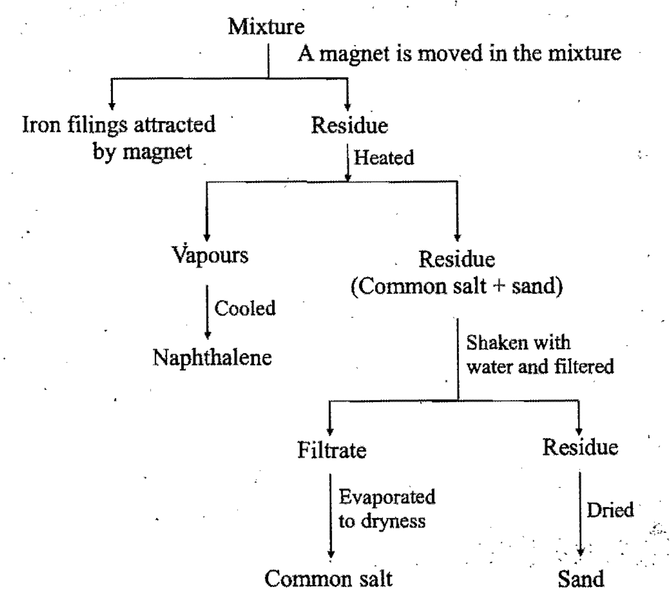
(*c*) *Powdered glass, ammonium chloride and potassium chloride.*

Solution:

(a) (i) Sulphur is soluble in carbon disulphide,
(ii) Potassium nitrate is soluble in water,
(iii) Charcoal is insoluble in carbon disulphide as well as in water.



(b) (i) Iron filings are separated by a magnet,
(ii) Naphthalene sublimes on heating,
(iii) Sand is insoluble in water.



(c) (i) Ammonium chloride sublimes on heating,
(ii) Potassium chloride is soluble in water,
(iii) Powdered glass is insoluble in water.

1.6 ALLOYS

When two or more elements are melted together and resulting liquid is allowed to solidify, the product so obtained is called an alloy if it possesses metallic properties. An alloy may consist of a mixture of a metal with another metal, a metal with a non-metal or a metal with both metal and non-metal.

Alloys are prepared because they have properties more suitable for certain applications than do the simple metals. Alloys are used because they are harder and stronger, have desirable casting properties, special physical properties such as magnetic properties and resistance to corrosion in certain environments. Melting point of an alloy is normally lower than the melting point of either of the pure components. Thermal and electrical conductivities are normally reduced in alloys.

An alloy containing one component mercury is called amalgam. Most of the metals form amalgams. Iron, platinum, tungsten, etc., are few metals which do not form amalgams.

Alloys are mainly classified into two distinct types, namely ferrous and non-ferrous. Ferrous alloys always contain iron, carbon and one or two of the other elements such as manganese, nickel, chromium, copper, vanadium, molybdenum, tungsten, etc. When the percentage of carbon in the alloy is below 0.1, the alloy is termed the iron alloy and if it is above 0.1, the alloys are called steels. When iron is not present in the alloy, it is termed a non-ferrous alloy. Some of the important alloys have been listed below:

Alloy	Composition	Main uses
1. Brass	Cu 60–80%, Zn 20–40%	Utensils, condenser tubes, electrical goods, cartridge shell
2. Bronze	Cu 75–90%, Sn 10–25%	Coins, statues, utensils.

3. German silver	Cu 56%, Zn 24%, Ni 20%	Utensils, resistance coils
4. Gun metal	Cu 87%, Sn 10%, Zn 3%	Machine parts, guns
5. Rolled gold	Cu 95%, Al 5%	Artificial jewellery
6. Magnalium	Al 94%, Mg 6%	Balance beams, light instruments
7. Electron	Mg 95%, Zn 5%	Construction of aircraft
8. Duralumin	Al 95%, Cu 4%, Mn 0.5%, Mg 0.5%	Making aeroplanes
9. Type metal	Pb 82%, Sb 15%, Sn 3%	Making printing types
10. Solder	Pb 50–70%, Sn 30–50%	Soldering
11. Britannia	Sn 93%, Sb 5%, Cu 2%	Tableware
12. Wood's metal	Bi 50%, Pb 25%, Sn 12.5%, Cd 12.5%	Electric fuses and other safety devices
13. Nichrome	Ni 60%, Cr 15%, Fe 25%	Electrical resistances
14. Constantan	Ni 40%, Cu 60%	Electrical resistances
15. Monel metal	Ni 70%, Cu 30%	Chemical plants
16. Invar	Ni 35%, Steel 65%	Surveying instruments, pendulums, chronometers
17. Stainless steel	Fe 89.4%, Cr 10%, Mn 0.35%, C 0.25%	Utensils, ornamental pieces.

1.7 PHYSICAL AND CHEMICAL CHANGES

Matter undergoes two types of changes; physical and chemical. A physical change is one in which a substance changes its physical state but keeps its chemical identity. In physical change, a new substance does not come into existence. Mass remains the same. Physical properties are altered. This is a temporary change. For example, water shows all of its chemical properties whether it is in the form of ice or water or steam. Ice melts to form water and water can be converted again into ice by placing it in a freezer. When 10 g of ice melts, 10 g of water is obtained. Melting, evaporation, condensation, freezing, sublimation, distillation, passing of electric current through metallic conductor, making of magnet from an iron piece, are some examples of physical changes.

In a chemical change, a new substance or substances come into existence. The starting materials called reactants, are used up and new substances called products, are formed. The composition of the new substances is different from that of the starting materials. It is a permanent change as it is not easy to obtain the starting materials again from the products.

Energy is always released or absorbed when chemical or physical changes occur. Energy is required to melt ice and energy is required to boil water. Conversely, the condensation of steam to form liquid water always liberates energy, as does the freezing of liquid water to form ice. Chemical changes either release energy (exothermic) or adsorb energy (endothermic).

Chemical changes are of various types. The important ones are:

(i) Combination: Two or more substances react to form one product. When a compound is obtained by the direct reaction between elements, it is termed direct union or synthesis.

$$H_2 + Cl_2 = 2HCl \quad ; \quad C + O_2 = CO_2$$

$$2Mg + O_2 = 2MgO \quad ; \quad SO_2 + H_2O = H_2SO_3$$

(ii) Decomposition: When a compound is broken down into two or more simple constituents, the change is called decomposition. Often heat is utilised for the decomposition. Such decomposition is termed **thermal decomposition.**

$$2HgO = 2Hg + O_2\ ; \quad 2KClO_3 = 2KCl + 3O_2$$

$$CaCO_3 = CaO + CO_2;\ 2NaHCO_3 = Na_2CO_3 + H_2O + CO_2$$

(iii) Substitution: When one element enters into a compound by the replacement of the other element, the change is termed substitution.

$$CuSO_4 + Zn = Cu + ZnSO_4;\ C_2H_6 + Cl_2 = C_2H_5Cl + HCl$$

$$2KI + Cl_2 = 2KCl + I_2 \quad ; \quad Zn + 2HCl = ZnCl_2 + H_2$$

(iv) Addition: Something is added to a chemical substance without elimination.

$$KI + I_2 = KI_3$$

$$C_2H_4 + HBr = C_2H_5Br$$

$$CH_3CHO + HCN = CH_3CH\begin{matrix} \diagup OH \\ \diagdown CN \end{matrix}$$

(v) Internal rearrangement: When nothing is added or nothing is eliminated from a chemical substance but due to rearrangement of the various atoms present in a molecule, a new compound comes into existence. When ammonium cyanate is heated, a new substance urea is formed.

$$NH_4CNO = NH_2CONH_2$$

The chemical change is termed isomerisation, when one isomer is converted into another.

$$\underset{n\text{-Butane}}{CH_3-CH_2-CH_2-CH_3} \xrightarrow[300°C]{AlCl_3/HCl} \underset{\text{2-Methyl propane}}{CH_3-\overset{\overset{CH_3}{|}}{CH}-CH_3}$$

(vi) Polymerization: Two or more molecules of a substance combine to form a giant molecule,

$$\underset{\text{Acetylene}}{3C_2H_2} = \underset{\text{Benzene}}{C_6H_6}$$

$$\underset{\text{Formaldehyde}}{nHCHO} = \underset{\text{Para formaldehyde}}{(HCHO)_n}$$

(vii) Double decomposition: An exchange of partners occurs between two compounds.

$$AB + CD = AD + CB$$

$$BaCl_2 + Na_2SO_4 = BaSO_4 + 2NaCl$$

$$KOH + HCl = KCl + H_2O$$

The reaction is also termed neutralisation, *i.e.*, a reaction between an acid and a base to form a salt and water molecule.

$$FeCl_3 + \underbrace{3HOH}_{\text{Water}} = Fe(OH)_3 + 3HCl$$

Reaction of above type is termed hydrolysis.

1.8 LAWS OF CHEMICAL COMBINATION

In order to understand the composition of various compounds, it is necessary to have a theory which accounts for both qualitative and quantitative observations during chemical changes. Observations of chemical reactions were most significant in the development of a satisfactory theory of the nature of matter. These observations of chemical reactions are summarised in certain statements known as laws of chemical combination.

(i) Law of conservation of mass: The law was first stated by **Lavoisier** in 1774. It is also known as the law of indestructibility of matter. According to this law, **in all chemical changes, the total mass of a system remains constant or in a chemical change, mass is neither created nor destroyed.** This law was tested by **Landolt.** All chemical reactions follow this law. Thus, this law is the basis of all quantitative work in chemistry.

Example: 1.70 *g of silver nitrate dissolved in* 100 *g of water is taken.* 0.585 *g of sodium chloride dissolved in* 100 *g of water is added to it and chemical reaction occurs.* 1.435 *g of silver chloride and* 0.85 *g of sodium nitrate are formed.*

Solution: Total masses before chemical change

$$= \text{Mass of } AgNO_3 + \text{Mass of NaCl} + \text{Mass of water}$$

$$= 1.70\,g + 0.585\,g + 200.0\,g$$

$$= 202.285\ g$$

Total masses after the chemical reaction,

$$= \text{Mass of AgCl} + \text{Mass of } NaNO_3 + \text{Mass of water}$$

$$= 1.435\,g + 0.85\,g + 200.0\,g$$

$$= 202.285\,g$$

Thus, in this chemical change,

Total masses of reactants = Total masses of products

This relationship holds good when reactants are completely converted into products.

In case, the reacting materials are not completely consumed, the relationship will be

Total masses of reactants = Total masses of products + Masses of unreacted reactants

(ii) Law of definite or constant proportions: This law was presented by **Proust** in 1799 and may be stated as follows:

A chemical compound always contains the same element combined together in fixed proportion by mass, *i.e.,* a chemical compound has a fixed composition and it does not depend on the method of its preparation or the source from which it has been obtained.

For example, carbon dioxide can be obtained by using any one of the following methods:

(a) by heating calcium carbonate,
(b) by heating sodium bicarbonate,
(c) by burning carbon in oxygen,
(d) by reacting calcium carbonate with hydrochloric acid.

Whatever sample of carbon dioxide is taken, it is observed that carbon and oxygen are always combined in the ratio of 12 : 32 or 3 : 8.

The converse of this law that when same elements combine in the same proportion, the same compound will be formed, is not always true. For example, carbon, hydrogen and oxygen when combine in the ratio of 12 : 3 : 8 may form either ethyl alcohol (C_2H_5OH) or dimethyl ether (CH_3OCH_3) under different experimental conditions.

(iii) Law of multiple proportions: This law was put forward by **Dalton** in 1808. According to this law, **if two elements combine to form more than one compound, then the different masses of one element which combine with a fixed mass of the other element, bear a simple ratio to one another.**

Hydrogen and oxygen combine to form two compounds H_2O (water) and H_2O_2 (hydrogen peroxide).

In water,	Hydrogen 2 parts	Oxygen 16 parts
In hydrogen peroxide,	Hydrogen 2 parts	Oxygen 32 parts

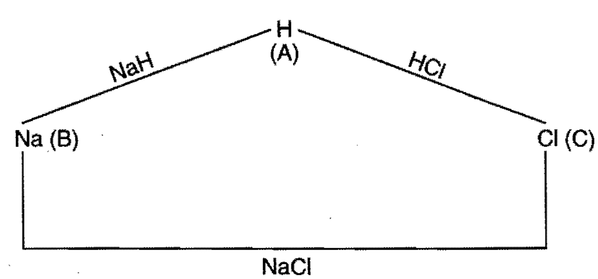
The masses of oxygen which combine with same mass of hydrogen in these two compounds bear a simple ratio 1 : 2.

Nitrogen forms five stable oxides.

N_2O	Nitrogen 28 parts	Oxygen 16 parts
N_2O_2	Nitrogen 28 parts	Oxygen 32 parts
N_2O_3	Nitrogen 28 parts	Oxygen 48 parts
N_2O_4	Nitrogen 28 parts	Oxygen 64 parts
N_2O_5	Nitrogen 28 parts	Oxygen 80 parts

The masses of oxygen which combine with same mass of nitrogen in the five compounds bear a ratio 16 : 32 : 48 : 64 : 80 or 1 : 2 : 3 : 4 : 5.

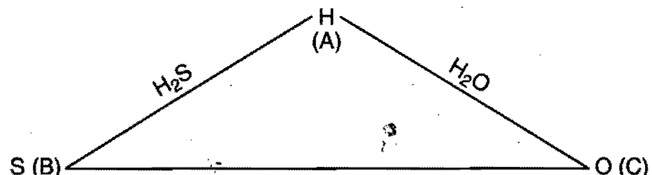
(iv) Law of reciprocal proportions: This law was given by **Richter** in 1794. The law states that **when definite mass of an element *A* combines with two other elements *B* and *C* to form two compounds and if *B* and *C* also combine to form a compound, their combining masses are in same proportion or bear a simple ratio to the masses of *B* and *C* which combine with a constant mass of *A*.**



For example, hydrogen combines with sodium and chlorine to form compounds NaH and HCl respectively.

In NaH,	Sodium 23 parts	Hydrogen one part
In HCl,	Chlorine 35.5 parts	Hydrogen one part

Sodium and chlorine also combine to form NaCl in which 23 parts of sodium and 35.5 parts of chlorine are present. These are the same parts which combine with one part of hydrogen in NaH and HCl respectively.



Hydrogen combines with sulphur and oxygen to form compounds H_2S and H_2O respectively.

In H_2S,	Hydrogen 2 parts	Sulphur 32 parts
In H_2O,	Hydrogen 2 parts	Oxygen 16 parts

Thus, according to this law, sulphur should combine with oxygen in the ratio of 32 : 16 or a simple multiple of it. Actually, both combine to form SO_2 in the ratio of 32 : 32 or 1 : 1.

The law of reciprocal proportions is a special case of a more general law, the law of equivalent masses, which can be stated as under :

"In all chemical reactions, substances always react in the ratio of their equivalent masses."

(v) Law of gaseous volumes: This law was enunciated by **Gay-Lussac** in 1808. According to this law, gases react with each other in the simple ratio of their volumes and if the product is also in gaseous state, the volume of the product also bears a simple ratio with the volumes of gaseous reactants when all volumes are measured under similar conditions of temperature and pressure.

$$\underset{1\text{ vol}}{H_2} + \underset{1\text{ vol}}{Cl_2} = \underset{2\text{ vol}}{2HCl} \qquad \text{ratio} \qquad 1:1:2$$

$$\underset{2\text{ vol}}{2H_2} + \underset{1\text{ vol}}{O_2} = \underset{2\text{ vol}}{2H_2O} \qquad \text{ratio} \qquad 2:1:2$$

$$\underset{2\text{ vol}}{2CO} + \underset{1\text{ vol}}{O_2} = \underset{2\text{ vol}}{2CO_2} \qquad \text{ratio} \qquad 2:1:2$$

$$\underset{1\text{ vol}}{N_2} + \underset{3\text{ vol}}{3H_2} = \underset{2\text{ vol}}{2NH_3} \qquad \text{ratio} \qquad 1:3:2$$

SOME SOLVED EXAMPLES

Example 3. *What mass of sodium chloride would be decomposed by 9.8 g of sulphuric acid, if 12 g of sodium bisulphate and 2.75 g of hydrogen chloride were produced in a reaction assuming that the law of conservation of mass is true?*

Solution: $NaCl + H_2SO_4 = NaHSO_4 + HCl$

According to law of conservation of mass,

Total masses of reactants = Total masses of products

Let the mass of NaCl decomposed be x g, so

$$x + 9.8 = 12.0 + 2.75$$
$$= 14.75$$
$$x = 4.95\text{ g}$$

Example 4. *In an experiment, 2.4 g of iron oxide on reduction with hydrogen yield 1.68 g of iron. In another experiment, 2.9 g of iron oxide give 2.03 g of iron on reduction with hydrogen. Show that the above data illustrate the law of constant proportions.*

Solution:

In the first experiment

The mass of iron oxide = 2.4 g

The mass of iron after reduction = 1.68 g

The mass of oxygen = Mass of iron oxide – Mass of iron

$= (2.4 - 1.68) = 0.72\text{ g}$

Ratio of oxygen and iron $= 0.72 : 1.68 = 1 : 2.33$

In the second experiment

The mass of iron oxide = 2.9 g

The mass of iron after reduction = 2.03 g

The mass of oxygen $= (2.9 - 2.03) = 0.87\text{ g}$

Ratio of oxygen and iron $= 0.87 : 2.03 = 1 : 2.33$

Thus, the data illustrate the law of constant proportions, as in both the experiments the ratio of oxygen and iron is the same.

Example 5. *Carbon combines with hydrogen to form three compounds A, B and C. The percentages of hydrogen in A, B and C are 25, 14.3 and 7.7 respectively. Which law of chemical combination is illustrated?*

Solution:

Compound	% of Hydrogen	% of Carbon
A	25.0	(100 – 25.0) = 75.0
B	14.3	(100 – 14.3) = 85.7
C	7.7	(100 – 7.7) = 92.3

In Compound *A*

25 parts of hydrogen combine with 75 parts of carbon

1 part of hydrogen combines with 75/25

= 3 parts of carbon

In Compound *B*

14.3 parts of hydrogen combine with 85.7 parts of carbon

1 part of hydrogen combines with 85.7/14.3

= 6.0 parts of carbon

In Compound *C*

7.7 parts of hydrogen combine with 92.3 parts of carbon

1 part of hydrogen combines with 92.3/7.7

= 12.0 parts of carbon

Thus, the masses of carbon in three compounds A, B and C, which combine with a fixed mass of hydrogen are in the ratio of 3 : 6 : 12 or 1 : 2 : 4. This is a simple ratio. Hence, the data illustrate the law of multiple proportions.

Example 6. *Two compounds each containing only tin and oxygen had the following composition:*

	Mass % of tin	*Mass % of oxygen*
Compound A	78.77	21.23
Compound B	88.12	11.88

Show how this data illustrate the law of multiple proportions?

Solution:

In Compound *A*

21.23 parts of oxygen combine with 78.77 parts of tin

1 part of oxygen combines with 78.77/21.23

= 3.7 parts of tin

In Compound *B*

11.88 parts of oxygen combine with 88.12 parts of tin

1 part of oxygen combines with 88.12/11.88

= 7.4 parts of tin

Thus, the mass of tin in compounds *A* and *B* which combine with a fixed mass of oxygen are in the ratio of 3.7 : 7.4 or 1 : 2. This is a simple ratio. Hence, the data illustrate the law of multiple proportions.

Example 7. *Illustrate the law of reciprocal proportions from the following data: KCl contains 52.0% potassium, KI contains 23.6% potassium and ICl contains 78.2% iodine.*

Solution: **In KCl:** Potassium 52.0%,

Chlorine $(100-52) = 48\%$

In KI: Potassium 23.6%;

Iodine $(100-23.6) = 76.4\%$

23.6 parts of potassium combine with 76.4 parts of iodine

52.0 parts of potassium will combine with

$(76.4/23.6) \times 52.0 = 168.3$ parts of iodine.

The ratio of masses of chlorine and iodine which combines with same mass of potassium = 48 : 168.3 or 1 : 3.5

In ICl: Iodine = 78.2% and chlorine

$= (100-78.2) = 21.8\%$

The ratio of chlorine and iodine in ICl = 21.8 : 78.2 = 1 : 3.5. Hence, the data illustrate the law of reciprocal proportions.

Example 8. *Zinc sulphate crystals contain 22.6% of zinc and 43.9% of water. Assuming the law of constant proportions to be true, how much zinc should be used to produce 13.7 g of zinc sulphate and how much water will they contain?*

Solution: 100 g of zinc sulphate crystals are obtained from

= 22.6 g zinc

1 g of zinc sulphate crystals will be obtained from

= 22.6 / 100 g zinc

13.7 g of zinc sulphate crystals will be obtained from

$$= \frac{22.6}{100} \times 13.7$$

= 3.0962 g of zinc

100 g of zinc sulphate crystals contain water

= 43.9 g

1 g of zinc sulphate crystals contain water

= 43.9 / 100 g

13.7 g of zinc sulphate crystals shall contain water

$$= \frac{43.9}{100} \times 13.7 = 6.0143\text{ g}$$

Example 9. *Carbon monoxide reacts with oxygen to form carbon dioxide according to the equation, $2CO + O_2 = 2CO_2$. In an experiment, 400 mL of carbon monoxide and 180 mL of oxygen were allowed to react, when 80% of carbon monoxide was transformed to carbon dioxide.*

All the volumes were measured under the same conditions of temperature and pressure. Find out the composition of the final mixture.

Solution:
$$\underset{2\text{ vol}}{2CO} + \underset{1\text{ vol}}{O_2} = \underset{2\text{ vol}}{2CO_2}$$

From the above equation, it is observed that volume of oxygen required for the transformation of carbon monoxide into carbon dioxide is half the volume of carbon monoxide and the volume of carbon dioxide produced is same as that of carbon monoxide.

Volume of carbon monoxide transformed

$$= \frac{80 \times 400}{100} = 320\text{ mL}$$

Hence, volume of oxygen required for transformation

$$= \frac{1}{2} \times 320 = 160\text{ mL}$$

Volume of carbon dioxide produced

= 320 mL

So, the composition of final mixture is

Carbon monoxide = (400 − 320)

= 80 mL

Carbon dioxide = 320 mL

Oxygen = 180 − 160 = 20 mL

Example 10. *How much volume of oxygen will be required for complete combustion of 40 mL of acetylene (C_2H_2) and how much volume of carbon dioxide will be formed? All volumes are measured at NTP.*

Solution:
$$\begin{array}{ccccccc} 2C_2H_2 & + & 5O_2 & = & 4CO_2 & + & 2H_2O \\ 2\text{ vol} & & 5\text{ vol} & & 4\text{ vol} & & \\ 40\text{ mL} & & \frac{5}{2} \times 40\text{ mL} & & \frac{4}{2} \times 40\text{ mL} & & \\ 40\text{ mL} & & 100\text{ mL} & & 80\text{ mL} & & \end{array}$$

So, for complete combustion of 40 mL of acetylene, 100 mL of oxygen are required and 80 mL of carbon dioxide is formed.

1.9 DALTON'S ATOMIC THEORY

The concept that matter is composed of very small particles was given by Indian and Greek philosophers. As early as 400 to 500 B.C. the Greek philosopher **Democritus** suggested that matter cannot be forever divided into smaller and smaller parts. The ultimate particles were considered as indivisible. These particles were called atoms. The word atom has been derived from the Greek word '*atomos*' meaning 'indivisible'. These early ideas, however, were not based on experiments but were mere speculations. The existence of atoms was accepted by Boyle in his book '*The Sceptical Chymist*' (1661) and by **Newton** in his books '*Principia*' and '*Opticks*' (1704). The old ideas were put on a scientific scale by **John Dalton** in the years 1803 to 1808 in the form of a theory known as **Dalton's Atomic Theory** which is a landmark in the history of chemistry. The main points of Dalton's atomic theory are:

(i) Elements consist of minute, indivisible, indestructible particles called atoms.
(ii) Atoms of an element are identical to each other. They have the same mass and size.
(iii) Atoms of different elements differ in properties and have different masses and sizes.
(iv) Compounds are formed when atoms of different elements combine with each other in simple numerical ratios such as one-to-one, one-to-two, two-to-three and so on.
(v) Atoms cannot be created, destroyed or transformed into atoms of other elements.
(vi) The relative numbers and kind of atoms are always the same in a given compound.

The theory convincingly explained the various laws of chemical combination, but the theory has undergone a complete shake up with the modern concept of structure of atom. However, the Daltonian atom still retains its significance as the unit participating in chemical reactions. The following are the modified views regarding Dalton's atomic theory:

(i) The atom is no longer supposed to be indivisible. The atom is not a simple particle but a complex one.
(ii) Atoms of the element may not necessarily possess the same mass but possess the same atomic number and show similar chemical properties **(Discovery of isotopes).**
(iii) Atoms of the different elements may possess the same mass but they always have different atomic numbers and differ in chemical properties **(Discovery of isobars).**
(iv) Atoms of one element can be transmuted into atoms of other element. **(Discovery of artificial transmutation).**
(v) In certain organic compounds, like proteins, starch, cellulose, etc., the ratio in which atoms of different elements combine cannot be regarded as simple. There are a number of compounds which do not follow the law of constant proportions. Such compounds are called **non-stoichiometric compounds.**

1.10 ATOMS, MOLECULES AND FORMULAE

An atom is the smallest particle of an element. The atom of hydrogen is the smallest and the lightest. Atoms take part in chemical combination and remain as indivisible. All atoms do not occur free in nature. **Avogadro** introduced the idea of another kind of particles called the molecules. **A molecule is the smallest particle of an element or compound that can have a stable and independent existence.** A molecule of an element consists of one or more atoms of the same element. Certain elements are capable of existence as single atoms and their atoms can be regarded as molecules. A molecule of an element that consists of one atom only is called monoatomic molecule as in the case of inert gases. Oxygen is not stable in atomic form but is stable in molecular form. A molecule of oxygen is diatomic in nature, *i.e.*, its molecule consists of two oxygen atoms. Hydrogen, nitrogen, fluorine, chlorine, bromine, iodine are also diatomic like oxygen. Some elements exist in more complex molecular forms. The molecule of phosphorus consists of four phosphorus atoms and the molecule of sulphur consists of eight sulphur atoms. Such molecules having more than two atoms are said to be polyatomic. A representation of the molecule of an element involves use of a subscript to the right of the elemental symbol. The diatomic molecule of chlorine is represented as Cl_2, whereas molecules of phosphorus and sulphur are represented as P_4 and S_8, respectively.

The molecule is the smallest possible unit of a compound which shows the properties of the compound. The molecules of all compounds contain two or more different types of atoms. These differ from the molecules of elements which contain only one type of atoms.

Thus, it becomes clear that atoms are the components of molecules and the molecules are components of elements or compounds.

The formula is a group of symbols of elements which represents one molecule of a substance. The formula of a substance represents its chemical composition. Water consists of molecules containing two hydrogen atoms and one oxygen atom which are represented as H_2O. The subscript to the right of the symbol for hydrogen indicates the number of hydrogen atoms contained in a molecule. No subscript follows the symbol for oxygen which means, by convention, that only one atom of oxygen is contained in the molecule.

The subscripts representing the number of atoms contained in a molecule of a compound are in no way related to the number of atoms present in the molecule of a free element. Although both hydrogen and oxygen are composed of diatomic molecules, a water molecule contains only one atom of oxygen and two atoms of hydrogen. The two hydrogen atoms present in H_2O are not molecular hydrogen but rather two hydrogen atoms that have chemically combined with an oxygen atom.

For a chemical formula to be correct, it must contain two pieces of information: (i) it must indicate the elements in the make up of the compound, and (ii) it must indicate the combining ratio of atoms of these elements in the particular compound. The first information is provided by including in the formula correct chemical symbols for all the elements in the compound. The second piece of information is provided by subscripts, *i.e.*, numbers written to the right slightly below the chemical symbols of the elements.

Nitric acid is a combination of hydrogen, nitrogen and oxygen giving a base formula HNO. These elements combine in the ratio 1 : 1 : 3. Therefore, the correct formula for nitric acid is HNO_3.

Some compounds are composed of ions rather than of molecules. Ions differ from atoms and molecules by being electrically charged particles of matter. The charges may be positive or negative and generally vary in magnitude. The positively charged ions are called cations and negatively charged ions are called anions. Simple cations and anions come into existence by loss and acceptance of an electron or electrons by neutral atoms respectively. Ions that consist of several atoms held together by chemical bonds similar to those involved in the molecules are called polyatomic ions or complex ions. These complex ions differ from molecules in the sense that they bear a charge. Some of the common complex ions are:

NO_3^-	Nitrate	PO_4^{3-}	Phosphate	NH_4^+	Ammonium
SO_4^{2-}	Sulphate	ClO_4^-	Perchlorate	PH_4^+	Phosphonium
SO_3^{2-}	Sulphite	CO_3^{2-}	Carbonate	MnO_4^-	Permanganate

When ions are present in a compound, the number of positive charges on a cation must balance with the negative charges on an anion to produce electrically neutral matter. Since, the charge on the anion may not always be equal to that on the cation, the number of anions will not always be equal to the number of cations.

Calcium nitrate consists of calcium and nitrate ions. Each calcium ion carries 2 units positive charge while each nitrate ion carries 1 unit negative charge. Thus, to make net charge zero, two nitrate ions will link with one calcium ion and the formula will be $Ca(NO_3)_2, [Ca^{2+} + 2NO_3^-]$. Names and formulae of some common chemical compounds are listed below :

Common Name	Chemical Name	Chemical Formula
Alum	Ammonium aluminium sulphate	$(NH_4)_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$
Aspirin	Acetyl salicylic acid	$C_9H_8O_4$
Battery acid or oil of vitriol	Sulphuric acid	H_2SO_4
Blue vitriol	Copper sulphate	$CuSO_4 \cdot 5H_2O$
Baking soda	Sodium bicarbonate	$NaHCO_3$
Bleaching powder	Calcium chlorohypochlorite	$CaOCl_2$
Borax	Sodium tetraborate	$Na_2B_4O_7 \cdot 10H_2O$
Butter of tin	Stannic chloride	$SnCl_4 \cdot 5H_2O$
Caustic soda	Sodium hydroxide	$NaOH$
Caustic potash	Potassium hydroxide	KOH
Carbolic acid	Phenol	C_6H_5OH
Chile saltpetre	Sodium nitrate	$NaNO_3$
Carborundum	Silicon carbide	SiC
Corrosive sublimate	Mercuric chloride	$HgCl_2$
Calomel	Mercurous chloride	Hg_2Cl_2
Dry ice	Carbon dioxide (solid)	CO_2
Formalin	Formaldehyde (40% solution)	$HCHO$
Grain alcohol (Spirit)	Ethyl alcohol	C_2H_5OH
Green vitriol	Ferrous sulphate	$FeSO_4 \cdot 7H_2O$
Gypsum	Calcium sulphate	$CaSO_4 \cdot 2H_2O$
Gammexane (BHC)	Benzene hexachloride	$C_6H_6Cl_6$
Hydrolith	Calcium hydride	CaH_2
Hypo (Antichlor)	Sodium thiosulphate	$Na_2S_2O_3 \cdot 5H_2O$
Indian nitre	Potassium nitrate	KNO_3
Limestone	Calcium carbonate	$CaCO_3$
Lunar caustic	Silver nitrate	$AgNO_3$
Laughing gas	Nitrous oxide	N_2O
Litharge	Lead monoxide	PbO
Muratic acid	Hydrochloric acid	HCl
Mohr's salt	Ferrous ammonium sulphate	$FeSO_4(NH_4)_2SO_4 \cdot 6H_2O$
Milk of magnesia	Magnesium hydroxide	$Mg(OH)_2$
Microcosmic salt	Sodium ammonium hydrogen ortho-phosphate	$Na(NH_4)HPO_4$
Marsh gas (Damp fire)	Methane	CH_4
Oleum	Sulphuric acid (Fuming)	$H_2S_2O_7$
Oxone	Sodium peroxide	Na_2O_2
Plaster of Paris	Calcium sulphate hemihydrate	$CaSO_4 \cdot \frac{1}{2}H_2O$
Philosphers's wool	Zinc oxide	ZnO
Phosgene	Carbonyl chloride	$COCl_2$
Pearl ash	Potassium carbonate	K_2CO_3
Pyrene	Carbon tetrachloride	CCl_4
Picric acid	2,4,6-Trinitrophenol	$C_6H_2(OH)(NO_2)_3$
Quick lime	Calcium oxide	CaO
Red lead (Minium)	Lead tetroxide	Pb_3O_4
Sugar	Sucrose	$C_{12}H_{22}O_{11}$
Slaked lime (Milk of lime)	Calcium hydroxide	$Ca(OH)_2$
Sal ammoniac	Ammonium chloride	NH_4Cl
Sugar of lead	Lead acetate	$(CH_3COO)_2Pb$
Sand	Silicon dioxide	SiO_2
Table salt (Common salt)	Sodium chloride	$NaCl$
TEL	Tetra-ethyl lead	$Pb(C_2H_5)_4$
Tear gas	Chloropicrin	CCl_3NO_2
Washing soda	Sodium carbonate	$Na_2CO_3 \cdot 10H_2O$
Water glass	Sodium silicate	Na_2SiO_3
White vitriol	Zinc sulphate	$ZnSO_4 \cdot 7H_2O$

1.11 ATOMIC AND MOLECULAR MASS

One of the most important concepts derived from Dalton's atomic theory is that of atomic mass, *i.e.*, each element has a characteristic atomic mass. As atoms are very tiny particles, their absolute masses are difficult to measure. However, it is possible to determine the relative masses of different atoms if a small unit of mass is taken as a standard. For this purpose, mass of one atom of hydrogen was assumed as unity and was accepted as standard. **The atomic mass of an element can be defined as the number which indicates how many times the mass of one atom of the element is heavier in comparison to the mass of one atom of hydrogen.**

$$A = \text{Atomic mass of an element}$$
$$= \frac{\text{Mass of one atom of the element}}{\text{Mass of one atom of hydrogen}}$$

In 1858, oxygen atom was adopted as a standard on account of the following reasons:

(i) It is much easier to obtain compounds of elements with oxygen than with hydrogen as oxygen is more reactive than hydrogen.

(ii) The atomic masses of most of the elements become approximately whole numbers but with hydrogen as standard the atomic masses of most of the elements are fractional.

The mass of one atom of natural oxygen was taken to be 16.0.

Thus, atomic mass of an element

$$= \frac{\text{Mass of one atom of the element}}{\frac{1}{16}\text{th part of the mass of one atom of oxygen}}$$
$$= \frac{\text{Mass of one atom of the element}}{\text{Mass of one atom of oxygen}} \times 16$$

By accepting oxygen as a standard, the atomic mass of hydrogen comes as 1.008, sodium 22.991 and sulphur 32.066.

In 1961, the International Union of Chemists selected a new unit for expressing the atomic masses. They accepted the stable isotope of carbon (^{12}C) with mass number of 12 as the standard. **Atomic mass of an element can be defined as the number which indicates how many times the mass of one atom of the element is heavier in comparison to $\frac{1}{12}$th part of the mass of one atom of carbon-12 (^{12}C).**

$$A = \text{Atomic mass of an element}$$
$$= \frac{\text{Mass of one atom of the element}}{\frac{1}{12}\text{th part of the mass of one atom of carbon-12}}$$
$$= \frac{\text{Mass of one atom of the element}}{\text{Mass of one atom of carbon-12}} \times 12$$

[The quantity 'A' was formerly known as atomic weight. However, this term is no longer used as the word 'weight' means gravitational force.]

Atomic mass unit: The quantity $\frac{1}{12}$ mass of an atom of carbon-12 (^{12}C) is known as the atomic mass unit and is abbreviated as amu. The actual mass of one atom of carbon-12 is 1.9924×10^{-23} g or 1.9924×10^{-26} kg.

*The term Dalton is used for one atomic mass unit, 1 Dalton = 1 amu.

Thus,

$$1\,\text{amu}^* = \frac{1.9924 \times 10^{-23}}{12} = 1.66 \times 10^{-24}\text{ g or } 1.66 \times 10^{-27}\text{ kg}$$

$$A = \text{Atomic mass of an element}$$
$$= \frac{\text{Mass of one atom of the element}}{1\,\text{amu}}$$

The atomic masses of some elements on the basis of carbon-12 are given below:

Hydrogen	1.008 amu	Iron	55.847 amu
Oxygen	16.00 amu	Sodium	22.989 amu
Chlorine	35.453 amu	Zinc	65.38 amu
Magnesium	24.305 amu	Silver	107.868 amu
Copper	63.546 amu		

The actual mass of an atom of an element

= The atomic mass of an element in amu $\times 1.66 \times 10^{-24}$ g

So, the actual mass of hydrogen atom

$$= 1.008 \times 1.66 \times 10^{-24} = 1.6736 \times 10^{-24}\text{ g}$$

Similarly, the actual mass of oxygen atom

$$= 16 \times 1.66 \times 10^{-24} = 2.656 \times 10^{-23}\text{ g}$$

It is clear from the above list of atomic masses that atomic masses of a number of elements are not nearly whole numbers. Actually, the above values are average relative masses. Most of the elements occur in nature as a mixture of isotopes. (Isotopes—the atoms of the same element having different atomic masses). With very few exceptions, however, elements have constant mixtures of isotopes. Chlorine is found in nature as a mixture containing two isotopes Cl-35 (34.969 amu) and Cl-37 (36.966 amu). These are found in the ratio of 75.53% (Cl-35) and 24.47% (Cl-37). Therefore, the average relative mass of chlorine is calculated as:

$$(34.969 \times 0.7553) + (36.966 \times 0.2447) = 35.46\text{ amu}$$

Based on the average mass, the atomic mass of chlorine is 35.46 or 35.5 amu but it is never possible to have an atom having a relative mass 35.5 amu. It can have relative mass of about 35.0 or 37.0 amu depending on the particular isotope. Thus, average relative mass of any naturally occurring sample of chlorine is 35.46 or 35.5 amu as it is a mixture of two isotopes present in definite proportion. The same reasoning applies to all other elements.

The average atomic masses of various elements are determined by multiplying the atomic mass of each isotope by its fractional abundance and adding the values thus obtained. The fractional abundance is determined by dividing percentage abundance by hundred.

$$\text{Average isotopic mass} = \frac{m \times a + n \times b}{m + n}$$

here, a, b are atomic masses of isotopes in the ratio $m : n$.

$$\text{Average isotopic mass} = \frac{x}{100} \times a + \frac{y}{100} \times b$$

here, x, y are percentage abundance of the two isotopes $(y = 100 - x)$.

Example 11. *Boron has two isotopes boron-10 and boron-11 whose percentage abundances are 19.6% and 80.4% respectively. What is the average atomic mass of boron?*

Solution:

$$\text{Contribution of boron-10} = 10.0 \times 0.196 = 1.96 \text{ amu}$$

$$\text{Contribution of boron-11} = 11.0 \times 0.804 = 8.844 \text{ amu}$$

$$\text{Adding both} = 1.96 + 8.844 = 10.804 \text{ amu}$$

Thus, the average atomic mass of boron is 10.804 amu.

Example 12. *Carbon occurs in nature as a mixture of carbon-12 and carbon-13. The average atomic mass of carbon is 12.011. What is the percentage abundance of carbon-12 in nature?*

Solution: Let x be the percentage abundance of carbon-12; then $(100 - x)$ will be the percentage abundance of carbon-13.

Therefore, $$\frac{12x}{100} + \frac{13(100-x)}{100} = 12.011$$

or $$12x + 1300 - 13x = 1201.1$$

or $$x = 98.9$$

Abundance of carbon-12 is 98.9%.

Gram-atomic Mass or Gram Atom

When numerical value of atomic mass of an element is expressed in grams, the value becomes gram-atomic mass or gram atom. The atomic mass of oxygen is 16 while gram-atomic mass or gram atom of oxygen is 16 g. Similarly, the gram-atomic masses of hydrogen, chlorine and nitrogen are 1.008 g, 35.5 g and 14.0 g respectively. **Gram-atomic mass or gram atom of every element consists of same number of atoms.** This number is called **Avogadro's number.** The value of Avogadro's number is 6.02×10^{23}.

Absolute mass of one oxygen atom

$$= 16 \text{ amu} = 16 \times 1.66 \times 10^{-24} \text{ g}$$

Therefore, the mass of 6.02×10^{23} atoms of oxygen will be

$$= 16 \times 1.66 \times 10^{-24} \times 6.02 \times 10^{23}$$

$$= 16 \text{ g (gram-atomic mass)}$$

Thus, gram-atomic mass can be defined as the absolute mass in grams of 6.02×10^{23} atoms of any element.

Number of gram atoms of any element can be calculated with the help of the following formula:

$$\text{No. of gram atoms} = \frac{\text{Mass of the element in grams}}{\text{Atomic mass of the element in grams}}$$

Molecular Mass

Like an atom, a molecule of a substance is also a very small particle possessing a mass of the order of 10^{-24} to 10^{-22} g. Similar to atomic mass, molecular mass is also expressed as a relative mass with respect to the mass of the standard substance which is an atom of hydrogen or an atom of oxygen or an atom of carbon-12. The molecular mass of a substance may be defined as the mass of a molecule of a substance relative to the mass of an atom of hydrogen as 1.008 or of oxygen taken as 16.00 or the mass of one atom of carbon taken as 12. **Molecular mass is a number which indicates how many times one molecule of a substance is heavier in comparison to $\frac{1}{16}$th of the mass of oxygen atom or $\frac{1}{12}$th of the mass of one atom of carbon-12.**

$$M = \text{Molecular mass}$$
$$= \frac{\text{Mass of one molecule of the substance}}{\frac{1}{12}\text{th mass of one atom of carbon-12}}$$

The mass of a molecule is equal to sum of the masses of the atoms present in a molecule. One molecule of water consists of 2 atoms of hydrogen and one atom of oxygen. Thus, molecular mass of water $= (2 \times 1.008) + 16.00 = 18.016$ amu. One molecule of H_2SO_4 (sulphuric acid) consists of 2 atoms of hydrogen, one atom of sulphur and four atoms of oxygen. Thus, the molecular mass of sulphuric acid is

$$= (2 \times 1.008) + 32.00 + (4 \times 16.00)$$
$$= 98.016 \text{ or } 98.016 \text{ amu}$$

Gram-molecular Mass or Gram Molecule

A quantity of substance whose mass in grams is numerically equal to its molecular mass is called gram-molecular mass. In other words, molecular mass of a substance expressed in grams is called gram-molecular mass or gram molecule. For example, the molecular mass of chlorine is 71 and, therefore, its gram-molecular mass or gram molecule is 71 g.

Similarly, molecular mass of oxygen (O_2) is 32, *i.e.*, $2 \times 16 = 32$ amu.

$$\text{Gram-molecular mass of oxygen} = 32 \text{ g}$$

Molecular mass of nitric acid (HNO_3) is 63, *i.e.*,

$$= 1 + 14 + 3 \times 16 = 63 \text{ amu}$$

$$\text{Gram-molecular mass of nitric acid} = 63 \text{ g}$$

Gram-molecular mass should not be confused with the mass of one molecule of the substance in grams. The mass of one molecule of a substance is known as its **actual mass.** For example, the actual mass of one molecule of oxygen is equal to $32 \times 1.66 \times 10^{-24}$ g, *i.e.*, 5.32×10^{-23} g.

The number of gram molecules of a substance present in a given mass of a substance can be determined by the application of following formula:

No. of gram molecules

$$= \frac{\text{Mass of a substance in grams}}{\text{Molecular mass of the substance in grams}}$$

$$\text{Mass of single molecule} = \frac{\text{Molar mass in grams}}{6.023 \times 10^{23}}$$

$$= \text{Molar mass in amu} \times 1.66 \times 10^{-24} \text{ grams}$$

Example 13. *Calculate the mass of 2.5 gram atoms of oxygen.*

Solution: We know that,

$$\text{No. of gram atoms} = \frac{\text{Mass of the element in grams}}{\text{Atomic mass of the element in grams}}$$

So, Mass of oxygen $= 2.5 \times 32 = 80.0\,g$

Example 14. *Calculate the gram atoms in 2.3 g of sodium.*

Solution: No. of gram atoms $\frac{2.3}{23} = 0.1$

[Atomic mass of sodium = 23 g]

Example 15. *Calculate the mass of 1.5 gram molecule of sulphuric acid.*

Solution: Molecular mass of

$H_2SO_4 = 2 \times 1 + 32 + 4 \times 16 = 98.0\,amu$

Gram-molecular mass of $H_2SO_4 = 98.0\,g$

Mass of 1.5 gram molecule of $H_2SO_4 = 98.0 \times 1.5 = 147.0\,g$

Example 16. *Calculate the actual mass of one molecule of carbon dioxide (CO_2).*

Solution: Molecular mass of CO_2 = 44 amu

$1\,amu = 1.66 \times 10^{-24}\,g$

So, The actual mass of $CO_2 = 44 \times 1.66 \times 10^{-24}$

$= 7.304 \times 10^{-23}\,g$

1.12 AVOGADRO'S HYPOTHESIS

According to Dalton's atomic theory, elements react with each other in the simple ratio of their atoms. Gay-Lussac proposed that gases combine in simple ratio of their volumes. In an attempt to correlate Dalton's atomic theory with Gay-Lussac law of gaseous volumes, **Berzelius** stated that **under similar conditions of temperature and pressure, equal volume of all gases contain the same number of atoms.** This hypothesis was subsequently found to be incorrect as it failed to interpret the experimental results and contradicted the very basic assumption of Dalton's atomic theory, *i.e.*, an atom is indivisible. For example, the formation of hydrogen chloride from hydrogen and chlorine could not be explained on the basis of Berzelius hypothesis.

Hydrogen +	Chlorine =	Hydrogen chloride
1 vol	1 vol	2 vol
n atoms	n atoms	$2n$ compound atoms
1 atom	1 atom	2 compound atoms
½ atom	½ atom	1 compound atom

i.e., for the formation of 1 compound atom of hydrogen chloride, ½ atom of hydrogen and ½ atom of chlorine are needed. In other words, each atom of hydrogen and chlorine has been divided which is against Dalton's atomic theory. Thus, the hypothesis of Berzelius was discarded.

The Italian scientist, **Amedeo Avogadro,** in 1811, solved the above problem by proposing two types of particles from which whole of the matter is composed of.

(i) Atom: The smallest particle of an element that can take part in chemical change but generally cannot exist freely as such.

(ii) Molecule: The smallest particle of a substance (element or compound) which has free or independent existence and possesses all characteristic properties of the substance. A molecule of an element is composed of like atoms while a molecule of a compound contains fixed number of atoms of two or more different elements. A molecule may be broken down into its constituent atoms but the atom is indivisible during a chemical change.

Avogadro after making the above differentiation, presented a hypothesis known as **Avogadro hypothesis** which can be stated as follows:

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

Avogadro hypothesis explains successfully the formation of hydrogen chloride.

Hydrogen +	Chlorine =	Hydrogen chloride
1 vol	1 vol	2 vol
n molecules	n molecules	$2n$ molecules
1 molecule	1 molecule	2 molecules
$\frac{1}{2}$ molecule	$\frac{1}{2}$ molecule	1 molecule
1 atom	1 atom	1 molecule

(Both hydrogen and chlorine are diatomic in nature.)

Thus, the hypothesis explains that the molecules of reacting gases break up into constituent atoms during chemical change which then combine to form new molecules of the product or products.

Applications of Avogadro's hypothesis

(i) Atomicity*: Atomicity means number of atoms present in one molecule of an elementary gas. Hydrogen, oxygen, nitrogen, chlorine, etc., are diatomic in nature. Noble gases are monoatomic while ozone is triatomic in nature. Avogadro's hypothesis helps in determining the atomicity of elements.

(ii) Relationship between molecular mass and vapour density: The vapour density of any gas is the ratio of the densities of the gas and hydrogen under similar conditions of temperature and pressure.

$$\text{Vapour Density (V.D.)} = \frac{\text{Density of gas}}{\text{Density of hydrogen}}$$

$$= \frac{\text{Mass of a certain volume of the gas}}{\text{Mass of same volume of hydrogen at the same temp. and pressure}}$$

If n molecules are present in the given volume of a gas and hydrogen under similar conditions of temperature and pressure,

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

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

*Atomicity can be ascertained with the values of ratio of two specific heats of gases $\left(\frac{C_P}{C_V}\right)$

$\frac{C_P}{C_V} = 1.66$ (Monoatomic), $\frac{C_P}{C_V} = 1.40$ (Diatomic), $\frac{C_P}{C_V} = 1.33$ (Polyatomic)

$$= \frac{\text{Molecular mass of gas}}{\text{Molecular mass of hydrogen}}$$

$$= \frac{\text{Mol. mass}}{2}$$

(since, mol. mass of hydrogen $= 2$)

Hence, $2 \times \text{V.D.} = \text{Mol. mass}$

This formula can be used for the determination of molecular masses of volatile substances from vapour density. Vapour density is measured mainly by two methods:

(a) Victor Meyer and (b) Duma's methods.

(iii) Gram-molecular volume: 1 g mole of any gas occupies 22.4 litres or 22400 mL of volume at NTP or STP conditions.*

The density of hydrogen at NTP is 0.00009 g mL^{-1}. Thus, 0.00009 g of hydrogen will occupy volume at NTP = 1 mL

$$1 \text{ g of hydrogen occupies volume at NTP} = \frac{1}{0.00009} \text{ mL}$$

1 g mole of hydrogen (2.016 g) occupies volume at NTP

$$= \frac{2.016}{0.00009} = 22400 \text{ mL} = 22.4 \text{ litre}$$

According to Avogadro's hypothesis, equal volumes of different gases contain same number of molecules under similar conditions of temperature and pressure. Thus, 22.4 litre or 22400 mL of any gas at NTP will contain one gram mole or its molecular mass in grams.

Loschmidt number : Number of molecules in 1 cm^3 or 1 mL of a gas at S.T.P. is known as Loschmidt number.

$$\text{Loschmidt number} = \frac{6.023 \times 10^{23}}{22400}$$

$$= 2.68 \times 10^{18} \text{ molecules mL}^{-1}$$

(iv) Molecular formula: Avogadro's hypothesis helps in finding the molecular formulae of gases. Under similar conditions of temperature and pressure, 2 volumes of ozone after decomposition give 3 volumes of oxygen.

$$\underset{2\text{ vol}}{\text{Ozone}} \xrightarrow{\text{Decomposition}} \underset{3\text{ vol}}{\text{Oxygen}}$$

2 molecules	3 molecules
1 molecule	3/2 molecules
1 molecule	3 atoms

Thus, the formula of ozone is O_3.

1.13 MOLE CONCEPT

For the counting of articles, the unit **dozen** or unit **gross** is commonly used irrespective of their nature. For example, one dozen pencils means 12 pencils or one dozen apples means 12 apples or one gross books means 144 books or one gross oranges means 144 oranges. In a similar way, for counting of atoms, molecules, ions, etc., chemists use the unit mole. The term mole was introduced by **Ostwald** in 1896. This is the Latin word 'moles' meaning heap or pile. **A mole (mol) is defined as the number of atoms in 12.00 g of carbon-12.** The number of atoms in 12 g of carbon-12 has been found experimentally to be 6.02×10^{23}. This number is also known as Avogadro's number named in honour of **Amedeo Avogadro** (1776 – 1856).

Thus, a mole contains 6.02×10^{23} units. These units can be atoms, molecules, ions, electrons or anything else.

1 mole of hydrogen atoms means 6.02×10^{23} hydrogen atoms.

1 mole of hydrogen molecules means 6.02×10^{23} hydrogen molecules.

1 mole of potassium ions means 6.02×10^{23} potassium ions.

1 mole of electrons means 6.02×10^{23} electrons.

The type of entity must be specified when the mole designation is used. A mole of oxygen atoms contains 6.02×10^{23} oxygen atoms and a mole of oxygen molecules contains 6.02×10^{23} oxygen molecules. Therefore, a mole of oxygen molecules is equal to two moles of oxygen atoms, *i.e.*, $2 \times 6.02 \times 10^{23}$ oxygen atoms.

How much does one mole weigh? That depends on the nature of particles (units). **The mass of one mole atoms of any element is exactly equal to the atomic mass in grams (gram-atomic mass or gram atom) of that element.**

For example, the atomic mass of aluminium is 27 amu. One amu is equal to 1.66×10^{-24} g. One mole of aluminium contains 6.02×10^{23} aluminium atoms.

$$\text{Mass of one atom aluminium} = 27 \times 1.66 \times 10^{-24} \text{ g}$$

$$\text{Mass of one mole aluminium} = 27 \times 1.66 \times 10^{-24} \times 6.02 \times 10^{23}$$

$$= 27 \text{ g}$$

This is the atomic mass of aluminium in grams or it is one gram atomic mass or one gram atom of aluminium.

Similarly, **the mass of 6.02×10^{23} molecules (1 mole) of a substance is equal to its molecular mass in grams or gram-molecular mass or gram molecule.** For example, molecular mass of water is 18 amu. Thus, mass of one mole of water will be $18 \times 1.66 \times 10^{-24} \times 6.02 \times 10^{23}$, *i.e.*, 18 g. This is the molecular mass of water in grams or one gram-molecular mass or one gram molecule.

Mole concept is also applicable to ionic compounds which do not contain molecules. In such cases, the formula of an ionic compound represents the ratio between constituent ions. The mass of 6.02×10^{23} formula units represents one mole of an ionic compound.

* 0°C or 273 K temperature and one atmosphere or 760 mm of Hg or 76 cm of Hg pressure are known as the standard conditions of temperature and pressure (STP) or normal conditions of temperature and pressure (NTP).

One mole of $BaCl_2 = 6.02 \times 10^{23}$ $BaCl_2$ units

$= 208.2$ g $BaCl_2$

= Molecular mass (formula mass) of $BaCl_2$

$= 6.02 \times 10^{23}$ Ba^{2+} ions $+ 2 \times 6.02 \times 10^{23}$ Cl^- ions

$= 137.2 + 71.0 = 208.2$ g

One mole of a substance will have mass equal to formula mass of that substance expressed in grams.

It has been established by Avogadro's hypothesis that one gram-molecular mass of any gaseous substance occupies a volume of 22.4 litres at NTP. One gram-molecular mass is nothing but one mole of substance. Thus, one mole, *i.e.*, 6.02×10^{23} molecules of any gaseous substance occupies 22.4 litres as volume at NTP.

The following formulae satisfy the above discussion.

1 mole of a substance $= 6.02 \times 10^{23}$ particles of the substance

Number of moles of a substance

$$= \frac{\text{Mass of substance in gram}}{\text{Mass of one mole of the substance in gram}}$$

Further, $\quad \text{Number of moles} = \dfrac{\text{No. of particles}}{6.02 \times 10^{23}}$

Thus,

$$\frac{\text{No. of particles}}{6.02 \times 10^{23}} = \frac{\text{Mass of substance in gram}}{\text{Mass of one mole of the substance in gram}}$$

Mass of one atom of an element

$$= \frac{\text{Gram atom of an element}}{6.02 \times 10^{23}}$$

Mass of one molecule of a substance

$$= \frac{\text{Gram-molecular mass of the substance}}{6.02 \times 10^{23}}$$

Number of molecules

$$= \frac{\text{Volume of gas in litres at NTP}}{22.4} \times 6.02 \times 10^{23}$$

SOME SOLVED EXAMPLES

Example 17. *A piece of copper weighs* 0.635 g. *How many atoms of copper does it contain?* **[CEE (Bihar) 1992]**

Solution: Gram-atomic mass of copper = 63.5 g

$$\text{Number of moles in 0.635 g of copper} = \frac{0.635}{63.5} = 0.01$$

Number of copper atoms in one mole $= 6.02 \times 10^{23}$

Number of copper atoms in 0.01 moles $= 0.01 \times 6.02 \times 10^{23}$

$= 6.02 \times 10^{21}$

Example 18. *How many molecules of water and oxygen atoms are present in* 0.9 g *of water?*

Solution: Gram-molecular mass of water = 18 g

$$\text{Number of moles in 0.9 g of water} = \frac{0.9}{18} = 0.05$$

Number of water molecules in one mole of water

$= 6.02 \times 10^{23}$

Number of molecules of water in 0.05 moles

$= 0.05 \times 6.02 \times 10^{23}$

$= 3.010 \times 10^{22}$

As one molecule of water contains one oxygen atom,

So, number of oxygen atoms in 3.010×10^{22} molecule of water $= 3.010 \times 10^{22}$

Example 19. *Calculate the mass of a single atom of sulphur and a single molecule of carbon dioxide.*

Solution:

Gram-atomic mass of sulphur = 32 g

$$\text{Mass of one sulphur atom} = \frac{\text{Gram-atomic mass}}{6.02 \times 10^{23}}$$

$$= \frac{32}{6.02 \times 10^{23}} = 5.33 \times 10^{-23} \text{ g}$$

Formula of carbon dioxide $= CO_2$

Molecular mass of $CO_2 = 12 + 2 \times 16 = 44$

Gram-molecular mass of $CO_2 = 44$ g

$$\text{Mass of one molecule of } CO_2 = \frac{\text{Gram-molecular mass}}{6.02 \times 10^{23}}$$

$$= \frac{44}{6.02 \times 10^{23}} = 7.308 \times 10^{-23} \text{ g}$$

Example 20. *What is the mass of* 3.01×10^{22} *molecules of ammonia?*

Solution: Gram-molecular mass of ammonia = 17 g

Number of molecules in 17 g (one mole) of $NH_3 = 6.02 \times 10^{23}$

Let the mass of 3.01×10^{22} molecules of NH_3 be $= x$ g

So, $$\frac{3.01 \times 10^{22}}{6.02 \times 10^{23}} = \frac{x}{17}$$

or $$x = \frac{17 \times 3.01 \times 10^{22}}{6.02 \times 10^{23}} = 0.85 \text{ g}$$

Example 21. *From* 200 *mg of* CO_2, 10^{21} *molecules are removed. How many moles of* CO_2 *are left?*

Solution:

Gram-molecular mass of $CO_2 = 44$ g

$$\text{Mass of } 10^{21} \text{ molecules of } CO_2 = \frac{44}{6.02 \times 10^{23}} \times 10^{21} = 0.073 \text{ g}$$

Mass of CO_2 left $= (0.2 - 0.073) = 0.127$ g

$$\text{Number of moles of } CO_2 \text{ left} = \frac{0.127}{44} = 2.88 \times 10^{-3}$$

Example 22. *How many molecules and atoms of oxygen are present in 5.6 litres of oxygen (O_2) at NTP?*

Solution: We know that, 22.4 litres of oxygen at NTP contain 6.02×10^{23} molecules of oxygen.

So, 5.6 litres of oxygen at NTP contain

$$= \frac{5.6}{22.4} \times 6.02 \times 10^{23} \text{ molecules}$$

$$= 1.505 \times 10^{23} \text{ molecules}$$

1 molecule of oxygen contains = 2 atoms of oxygen

So, 1.505×10^{23} molecules of oxygen contain

$$= 2 \times 1.505 \times 10^{23} \text{ atoms}$$

$$= 3.01 \times 10^{23} \text{ atoms}$$

Example 23. *How many electrons are present in 1.6 g of methane?*

Solution: Gram-molecular mass of methane,

$$(CH_4) = 12 + 4 = 16 \text{ g}$$

Number of moles in 1.6 g of methane

$$= \frac{1.6}{16} = 0.1$$

Number of molecules of methane in 0.1 mole

$$= 0.1 \times 6.02 \times 10^{23}$$

$$= 6.02 \times 10^{22}$$

One molecule of methane has $= 6 + 4 = 10$ electrons

So, 6.02×10^{22} molecules of methane have

$$= 10 \times 6.02 \times 10^{22} \text{ electrons}$$

$$= 6.02 \times 10^{23} \text{ electrons}$$

Example 24. *The electric charge on the electron is 1.602×10^{-19} coulomb. How much charge is present on 0.1 mole of Cu^{2+} ions?*

Solution: Charge on one mole of electrons

$$= 6.02 \times 10^{23} \times 1.602 \times 10^{-19} \text{ coulomb}$$

$$= 96500 \text{ coulomb} = 1 \text{ faraday}$$

Charge on one mole of Cu^{2+} ions

$$= 2 \times 96500 \text{ coulomb} = 2 \text{ faraday}$$

Charge on 0.1 mole of Cu^{2+} ions

$$= 0.1 \times 2 = 0.2 \text{ faraday}$$

Example 25. *How many years it would take to spend one Avogadro's number of rupees at a rate of 10 lakh of rupees in one second?* **(MLNR 1990)**

Solution: Number of rupees spent in one second $= 10^6$

Number of rupees spent in one year

$$= 10^6 \times 60 \times 60 \times 24 \times 365$$

Avogadro's number of rupees will be spent

$$= \frac{6.02 \times 10^{23}}{10^6 \times 60 \times 60 \times 24 \times 365}$$

$$= 19.089 \times 10^9 \text{ years} = 1.9089 \times 10^{10} \text{ years}$$

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

1. 116 mg of a compound on vaporisation in Victor Meyer's apparatus displaces 44.8 mL of air measured at STP. The molecular mass of the compound is: **[CEE (Kerala) 2004]**
(a) 116 (b) 232 (c) 58 (d) 44.8 (e) 46.4
[Ans. (c)]
[Hint: Molar mass of compound
$$= \text{Mass of 22400 mL vapour at STP}$$
$$= \frac{0.116 \times 22400}{44.8} = 58]$$

2. A gas has a vapour density 11.2. The volume occupied by 1 g of the gas at NTP is: **(JCECE 2004)**
(a) 1 L (b) 11.2 L (c) 22.4 L (d) 4 L
[Ans. (a)]
[Hint: Molar mass $= 2 \times 11.2 = 22.4$ g
Volume of 1 g compound at STP $= \frac{22.4}{22.4} = 1$ L]

3. 3 g of hydrocarbon on combustion with 11.2 g of oxygen produce 8.8 g of CO_2 and 5.4 g of H_2O. The data illustrate the law of:
(a) conservation of mass (b) multiple proportions
(c) constant proportions (d) reciprocal proportions
[Ans. (a)]
[Hint: Σ Masses of reactants = Σ Masses of products
(3 + 11.2) g (8.8 + 5.4) g
Hence, law of conservation of mass is verified.]

4. The maximum number of molecules is present in: **[CBSE (PMT) 2004; Manipal (Medical) 2007]**
(a) 15 L of H_2 gas at STP (b) 5 L of N_2 gas at STP
(c) 0.5 g of H_2 gas (d) 10 g of O_2 gas
[Ans. (a)]
[Hint:
Number of molecules in 15 L $H_2 = \frac{15}{22.4} \times N = 0.669\,N$
Number of molecules in 5 L $N_2 = \frac{5}{22.4} \times N = 0.223\,N$
Number of molecules in 0.5 g $H_2 = \frac{0.5}{2} \times N = 0.25\,N$
Number of molecules in 10 g $O_2 = \frac{10}{32} \times N = 0.312\,N$]

5. Insulin contains 3.4% sulphur. Then, the minimum molecular mass of the insulin is about:
(a) 940 amu (b) 9400 amu
(c) 3600 amu (d) 970 amu
[Ans. (a)]
[Hint: $\because$ 3.4 g sulphur is present in 100 g insulin
$\therefore$ 32 g sulphur will be present in $\frac{100}{3.4} \times 32$ g insulin = 940
$\therefore$ Molar mass of insulin is about 940 amu]

6. 25 g of MCl_4 contains 0.5 mol chlorine then its molecular mass is: **(DPMT 2007)**
(a) 100g mol^{-1} (b) 200g mol^{-1}
(c) 150g mol^{-1} (d) 400g mol^{-1}
[**Ans.** (b)]
[**Hint:** 1 mol of MCl_4 contains 4 mol of chlorine
$\because$ 0.5 mol chlorine is present in 25 g of MCl_4
$\therefore$ 4 mol chlorine will be present in $\frac{25}{0.5} \times 4$, *i.e.*, 200 g of MCl_4.]

1.14 EQUIVALENT MASSES OR CHEMICAL EQUIVALENTS

Equivalent mass of a substance (element or compound) is defined as **the number of parts by mass of the substance which combine or displace 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 or 108 parts by mass of silver.**

The equivalent mass is a pure number. When the equivalent mass of a substance is expressed in grams, it is called gram equivalent mass. For example, equivalent mass of sodium is 23, hence, its gram equivalent mass is 23 g.

The equivalent mass of a substance may have different values under different conditions. The equivalent mass of an element may vary with change of valency. For example, copper forms two oxides CuO and Cu_2O. In CuO, 63.5 parts of copper combine with 16 parts of oxygen. Thus, equivalent mass of copper in this oxide is $\frac{63.5}{2} = 31.75$. In Cu_2O, 2×63.5 parts of copper combine with 16 parts of oxygen; thus, the equivalent mass of copper in this oxide is:

$$\frac{2 \times 63.5}{2} = 63.5$$

Relation between atomic mass, equivalent mass and valency: Suppose an element X combines with hydrogen to form a compound, XH_n, where n is the valency of the element X.

n parts by mass of hydrogen combine with atomic mass of element X.

1 part by mass of hydrogen combines with

$$\frac{\text{Atomic mass of element}}{n}$$

By above definition, $\frac{\text{Atomic mass of element}}{n}$ is the equivalent mass of the element.

Thus, $$\text{Equivalent mass} = \frac{\text{Atomic mass}}{n}$$

or $$\text{Atomic mass} = \text{Equivalent mass} \times \text{Valency}$$

Note: Detailed discussion on equivalent masses of compounds (acids, bases, salts, oxidising agents, reducing agents, etc.,) will be taken in chapter on volumetric analysis.

The following methods are used for the determination of equivalent mass of elements.

(i) Hydrogen displacement method: This method is used for those elements which can evolve hydrogen from acids, *i.e.*, active metals. A known mass of the active metal is reacted with dilute mineral acid. Hydrogen gas thus evolved is measured under experimental conditions. The volume of hydrogen is then reduced to NTP conditions. The mass of liberated hydrogen is determined using density of hydrogen (0.00009 at NTP).

$$\text{Equivalent mass} = \frac{\text{Mass of element}}{\text{Mass of hydrogen}} \times 1.008$$

$$= \frac{\text{Mass of element} \times 1.008}{\text{Volume in mL of hydrogen displaced at NTP} \times 0.00009}$$

$$= \frac{\text{Mass of element} \times 11200}{\text{Volume in mL of hydrogen displaced at NTP}}$$

(ii) Oxide formation method: A known mass of the element is changed into oxide directly or indirectly. The mass of oxide is noted.

$$\text{Mass of oxygen} = (\text{Mass of oxide} - \text{Mass of element})$$

Thus, the equivalent mass of the element

$$= \frac{\text{Mass of element}}{(\text{Mass of oxide} - \text{Mass of element})} \times 8$$

$$= \frac{\text{Mass of element}}{\text{Mass of oxygen}} \times 8$$

(iii) Chloride formation method: A known mass of the element is changed into chloride directly or indirectly. The mass of the chloride is determined.

$$\text{Mass of chlorine} = (\text{Mass of chloride} - \text{Mass of element})$$

Thus, the equivalent mass of the element

$$= \frac{\text{Mass of element} \times 35.5}{(\text{Mass of chloride} - \text{Mass of element})}$$

$$= \frac{\text{Mass of element} \times 35.5}{\text{Mass of chlorine}}$$

(iv) Metal to metal displacement method: A more active metal can displace less active metal from its salt's solution. For example, when zinc is added to copper sulphate, copper is precipitated. A known mass of active metal is added to the salt's solution of less active metal. The precipitated metal after drying is accurately weighed. The masses of the displacing metal and the displaced metal bear the same ratio as their equivalent masses. If E_1 and E_2 are the equivalent masses of two elements and m_1 and m_2 their respective masses, then,

$$\frac{m_1}{m_2} = \frac{E_1}{E_2}$$

Knowing the equivalent mass of one metal, the equivalent mass of the other metal can be calculated.

(v) Double decomposition method: This method is based on the following points:

(a) The mass of the compound reacted and the mass of product formed are in the ratio of their equivalent masses.

(b) The equivalent mass of the compound (electrovalent) is the sum of equivalent masses of its radicals.

(c) The equivalent mass of a radical is equal to the formula mass of the radical divided by its charge.

$$AB + CD \longrightarrow \underset{\text{ppt.}}{AD} + CB$$

$$\frac{\text{Mass of } AB}{\text{Mass of } AD} = \frac{\text{Equivalent mass of } AB}{\text{Equivalent mass of } AD}$$

$$= \frac{\text{Eq. mass of } A + \text{Eq. mass of } B}{\text{Eq. mass of } A + \text{Eq. mass of } D}$$

Knowing the equivalent masses of B and D, equivalent mass of A can be calculated.

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

7. An unknown element forms an oxide. What will be the equivalent mass of the element if the oxygen content is 20% by mass: [JEE (WB) 2008]
(a) 16 (b) 32 (c) 8 (d) 64
[Ans. (b)]
[Hint: Equivalent mass of element $= \frac{\text{Mass of element}}{\text{Mass of oxygen}} \times 8$
$= \frac{80}{20} \times 8 = 32$]

8. A metal M of equivalent mass E forms an oxide of molecular formula M_xO_y. The atomic mass of the metal is given by the correct equation: [PMT (Kerala) 2008]
(a) $2E(y/x)$ (b) xyE
(c) E/y (d) y/E
(e) $\frac{E}{2} \times \frac{x}{y}$
[Ans. (a)]
[Hint: Let atomic mass of metal M is 'a'.
Mass of metal $= a \times x$
Mass of oxygen $= 16 \times y$
Equivalent mass of element $= \frac{\text{Mass of element}}{\text{Mass of oxygen}} \times 8$
$E = \frac{ax}{16y} \times 8$
$a = 2E\left(\frac{y}{x}\right)$]

9. The percentage of an element M is 53 in its oxide of molecular formula M_2O_3. Its atomic mass is about: [PET (Kerala) 2008]
(a) 45 (b) 9 (c) 18 (d) 38
(e) 27
[Ans. (e)]
[Hint: Equivalent mass of element $= \frac{\text{Mass of element}}{\text{Mass of oxygen}} \times 8$
$= \frac{53}{47} \times 8 \simeq 9$
Atomic mass = Equivalent mass × Valency
$= 9 \times 3 = 27$ amu.]

10. The equivalent weight of a metal is double than that of oxygen. How many times is the weight of its oxide greater than the weight of metal?
(a) 4 (b) 2 (c) 3 (d) 1.5
[Ans. (d)]
[Hint: Equivalent mass of metal $= 16 = \frac{x}{n}$
Where x = atomic mass of metal
n = valency of metal
Molecular formula of metal oxide $= M_2O_n$
$\frac{\text{Mass of metal oxide}}{\text{Mass of metal}} = \frac{2(16n) + 16(n)}{2(16n)} = 1.5$]

1.15 METHODS FOR THE DETERMINATION OF ATOMIC MASS

(i) Dulong and Petit's Law: According to this law, the product of atomic mass and specific heat of a solid element is approximately equal to 6.4. The product of atomic mass and specific heat is called atomic heat. Thus,

$$\text{Atomic mass} \times \text{Specific heat} = 6.4$$

or

$$\text{Atomic mass (approximate)} = \frac{6.4}{\text{Specific heat}}$$

In above formula, the specific heat must be in cal/g unit.

The equivalent mass of the element is determined experimentally and the valency, which is always a whole number, can be obtained by dividing approximate atomic mass with the equivalent mass and changing the value so obtained to the nearest whole number. In this way, exact atomic mass can be determined by multiplying equivalent mass with valency.

Example 26. *A chloride of an element contains 49.5% chlorine. The specific heat of the element is 0.056. Calculate the equivalent mass, valency and atomic mass of the element.*

Solution: Mass of chlorine in the metal chloride = 49.5%

$$\text{Mass of metal} = (100 - 49.5) = 50.5$$

$$\text{Equivalent mass of the metal} = \frac{\text{Mass of metal}}{\text{Mass of chlorine}} \times 35.5$$

$$= \frac{50.5}{49.5} \times 35.5 = 36.21$$

According to Dulong and Petit's law,

$$\text{Approximate atomic mass of the metal} = \frac{6.4}{\text{Specific heat}}$$

$$= \frac{6.4}{0.056} = 114.3$$

$$\text{Valency} = \frac{\text{Approximate atomic mass}}{\text{Equivalent mass}} = \frac{114.3}{36.21} = 3.1 \approx 3$$

Hence, exact atomic mass $= 36.21 \times 3 = 108.63$

Example 27. *On dissolving 2.0 g of metal in sulphuric acid, 4.51 g of the metal sulphate was formed. The specific heat of the metal is 0.057 cal g^{-1}. What is the valency of the metal and exact atomic mass?*

Solution: Equivalent mass of SO_4^{2-} radical

$$= \frac{\text{Ionic mass}}{\text{Valency}} = \frac{96}{2} = 48$$

Mass of metal sulphate $= 4.51\,g$

Mass of metal $= 2.0\,g$

Mass of sulphate radical $= (4.51 - 2.0) = 2.51\,g$

2.51 g of sulphate combine with 2.0 g of metal.

So, 48 g of sulphate will combine with

$$= \frac{2}{2.51} \times 48 = 38.24 \text{ g metal}$$

Equivalent mass of metal = 38.24

According to Dulong and Petit's law,

$$\text{Approximate atomic mass} = \frac{6.4}{\text{Specific heat}} = \frac{6.4}{0.057} = 112.5$$

$$\text{Valency} = \frac{\text{Approximate atomic mass}}{\text{Equivalent mass}}$$

$$= \frac{112.5}{38.24} = 2.9 \approx 3$$

Exact atomic mass $= 38.24 \times 3 = 114.72$

(ii) Cannizzaro's method: Atomic mass of an element may be defined as the smallest mass of the element present in the molecular mass of any one of its compounds. For this purpose, the following steps are followed:

(a) Molecular masses of a number of compounds in which the element is present are determined.

(b) Each compound is analysed. Mass of the element is determined in the molecular mass of each compound.

(c) The lowest mass of the element is taken its atomic mass.

The following table shows the application of this method:

Compound	Vapour density (V.D.)	Molecular mass = 2 V.D.	% of carbon by mass in compound	Mass of carbon in one molecular mass of the compound
Methane	8	16	75.0	$\frac{75.0 \times 16}{100} = 12$ g
Ethane	15	30	80.0	$\frac{80.0 \times 30}{100} = 24$ g
Carbon monoxide	14	28	42.9	$\frac{42.9 \times 28}{100} = 12$ g
Carbon dioxide	22	44	27.3	$\frac{27.3 \times 44}{100} = 12$ g
Propane	22	44	81.8	$\frac{81.8 \times 44}{100} = 36$ g

Least mass of carbon is 12 g.

Thus, the atomic mass of carbon is 12.

(iii) The law of isomorphism: Isomorphous substances form crystals which have same shape and size and can grow in the saturated solution of each other. They have a property of forming mixed crystals. Isomorphous substances have same composition, *i.e.*, they have same number of atoms arranged similarly.

Examples of isomorphous compounds are:

(a) K_2SO_4 and K_2CrO_4 (potassium sulphate and potassium chromate)

(b) $ZnSO_4 \cdot 7H_2O$ and $FeSO_4 \cdot 7H_2O$ (zinc sulphate and ferrous sulphate)

(c) $KClO_4$ and $KMnO_4$ (potassium perchlorate and potassium permanganate)

(d) $K_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$ and $K_2SO_4 \cdot Cr_2(SO_4)_3 \cdot 24H_2O$ (potash alum and chrome alum).

The following conclusions have been deduced from the phenomenon of isomorphism:

(i) Masses of two elements that combine with same mass of other elements in their respective compounds are in the ratio of their atomic masses.

$$\frac{\text{Mass of one element } (A) \text{ that combines with a certain mass of other elements}}{\text{Mass of other element } (B) \text{ that combines with the same mass of other elements}} = \frac{\text{Atomic mass of } A}{\text{Atomic mass of } B}$$

(ii) The valencies of the elements forming isomorphous compounds are the same.

Example 28. *Potassium chromate is isomorphous to potassium sulphate* (K_2SO_4) *and is found to contain* 26.78% *chromium. Calculate the atomic mass of chromium* ($K = 39.10$).

Solution: Since, the formula of potassium sulphate is K_2SO_4, so the formula of potassium chromate should be K_2CrO_4 as it is isomorphous to K_2SO_4.

If the atomic mass of chromium is A, then

formula mass of potassium chromate should be

$$= 2 \times 39.1 + A + 64 = (142.2 + A)$$

$$\%\text{ of chromium} = \frac{A}{(142.2 + A)} \times 100$$

So, $$\frac{100A}{(142.2 + A)} = 26.78$$

$$100A = 26.78\,(142.2 + A)$$

or $$A = \frac{26.78 \times 142.2}{73.22} = 52.00$$

(iv) Atomic mass from vapour density of a chloride: The following steps are involved in this method:

(a) Vapour density of the chloride of the element is determined.

(b) Equivalent mass of the element is determined.

Let the valency of the element be x. The formula of its chloride will be MCl_x.

$$\text{Molecular mass} = \text{Atomic mass of } M + 35.5x$$

$$= A + 35.5x$$

$$\text{Atomic mass} = \text{Equivalent mass} \times \text{Valency}$$

$$A = E \times x$$

$$\text{Molecular mass} = E \times x + 35.5x$$

$$2\,\text{V.D.} = x(E + 35.5)$$

$$x = \frac{2\,\text{V.D.}}{E + 35.5}$$

Knowing the value of valency, the atomic mass can be determined.

Example 29. *One gram of a chloride was found to contain* 0.835 *g of chlorine. Its vapour density is* 85. *Calculate its molecular formula.*

Solution: Mass of metal chloride = 1g

$$\text{Mass of chlorine} = 0.835\,\text{g}$$

$$\text{Mass of metal} = (1 - 0.835) = 0.165\ \text{g}$$

$$\text{Equivalent mass of metal} = \frac{0.165 \times 35.5}{0.835} = 7.01$$

$$\text{Valency of the metal} = \frac{2\,\text{V.D.}}{E + 35.5} = \frac{2 \times 85}{7.01 + 35.5} = 4$$

$$\text{Formula of the chloride} = MCl_4$$

Example 30. *The oxide of an element contains* 32.33 *per cent of the element and the vapour density of its chloride is* 79. *Calculate the atomic mass of the element.*

Solution: Mass of the element = 32.33 parts

Mass of oxygen = (100 − 32.33) = 67.67 parts

$$\text{Equivalent mass of the element} = \frac{32.33}{67.67} \times 8 = 3.82$$

$$\text{Valency of the element} = \frac{2\,\text{V.D.}}{E + 35.5} = \frac{2 \times 79}{3.82 + 35.5} = 4$$

Hence, the atomic mass of the element = 3.82 × 4

= 15.28

1.16 TYPES OF FORMULAE

As already stated in section 1.10, a formula is a group of symbols of the elements which represents one molecule of the substance. Formula represents chemical composition of the substance. There are three kinds of formulae in the case of compounds.

(i) Empirical formula: It represents the simplest relative whole number ratio of atoms of each element present in the molecule of the substance. For example, CH is the empirical formula of benzene in which ratio of the atoms of carbon and hydrogen is 1 : 1. It also indicates that the ratio of carbon and hydrogen is 12 : 1 by mass.

(ii) Molecular formula: Molecular formula of a compound is one which expresses as the actual number of atoms of each element present in one molecule. C_6H_6 is the molecular formula of benzene indicating that six carbon atoms and six hydrogen atoms are present in a molecule of benzene. Thus,

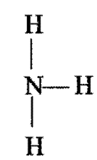
$$\text{Molecular formula} = n \times \text{Empirical formula}$$

where, $n = \dfrac{\text{Molecular formula mass}}{\text{Empirical formula mass}}$

Molecular formula gives the following informations:

(i) Various elements present in the molecule.
(ii) Number of atoms of various elements in the molecule.
(iii) Mass ratio of the elements present in the molecule. The mass ratio of carbon and oxygen in CO_2 molecule is 12 : 32 or 3 : 8.
(iv) Molecular mass of the substance.
(v) The number written before the formula indicates the number of molecules, *e.g.*, $2CO_2$ means 2 molecules of carbon dioxide.

(iii) Structural formula: It represents the way in which atoms of various elements present in the molecule are linked with one another. For example, ammonia is represented as:



The formula indicates that three hydrogen atoms are linked to one nitrogen atom by three single covalent bonds.

1.17 PERCENTAGE COMPOSITION OF A COMPOUND

Percentage composition of a compound is the relative mass of the each of the constituent element in 100 parts of it. It is readily calculated from the formula of the compound. Molecular mass of a compound is obtained from its formula by adding up the masses of all the atoms of the constituent elements present in the molecule.

Let the molecular mass of a compound be M and X be the mass of an element in the molecule.

$$\text{Percentage of element} = \frac{\text{Mass of element}}{M} \times 100 = \frac{X}{M} \times 100$$

Example 31. *Calculate the percentage composition of calcium nitrate.*

Solution: The formula of calcium nitrate is $Ca(NO_3)_2$.

Thus, the formula mass or molecular mass

= At. mass of Ca + 2 × At. mass of N + 6 × At. mass of oxygen

= 40 + 2 × 14 + 6 × 16

= 164

$$\%\text{ of Ca} = \frac{40}{164} \times 100 = 24$$

$$\%\text{ of N} = \frac{28}{164} \times 100 = 17$$

$$\%\text{ of O} = 100 - (24 + 17) = 59$$

Example 32. *Determine the percentage of water of crystallisation, iron, sulphur and oxygen in pure ferrous sulphate* ($FeSO_4 \cdot 7H_2O$).

Solution: The formula mass of ferrous sulphate

$$= \text{At. mass of Fe} + \text{At. mass of S} + 4 \times \text{At. mass of oxygen} + 7 \times \text{Mol. mass of } H_2O$$

$$= 56.0 + 32.0 + 4 \times 16.0 + 7 \times 18.0$$

$$= 278.0$$

So, $\%$ of water of crystallisation $= \dfrac{126}{278} \times 100 = 45.32$

$$\% \text{ of iron} = \frac{56}{278} \times 100 = 20.14$$

$$\% \text{ of sulphur} = \frac{32}{278} \times 100 = 11.51$$

$$\% \text{ of oxygen} = \frac{64}{278} \times 100 = 23.02$$

(Oxygen present in water molecules is not taken into account.)

Example 33. *It is found that* 16.5 g *of metal combine with oxygen to form* 35.60 g *of metal oxide. Calculate the percentage of metal and oxygen in the compound.*

Solution:

$$\text{Mass of oxygen in oxide} = (35.60 - 16.50) = 19.10 \text{ g}$$

$$\% \text{ of metal} = \frac{16.50}{35.60} \times 100 = 46.3$$

$$\% \text{ of oxygen} = \frac{19.10}{35.60} \times 100 = 53.7$$

Example 34. *Hydrogen and oxygen are combined in the ratio* 1 : 16 *by mass in hydrogen peroxide. Calculate the percentage of hydrogen and oxygen in hydrogen peroxide.*

Solution: 17 parts of hydrogen peroxide contain hydrogen = 1 part

100 parts of hydrogen peroxide contain hydrogen

$$= \frac{1}{17} \times 100 = 5.88$$

$$\% \text{ of oxygen} = (100 - 5.88) = 94.12$$

Example 35. *On analysis of an impure sample of sodium chloride, the percentage of chlorine was found to be* 45.5. *What is the percentage of pure sodium chloride in the given sample?*

Solution: The molecular mass of pure sodium chloride (NaCl)

$$= \text{At. mass of Na} + \text{At. mass of chlorine}$$

$$= (23 + 35.5) = 58.5$$

% of chlorine in pure NaCl

$$= \frac{35.5}{58.4} \times 100 = 60.6$$

Thus,

% of purity of NaCl in the sample

$$= \frac{45.5}{60.6} \times 100 = 75$$

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

11. A gas mixture contains 50% helium and 50% methane by volume. What is the percentage by mass of methane in the mixture? [CEE (Kerala) 2004]
(a) 19.97% (b) 20.05% (c) 50% (d) 75%
(e) 80.03%
[Ans. (e)]
[Hint: Molar and volume ratio will be same, *i.e.*, 1 : 1.
$\therefore$ Mass of 1 mole CH_4 and He will be 16 and 4 g respectively.

$$\text{Percentage by mass of } CH_4 = \frac{\text{Mass of } CH_4}{\text{Total mass}} \times 100$$

$$= \frac{16}{20} \times 100 \approx 80\%]$$

12. The atomic composition of the entire universe is approximately given in the table below:

Atom	% of total no. of atoms
H	93
He	7

Hydrogen atoms constitute what percentage of the universe by mass?
(a) 77% (b) 23% (c) 37% (d) 73%
[Ans. (a)]
[Hint: Mass of 93 'H' atoms = 93 amu
Mass of 7 'He' atoms = 28 amu

$$\% \text{ Hydrogen by mass} = \frac{93}{(93 + 28)} \times 100 = 77\%]$$

13. Which pair of species has same percentage of carbon?
(a) CH_3COOH and $C_6H_{12}O_6$
(b) CH_3COOH and C_2H_5OH
(c) $HCOOCH_3$ and $C_{12}H_{22}O_{11}$
(d) $C_6H_{12}O_6$ and $C_{12}H_{22}O_{11}$
[Ans. (a)]
[Hint: Percentage of carbon in acetic acid $= \dfrac{24}{60} \times 100 = 40\%$

$$\text{Percentage of carbon in } C_6H_{12}O_6 = \frac{72}{180} \times 100 = 40\%]$$

14. Which of the following alkanes has 75% of carbon?
(a) C_2H_6 (b) CH_4 (c) C_3H_8 (d) C_4H_{10}
[Ans. (b)]
[Hint: Percentage of carbon in methane $= \dfrac{12}{16} \times 100 = 75\%$]

15. Which of the following two oxides of nitrogen have 30.5% nitrogen?
(a) NO (b) NO_2 (c) N_2O_4 (d) N_2O_5
[Ans. (b) and (c)]
[Hint: Percentage of nitrogen in $NO_2 = \dfrac{14}{46} \times 100 = 30.5\%$

$$\text{Percentage of nitrogen in } N_2O_4 = \frac{28}{92} \times 100 = 30.5\%]$$

1.18 DETERMINATION OF EMPIRICAL AND MOLECULAR FORMULAE

The following steps are followed to determine the empirical formula of the compound :

(i) The percentage composition of the compound is determined by quantitative analysis.

(ii) The percentage of each element is divided by its atomic mass. It gives atomic ratio of the elements present in the compound.

(iii) The atomic ratio of each element is the divided by the minimum value of atomic ratio as to get the simplest ratio of the atoms of elements present in the compound.

(iv) If the simplest ratio is fractional, then values of simplest ratio of each element is multiplied by a smallest integer to get a simplest whole number for each of the element.

(v) To get the empirical formula, symbols of various elements present are written side by side with their respective whole number ratio as a subscript to the lower right hand corner of the symbol.

The molecular formula of a substance may be determined from the empirical formula if the molecular mass of the substance is known. The molecular formula is always a simple multiple of empirical formula and the value of simple multiple is obtained by dividing molecular mass with empirical formula mass.

Example 36. *Calculate the empirical formula for a compound that contains* 26.6% *potassium*, 35.4% *chromium and* 38.1% *oxygen.*

[*Given* $K = 39.1; Cr = 52; O = 16$]

Solution:

Element	Percentage	Atomic mass	Relative number of atoms	Simplest ratio	Simplest whole number ratio
Potassium	26.6	39.1	$\frac{26.6}{39.1} = 0.68$	$\frac{0.68}{0.68} = 1$	$1 \times 2 = 2$
Chromium	35.4	52.0	$\frac{35.4}{52} = 0.68$	$\frac{0.68}{0.68} = 1$	$1 \times 2 = 2$
Oxygen	38.1	16.0	$\frac{38.1}{16} = 2.38$	$\frac{2.38}{0.68} = 3.5$	$3.5 \times 2 = 7$

Therefore, empirical formula is $K_2Cr_2O_7$.

Example 37. *A compound contains* 34.8% *oxygen*, 52.2% *carbon and* 13.0% *hydrogen. What is the empirical formula mass of the compound?*

Solution:

Element	Percentage	Atomic mass	Relative number of atoms	Simplest ratio
Oxygen	34.8	16	$\frac{34.8}{16} = 2.175$	$\frac{2.175}{2.175} = 1$
Carbon	52.2	12	$\frac{52.2}{12} = 4.35$	$\frac{4.35}{2.175} = 2$
Hydrogen	13.0	1	$\frac{13.0}{1} = 13.0$	$\frac{13.0}{2.175} = 6$

The empirical formula is C_2H_6O.

Empirical formula mass $= (2 \times 12) + (6 \times 1) + 16 = 46$

Example 38. *A compound of carbon, hydrogen and nitrogen contains these elements in the ratio* 9 : 1 : 3.5. *Calculate the empirical formula. If its molecular mass is* 108, *what is the molecular formula?*

Solution:

Element	Element ratio	Atomic mass	Relative number of atoms	Simplest ratio
Carbon	9	12	$\frac{9}{12} = 0.75$	$\frac{0.75}{0.25} = 3$
Hydrogen	1	1	$\frac{1}{1} = 1$	$\frac{1}{0.25} = 4$
Nitrogen	3.5	14	$\frac{3.5}{14} = 0.25$	$\frac{0.25}{0.25} = 1$

The empirical formula $= C_3H_4N$

Empirical formula mass $= (3 \times 12) + (4 \times 1) + 14 = 54$

$$n = \frac{\text{Mol. mass}}{\text{Emp. mass}} = \frac{108}{54} = 2$$

Thus, molecular formula of the compound

$$= 2 \times \text{Empirical formula}$$
$$= 2 \times C_3H_4N = C_6H_8N_2$$

Example 39. *A carbon compound containing only carbon and oxygen has an approximate molecular mass of* 290. *On analysis, it is found to contain* 50% *by mass of each element. What is the molecular formula of the compound?*

Solution:

Element percentage	Atomic mass	Relative number of atoms	Simplest ratio	Simplest whole number ratio
Carbon 50.0	12	4.166	$\frac{4.166}{3.125} = 1.33$	4
Oxygen 50.0	16	3.125	$\frac{3.125}{3.125} = 1$	3

The empirical formula $= C_4O_3$

Empirical formula mass $= (4 \times 12) + (3 \times 16) = 96$

Molecular mass $= 290$

$$n = \frac{\text{Mol. mass}}{\text{Emp. mass}} = \frac{290}{96} = 3 \text{ approximately}$$

Molecular formula $= n \times$ Empirical formula

$= 3 \times C_4O_3 = C_{12}O_9$

Example 40. *A compound on analysis, was found to have the following composition: (i) Sodium* $= 14.31\%$, *(ii) Sulphur* $= 9.97\%$, *(iii) Oxygen* $= 69.50\%$, *(iv) Hydrogen* $= 6.22\%$. *Calculate the molecular formula of the compound assuming that whole of hydrogen in the compound is present as water of crystallisation. Molecular mass of the compound is* 322.

Solution:

Element	Percentage	Atomic mass	Relative number of atoms	Simplest ratio
Sodium	14.31	23	0.622	$\frac{0.622}{0.311}=2$
Sulphur	9.97	32	0.311	$\frac{0.311}{0.311}=1$
Hydrogen	6.22	1	6.22	$\frac{6.22}{0.311}=20$
Oxygen	69.50	16	4.34	$\frac{4.34}{0.311}=14$

The empirical formula $= Na_2SH_{20}O_{14}$

Empirical formula mass $= (2\times 23)+32+(20\times 1)+(14\times 16)$
$= 322$

Molecular mass $= 322$

Molecular formula $= Na_2SH_{20}O_{14}$

Whole of the hydrogen is present in the form of water. Thus, 10 water molecules are present in the molecule.

So, molecular formula $= Na_2SO_4\cdot 10H_2O$

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

16. An organic compound contains 49.30% carbon, 6.84% hydrogen and its vapour density is 73. Molecular formula of the compound is: **[CET (Kerala) 2004]**

(a) $C_3H_8O_2$ (b) $C_3H_{10}O_2$ (c) C_6H_9O (d) $C_4H_{10}O_2$
(e) $C_6H_{10}O_4$

[Ans. (e)]

[Hint: Molecular mass $= 2\times 73 = 146$

$$C=\frac{\%}{100}\times\frac{\text{Molecular mass}}{\text{Atomic mass}}=\frac{49.30}{100}\times\frac{146}{12}=6$$

$$H=\frac{\%}{100}\times\frac{\text{Molecular mass}}{\text{Atomic mass}}=\frac{6.84}{100}\times\frac{146}{1}=10$$

$$O=\frac{\%}{100}\times\frac{\text{Molecular mass}}{\text{Atomic mass}}=\frac{43.86}{100}\times\frac{146}{16}=4$$

Molecular formula $= C_6H_{10}O_4$

Molecular mass$= 12\times 6+10\times 1+16\times 4 = 146$

'Or'

Element	Percentage	Atomic mass	Relative number of atoms	Simplest ratio
Carbon	49.30	12	4.10	$1.5\times 2=3$
Hydrogen	6.84	1	6.84	$2.5\times 2=5$
Oxygen	43.86	16	2.74	$1\times 2=2$

The empirical formula $= C_3H_5O_2$

$$n=\frac{2\times 73}{73}=2$$

Molecular formula $= 2\times C_3H_5O_2 = C_6H_{10}O_4$]

17. A compound has an empirical formula C_2H_4O. An independent analysis gave a value of 132.16 for its molecular mass. What is the correct molecular formula? **[CET (Kerala) 2004]**

(a) $C_4H_4O_5$ (b) $C_{10}H_{12}$
(c) C_7O_3 (d) $C_6H_{12}O_3$
(e) $C_4H_8O_5$

[Ans. (d)]

[Hint: Molecular formula $= (C_2H_4O)_n$

$$n=\frac{\text{Molecular mass}}{\text{Empirical formula mass}}=\frac{132.16}{44}=3$$

Molecular formula $= (C_2H_4O)_3 = C_6H_{12}O_3$]

18. An organic compound containing C and H has 92.30% carbon. Its empirical formula is:

(a) CH (b) CH_3
(c) CH_2 (d) CH_4

[Ans. (a)]

[Hint: Percentage of carbon $=\frac{12}{13}\times 100 = 92.30\%$

'Or'

Element	Percentage	Atomic mass	Relative number of atoms	Simplest ratio
Carbon	92.30	12	7.69	1
Hydrogen	7.70	1	7.70	1

Empirical formula = CH]

19. Two oxides of a metal contain 50% and 40% of metal M respectively. If the formula of first oxide is MO, the formula of 2nd oxide will be:

(a) MO_2 (b) M_2O_3
(c) M_2O (d) M_2O_5

[Ans. (b)]

[Hint:

Compound 1		Compound 2	
M	O	M	O
50%	50%	40%	60%
50 g	50 g	40 g	60 g
1 g	$\frac{50}{50}=1$ g	1 g	$\frac{60}{40}=1.5$ g
2 g	2 g	2 g	3 g

Formula: MO M_2O_3]

20. Two elements X and Y have atomic mass 75 and 16 respectively. They combine to give a compound having 75.8% X. The formula of the compound is:

(a) XY (b) X_2Y
(c) X_2Y_2 (d) X_2Y_3

[Ans. (d)]

[Hint: Molecular mass of $X_2Y_3 = 2\times 75+3\times 16 = 198$

$$\text{Percentage of } X = \frac{150}{198} \times 100 = 75.80\%$$

'Or'

Element	Percentage	Atomic mass	Relative number of atoms	Simplest ratio
X	75.80	75	1.01	$1 \times 2 = 2$
Y	24.20	16	1.51	$1.5 \times 2 = 3$

Formula = X_2Y_3]

21. The crystalline salt $Na_2SO_4 \cdot xH_2O$ on heating loses 55.9% of its mass. The formula of crystalline salt is:
(a) $Na_2SO_4 \cdot 5H_2O$ (b) $Na_2SO_4 \cdot 7H_2O$
(c) $Na_2SO_4 \cdot 2H_2O$ (d) $Na_2SO_4 \cdot 10H_2O$
(e) $Na_2SO_4 \cdot 6H_2O$ **[PMT (Kerala) 2007]**
[Ans. (d)]
[Hint: Molecular mass of $Na_2SO_4 \cdot 10H_2O$
$= 46 + 96 + 180 = 322$ amu

$$\% \text{ by mass of } H_2O = \frac{180}{322} \times 100 = 55.9\%]$$

1.19 CHEMICAL EQUATION

A chemical equation is a symbolic representation of a chemical change.

The substances, in which the chemical change is brought, are called reactants and the substances which come into existence as the result of chemical change are called products. The relationship between reactants and products is represented in the form of a chemical equation. The symbols or formulae of the reactants are written on left hand side of equality (=) or $\rightarrow$ sign and the symbols or formulae of products on right hand side. The symbols or formulae on both the sides are added by + sign. Such an equation is known as skeleton equation. The equation becomes balanced when total number of atoms of various elements are made equal on both the sides. Gases are always written in molecular form.

$$KClO_3 \longrightarrow KCl + O_2$$

This is the skeleton equation as it only represents reactant and products involved in the chemical change but the following equation is a balanced equation as the number of atoms of various elements is equal on both sides.

$$\underset{\text{Reactant}}{2KClO_3} = \underbrace{2KCl + 3O_2}_{\text{Products}}$$

The following notations are also used in chemical equations as to provide more information about chemical change:

(i) Upper arrow ($\uparrow$) is written immediately after the gaseous product.
(ii) Lower arrow ($\downarrow$) is written immediately after the insoluble substance (solid) which deposits from a solution.
(iii) Symbols, (s) for solid, (l) for liquid and (g) for gas are also written to represent the physical state of the reactants and products.
(iv) Symbol ($aq.$) is written for substances dissolved in water.
(v) Symbol (Δ) is written over an arrow or over an equality sign to represent heating.

Information Obtained from Chemical Equation

A balanced chemical equation provides the following informations:

(i) What are the reactants and products involved in the chemical change?
(ii) The relative number of molecules of reactants and products.
(iii) The relative number by parts of mass of reactants and products.
(iv) Relative volumes of gaseous reactants and products.

For example, consider the following reaction:

$$CH_4(g) + 2O_2(g) = CO_2(g) + 2H_2O(g)$$

This equation tells us that methane and oxygen are reactants and carbon dioxide and water are products. One molecule of methane reacts with two molecules of oxygen to produce one molecule of CO_2 and two molecules of water or one mole of methane reacts with two moles of oxygen to produce one mole of carbon dioxide and two moles of water or 16 g of methane reacts with 64 g of oxygen to produce 44 g of CO_2 and 36 g of water. This equation also tells that 1 vol. of methane reacts with 2 vol. of oxygen to produce 1 vol. of CO_2 and 2 vol. of steam under similar conditions of temperature and pressure.

Limitations of Chemical Equation

A chemical equation fails to provide the following informations:

(i) Actual concentration of the reactants taken and the actual concentration of the products obtained.
(ii) Time taken for the completion of the chemical change.
(iii) Conditions applied for bringing the chemical change.
(iv) Whether the reaction is reversible or irreversible.

The following efforts have been made to make the chemical equations more informative by introducing:

(i) Experimental conditions: If a particular chemical change occurs under certain temperature and pressure conditions, these are mentioned above and below the ($\rightarrow$) or (=) sign.

$$N_2 + 3H_2 \xrightarrow[450°C]{200\text{ atm}} 2NH_3$$

If the reaction occurs in presence of a catalyst, it is written above the ($\rightarrow$) or (=) sign.

$$2SO_2 + O_2 \xrightarrow{Pt} 2SO_3$$

(ii) Heat evolved or absorbed: Heat evolved or absorbed in a chemical change can be represented by adding or subtracting the amount of heat on right hand side.

$$N_2 + O_2 \longrightarrow 2NO - 43.2\text{ kcals}$$
$$C + O_2 \longrightarrow CO_2 + 94.3\text{ kcals}$$

(iii) Reversible or irreversible nature: Reversible reactions are shown by changing the sign of equality (=) or arrow ($\rightarrow$) with sign of double arrow ($\rightleftharpoons$).

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Types of Chemical Equations

Chemical equations are of two types:

(i) Molecular equations
(ii) Ionic equations.

Molecular equations are those in which reactants and products are represented in the form of molecules.

$$BaCl_2 + Na_2SO_4 = BaSO_4 \downarrow + 2NaCl$$
$$2NaOH + H_2SO_4 = Na_2SO_4 \downarrow + 2H_2O$$

Ionic equations are those in which reactants and products are written in ionic form. The molecular equation

$$BaCl_2 + Na_2SO_4 = BaSO_4 \downarrow + 2NaCl$$

can be written in ionic form as:

$$Ba^{2+} + 2Cl^- + 2Na^+ + SO_4^{2-} = BaSO_4 \downarrow + 2Na^+ + 2Cl^-$$
$$Ba^{2+} + SO_4^{2-} = BaSO_4 \downarrow$$

Note: Calculations based on chemical equations have been dealt in the chapter 'Stoichiometry' in 'Inorganic Chemistry'.

1.20 MEASUREMENT IN CHEMISTRY: FUNDAMENTAL AND DERIVED UNITS

Chemistry is an experimental science. An experiment always involves observation of a phenomenon under certain set of conditions. The quantitative scientific observation generally requires the measurement of one or more physical quantities such as mass, length, density, volume, pressure, temperature, etc.

A physical quantity is expressed in terms of a number and a unit. Without mentioning the unit, the number has no meaning. For example, the distance between two points is "four" has no meaning unless a specific unit (inch, centimetre, metre, etc.,) is associated with the number. The units of physical quantities depend on three basic units, *i.e.*, units of mass, length and time. Since, these are independent units and cannot be derived from any other units, they are called **fundamental units.** It was soon realised that the three fundamental units cannot describe all the physical quantities such as temperature, intensity of luminosity, electric current and the amount of the substance. Thus, seven units of measurement, namely mass, length, time, temperature, electric current, luminous intensity and amount of substance are taken as **basic units.** All other units can be derived from them and are, therefore, called **derived units.** The units of area, volume, force, work, density, velocity, energy, etc., are all derived units.

SI Units of Measurement

Various systems of units were in use prior to 1960. The common ones are the following:

(i) The English or FPS system: The system uses the **foot,** the **pound** and the **second** for length, mass and time measurements respectively. It is not used now-a-days.

(ii) MKS system: Here M stands for **metre** (a unit of length), K for **kilogram** (a unit of mass) and S for **second** (a unit of time). This is a decimal system.

(iii) CGS system: Here the unit of length is **centimetre,** the unit of mass is **gram** and the unit of time is **second.** It is also a decimal system.

MKS system often known as metric system was very popular throughout the world, but the drawback with this system was that a number of different metric units for the same quantity were used in different parts of the world. In 1964, the National Bureau of Standards adopted a slightly modified version of the metric system, which had been officially recommended in 1960 by an international body, **General Conference of Weights and Measures.** This revised set of units is known as the **International System of Units** (abbreviated SI). Now the SI units have been accepted by the scientists all over the world in all branches of science, engineering and technology.

The SI system have seven basic units. The various fundamental quantities that are expressed by these units along with their symbols are tabulated below:

Basic physical quantity	Unit	Symbol
Length	Metre	m
Mass	Kilogram	kg
Time	Second	s
Temperature	Kelvin	K
Electric current	Ampere	amp or A
Luminous intensity	Candela	cd
Amount of substance	Mole	mol

Sometimes, submultiples and multiples are used to reduce or enlarge the size of the different units. The names and symbols of sub-multiples and multiples are listed in the table given below.

The name for the base unit for mass, the kilogram, already contains a prefix. The names of other units of mass are obtained by substituting other prefixes for prefix kilo. The names of no other base units contain prefixes.

The use of SI system is slowly growing, however, older systems are still in use. Furthermore, the existence of older units in scientific literature demands that one must be familiar with both old and new systems.

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

Greek Alphabets

Alpha	A	α	Nu	N	ν
Beta	B	β	Xi	Ξ	ξ
Gamma	Γ	γ	Omicron	O	o
Delta	Δ	δ	Pi	Π	π
Epsilon	E	ε	Rho	P	ρ
Zeta	Z	ζ	Sigma	Σ	σ
Eta	H	η	Tau	τ	τ

Theta	Θ	θ	Upsilon	Υ	υ
Iota	Ι	ι	Phi	Φ	ϕ
Kappa	Κ	κ	Chi	Χ	χ
Lambda	Λ	λ	Psi	Ψ	ψ
Mu	Μ	μ	Omega	Ω	ω

Numerical Prefix

Prefix	Value	Prefix	Value
Hemi	(1/2)	Deca	10
Mono	1	Undeca	11
Sesqui	$1\frac{1}{2}$	Dodeca	12
Di or Bi	2	Trideca	13
Tri	3	Tetradeca	14
Tetra	4	Pentadeca	15
Penta	5	Hexadeca	16
Hexa	6	Heptadeca	17
Hepta	7	Octadeca	18
Octa	8	Nonadeca	19
Nona	9	Eicosa	20

SI Units for Some Common Derived Quantities

(a) $\text{Area} = \text{length} \times \text{breadth}$

$= \text{m} \times \text{m} = \text{m}^2$ [square metre]

(b) $\text{Volume} = \text{length} \times \text{breadth} \times \text{height}$

$= \text{m} \times \text{m} \times \text{m} = \text{m}^3$ [cubic metre]

(c) $$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{\text{kg}}{\text{m}^3} = \text{kg m}^{-3}$$

(d) $$\text{Speed} = \frac{\text{distance covered}}{\text{time}} = \frac{\text{metre}}{\text{time}} = \text{m s}^{-1}$$

(e) $$\text{Acceleration} = \frac{\text{change in velocity}}{\text{time taken}} = \frac{\text{m s}^{-1}}{\text{s}} = \text{m s}^{-2}$$

(f) $\text{Force} = \text{mass} \times \text{acceleration}$

$= \text{kg} \times \text{m s}^{-2}$

$= \text{kg m s}^{-2}$ (Newton, abbreviated as N)

(g) $\text{Pressure} = \text{force per unit area}$

$$= \frac{\text{kg m s}^{-2}}{\text{m}^2} = \text{kg m}^{-1}\,\text{s}^{-2} \text{ or Nm}^{-2}$$

(Pascal—Pa)

(h) $\text{Energy} = \text{force} \times \text{distance travelled}$

$= \text{kg m s}^{-2} \times \text{m}$

$= \text{kg m}^2\,\text{s}^{-2}$ (joule—J)

Some Old Units Still in Use

The use of some of the old units is still permitted. The 'litre', for example, which is defined as 1 cubic decimetre is used frequently by chemists. Certain other units which are not a part of SI units are still retained for a limited period of time. The term atmosphere (atm), the unit of pressure, falls into this category. Few of the old units along with conversion factors are given below:

Length: The interatomic distances are reported in units of angstrom (Å), nanometre (nm) or picometre (pm).

$$1\,\text{Å} = 10^{-8}\text{ cm} = 10^{-10}\text{ m}$$

$$1\text{ nm} = 10^{-7}\text{ cm} = 10^{-9}\text{ m} = 10\,\text{Å}$$

$$1\text{ pm} = 10^{-10}\text{ cm} = 10^{-12}\text{ m} = 10^{-2}\,\text{Å}$$

$$1\text{ nm} = 10^{3}\text{ pm}$$

Mass: The basic unit of mass is generally taken as gram (g). The gram is 10^{-3} kg.

$$1\text{ kilogram (kg)} = 10^3\text{ g}$$

$$1\text{ milligram (mg)} = 10^{-3}\text{ g}$$

$$1\text{ microgram }(\mu\text{g}) = 10^{-6}\text{ g}$$

While dealing with atoms and molecules, the term atomic mass unit (amu) is used. One amu is taken exactly as $\frac{1}{12}$ of the mass of an atom of the carbon isotope, C^{12}.

$$1\text{ amu} = 1.6605 \times 10^{-24}\text{ g} = 1.6605 \times 10^{-27}\text{ kg}$$

Volume: The units of volume are reported as cubic centimetre (cm^3) and cubic decimetre (dm^3). Cubic decimetre is termed **litre** while cubic centimetre is termed **millilitre.**

$$1\text{ litre (lit or L)} = (10\text{ cm})^3 = 1000\text{ cm}^3 = 10^{-3}\text{ m}^3$$

$$1\text{ millilitre (mL)} = (1\text{ cm})^3 = 1\text{ cm}^3\text{ (cc)} = 10^{-6}\text{ m}^3$$

So, $\quad 1\text{ litre} = 1000\text{ mL}$

Temperature: The **celsius temperature scale** which is not a part of SI system, is employed in scientific studies. This scale is based on the assignment of 0°C to the normal freezing point of water and 100°C to the normal boiling point of water. The celsius scale was formerly called the **centigrade scale.**

The unit of temperature in SI system is **Kelvin.** A degree on the kelvin scale has the same magnitude as the degree on the celsius scale but zero on the kelvin scale is equal to –273.15°C. The temperature (0 K) is often referred to as **absolute zero.**

So, $\quad \text{K} = (°\text{C} + 273.15)$

or $\quad °\text{C} = (\text{K} - 273.15)$

There is another important temperature scale known as **fahrenheit scale.** In this scale, the normal freezing point of water is 32°F and normal boiling point is 212°F. Thus, 100°C equals 180°F. Both the scales are related by the following equations:

$$°\text{C} = \frac{5}{9} \times (°\text{F} - 32)$$ [since, 100 parts on celsius scale = 180 parts on fahrenheit scale]

$$°\text{F} = \frac{9}{5} \times (°\text{C}) + 32$$

Pressure: There are three non-SI units for pressure which are commonly used.

(a) Atmosphere (atm) is defined as the pressure exerted by a column of mercury of 760 mm or 76 cm height at 0°C.

(b) Torr is defined as the pressure exerted by a 1 mm column of mercury at 0°C.

(c) Millimetre of mercury (mm Hg).

These three units are related as:

$$1\,\text{atm} = 760\,\text{torr} = 760\,\text{mm Hg} = 76\,\text{cm Hg} = 1.013 \times 10^5\ \text{Pa}$$

Energy: Calorie has been used in the past as a unit of energy measurement. The calorie was defined as the amount of heat required to raise the temperature of one gram of water from 14.5°C to 15.5 °C. One calorie is defined as exactly equal to 4.184 joules.

$$1\,\text{cal} = 4.184\,\text{J} \quad \text{or} \quad 1\,\text{J} = 0.2390\,\text{cal}$$

$$1\,\text{kcal} = 1000\,\text{cal} = 4.184\,\text{kJ}$$

Conversion factors

$$1\,\text{angstrom}\,(\text{Å}) = 10^{-8}\ \text{cm} = 10^{-10}\ \text{m} = 10^{-1}\ \text{nm} = 10^2\ \text{pm}$$

$$1\,\text{inch} = 2.54\,\text{cm} \quad \text{or} \quad 1\,\text{cm} = 0.394\,\text{inch}$$

$$39.37\,\text{inch} = 1\,\text{metre} \qquad 1\,\text{km} = 0.621\,\text{mile}$$

$$1\,\text{kg} = 2.20\,\text{pounds (lb)} \qquad 1\,\text{g} = 0.0353\,\text{ounce (o)}$$

$$1\,\text{pound (lb)} = 453.6\,\text{g}$$

$$\begin{aligned} 1\,\text{atomic mass unit (amu)} &= 1.6605 \times 10^{-24}\,\text{g} \\ &= 1.6605 \times 10^{-27}\,\text{kg} \\ &= 1.492 \times 10^{-3}\ \text{erg} = 1.492 \times 10^{-10}\ \text{J} \\ &= 3.564 \times 10^{11}\ \text{cal} = 9.310 \times 10^8\ \text{eV} \\ &= 931.48\,\text{MeV} \end{aligned}$$

$$\begin{aligned} 1\,\text{atmosphere (atm)} &= 760\,\text{torr} = 760\,\text{mm Hg} = 76\,\text{cm Hg} \\ &= 1.01325 \times 10^5\ \text{Pa} \end{aligned}$$

$$\begin{aligned} 1\,\text{calorie (cal)} &= 4.1840 \times 10^7\ \text{erg} = 4.184\ \text{J} \\ &= 2.613 \times 10^{19}\ \text{eV} \end{aligned}$$

$$1\,\text{coulomb (coul)} = 2.9979 \times 10^9\ \text{esu}$$

$$1\,\text{curie (Ci)} = 3.7 \times 10^{10}\ \text{disintegrations sec}^{-1}$$

$$\begin{aligned} 1\,\text{electron volt (eV)} &= 1.6021 \times 10^{-12}\,\text{erg} = 1.6021 \times 10^{-19}\ \text{J} \\ &= 3.827 \times 10^{-20}\,\text{cal} \\ &= 23.06\,\text{kcal mol}^{-1} \end{aligned}$$

$$1\,\text{erg} = 10^{-7}\,\text{J} = 2.389 \times 10^{-8}\,\text{cal} = 6.242 \times 10^{11}\,\text{eV}$$

$$1\,\text{electrostatic unit (esu)} = 3.33564 \times 10^{-10}\ \text{coul}$$

$$1\,\text{faraday}\ (F) = 9.6487 \times 10^4\ \text{coul}$$

$$1\,\text{dyne (dyne)} = 10^{-5}\ \text{N}$$

$$1\,\text{joule} = 10^7\ \text{erg} = 0.2390\,\text{cal}$$

$$\begin{aligned} 1\,\text{litre} &= 1000\,\text{cc} = 1000\,\text{mL} = 1\,\text{dm}^3 \\ &= 10^{-3}\ \text{m}^3 \end{aligned}$$

Values of Some Useful Constants

Fundamental constant	Value in old units	Value in SI units
Avogadro's number (N)	6.023×10^{23} mol^{-1}	6.023×10^{23} mol^{-1}
Atomic mass unit (amu)	1.6605×10^{-24} g	1.6605×10^{-27} g
Bohr radius (a_0)	0.52918 Å = 0.52918 $\times 10^{-8}$ cm	5.2918×10^{-11} m
Boltzmann constant (k)	1.3807×10^{-16} erg deg^{-1}	1.3807×10^{-23} JK^{-1}
Charge on electron (e)	(–) 4.8029×10^{-10} esu	(–) 1.6021×10^{-19} coul
Charge to mass ratio e/m of electron	1.7588×10^8 coul g^{-1}	1.7588×10^{11} coul kg^{-1}
Electron rest mass (m_e)	9.1091×10^{-28} g	9.1091×10^{-31} kg
Gas constant (R)	0.0821 lit atm deg^{-1}mol^{-1} 8.314×10^7 erg deg^{-1}mol^{-1} $1.987 \simeq 2.0$ cal deg^{-1}mol^{-1}	8.314 J K^{-1} mol^{-1}
Molar volume at NTP (V_m)	22.4 L mol^{-1}	22.4×10^{-3} m^3 mol^{-1}
Planck's constant (h)	6.6252×10^{-27} erg sec	6.6252×10^{-34} J sec
Proton mass (m_p)	1.6726×10^{-24} g	1.6726×10^{-27} kg
Neutron mass (m_n)	1.67495×10^{-24} g	1.67495×10^{-27} kg
Rydberg constant (R_z)	109678 cm^{-1}	1.09678×10^7 m^{-1}
Velocity of light (c) in vacuum	2.9979×10^{10} cm sec^{-1} or 186281 miles sec^{-1}	2.9979×10^8 m sec^{-1}
Faraday (F)	9.6487×10^4 C/equiv. or 96500 C/equiv.	9.6487×10^4 C/equiv.
$\dfrac{1}{4\pi\varepsilon_0}$	1	0.8988×10^{10} N m^{2}C^{-2} or 9×10^9 N m^{2}C^{-2}

Derived SI Units

Quantity with Symbol	Unit (SI)	Symbol
Velocity (v)	metre per sec	m s^{-1}
Area (A)	square metre	m^2
Volume (V)	cubic metre	m^3
Density (ρ)	kilogram m^{-3}	kg m^{-3}
Acceleration (a)	metre per sec^2	m s^{-2}
Energy (E)	joule (J)	kg m^2 s^{-2}
Force (F)	newton (N)	kg m s^{-2}

Power (W)	watt (W)	$J\,s^{-1}$; $kg\,m^2\,s^{-3}$
Pressure (P)	pascal (Pa)	$N\,m^{-2}$
Resistance (R)	ohm (Ω)	$V\,A^{-1}$
Conduction (C)	ohm^{-1}, mho, siemens	$m^{-2}\,kg^{-1}\,s^3\,A^2$ or Ω^{-1}
Potential difference	volt (V)	$kg\,m^2\,s^{-3}\,A^{-1}$
Electrical charge	coulomb (C)	A-s (ampere-second)
Frequency (ν)	hertz (Hz)	cycle per sec
Magnetic flux × density	tesla (T)	$kg\,s^{-2}\,A^{-1} = N\,A^{-1}\,m^{-1}$

Popular Units and their SI Equivalents

Physical quantity	Unit with symbol	Equivalent in SI unit
Mass	1 amu	1 amu = 1.6605×10^{-27} kg
Energy	1 electron volt (eV)	1.602×10^{-19} joule
Length	1 Å	10^{-10} m (10^{-1} nm)
Volume	litre	$10^{-3}\,m^3 = dm^3$
Force	dyne	10^{-5} N
Pressure	1 atmosphere	760 torr (760 mm Hg)
		101325 pa or 10^5 pa
	1 bar	101325 pa or 10^5 pa
	1 torr	133.322 $N\,m^{-2}$
Dipole moment	debye, 10^{-18} esu-cm	3.324×10^{-30} cm
Magnetic flux density	gauss (G)	10^{-4} T
Area of nuclear cross section	1 barn	$10^{-28}\,m^2$
Nuclear Diameter	1 fermi (1 femto)	10^{-15} m

Significant Figures

There is always some degree of uncertainty in every scientific measurement except in counting. The uncertainty in measurement mainly depends upon two factors:

(i) Skill and accuracy of the observer,

(ii) Limitation of the measuring scale.

To indicate the precision of a measurement, scientists use the term **significant figures.** The significant figures in a number are all certain digits plus one doubtful digit. The number of significant figures gives the information that except the digit at extreme right, all other digits are precise or reproducible. For example, mass of an object is 11.24 g. This value indicates that actual mass of the object lies between 11.23 g and 11.25 g. Thus, one is sure of first three figures (1, 1 and 2) but the fourth figure is somewhat inexact. The total significant figures in this number are four.

The following rules are observed in counting the number of significant figures in a given measured quantity:

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

42.3 has three significant figures.

243.4 has four significant figures.

24.123 has five significant figures.

(ii) A zero becomes significant figure if it appears between two non-zero digits. For example,

5.03 has three significant figures.

5.604 has four significant figures.

4.004 has four significant figures.

(iii) Leading zeros or the zeros placed to the left of the number are never significant. For example,

0.543 has three significant figures.

0.045 has two significant figures.

0.006 has one significant figure.

(iv) Trailing zeros or the zeros placed to the right of the number are significant. For example,

433.0 has four significant figures.

433.00 has five significant figures.

343.000 has six significant figures.

(v) In exponential notation, the numerical portion gives the number of significant figures. For example,

1.32×10^{-2} has three significant figures.

1.32×10^{4} has three significant figures.

(vi) The non-significant figures in the measurements are rounded off.

(a) If the figure following the last number to be retained is less than 5, all the unwanted figures are discarded and the last number is left unchanged, *e.g.*,

5.6724 is 5.67 to three significant figures.

(b) If the figure following the last number to be retained is greater than 5, the last figure to be retained is increased by 1 unit and the unwanted figures are discarded, *e.g.*,

8.6526 is 8.653 to four significant figures.

(c) If the figure following the last number to be retained is 5, the last figure is increased by 1 only in case it happens to be odd. In case of even number the last figure remains unchanged.

2.3524 is 2.4 to two significant figures.

7.4511 is 7.4 to two significant figures.

Calculations Involving Significant Figures

In most of the experiments, the observations of various measurements are to be combined mathematically, *i.e.*, added, subtracted, multiplied or divided as to achieve the final result. Since, all the observations in measurements do not have the same precision, it is natural that the final result cannot be more precise than the least precise measurement. The following two rules should be followed to obtain the proper number of significant figures in any calculation.

Rule 1: The result of an addition or subtraction in the numbers having different precisions should be reported to the same number of decimal places as are present in the number having the least number of decimal places. The rule is illustrated by the following examples:

(a)

	33.3	← (has only one decimal place)
	3.11	
	0.313	
Sum	36.723	← (answer should be reported to one decimal place)

Correct answer = 36.7

(b)

	3.1421	
	0.241	
	0.09	← (has 2 decimal places)
Sum	3.4731	← (answer should be reported to 2 decimal places)

Correct answer = 3.47

(c)

	62.831	← (has 3 decimal places)
	− 24.5492	
Difference	38.2818	← (answer should be reported to 3 decimal places after rounding off)

Correct answer = 38.282

Rule 2: The answer to a multiplication or division is rounded off to the same number of significant figures as is possessed by the least precise term used in the calculation. Examples are:

(a)

142.06	
× 0.23	← (two significant figures)
32.6738	← (answer should have two significant figures)

Correct answer = 33

(b)

51.028	
× 1.31	← (three significant figures)
66.84668	

Correct answer = 66.8

(c) $\dfrac{0.90}{4.26} = 0.2112676$

Correct answer = 0.21

Note: (i) Same procedure is followed if an expression involves multiplication as well as division.

(ii) The presence of exact numbers in an expression does not affect the number of significant figures in the answer.

Examples are:

(a) $\dfrac{3.24 \times 0.0866}{5.046} = 0.055643$ (b) $\dfrac{4.28 \times 0.146 \times 3}{0.0418} = 44.84784$

Correct answer = 0.0556 Correct answer = 44.8

MISCELLANEOUS NUMERICAL EXAMPLES

Example 1. *0.44 g of a hydrocarbon on complete combustion with oxygen gave 1.8 g water and 0.88 g carbon dioxide. Show that these results are in accordance with the law of conservation of mass.*

Solution: A hydrocarbon is a compound which consists of carbon and hydrogen only. It undergoes combustion forming carbon dioxide and water as products.

Formula of carbon dioxide = CO_2;

Molecular mass = $12 + 32 = 44$ g

Formula of water = H_2O;

Molecular mass = $2 + 16 = 18$ g

$$\text{Mass of carbon in 0.88 g of } CO_2 = \frac{12}{44} \times 0.88 = 0.24 \text{ g}$$

$$\text{Mass of hydrogen in 1.8 g of } H_2O = \frac{2}{18} \times 1.8 = 0.20 \text{ g}$$

Total masses of carbon and hydrogen in the products

$= 0.24 + 0.20 = 0.44$ g

This is equal to the mass of hydrocarbon before combustion.

Thus, the results are in accordance with the law of conservation of mass.

Example 2. *Calcium carbonate decomposes completely, on heating, into lime (CaO) and carbon dioxide (CO_2). 1 kg of calcium carbonate is completely decomposed by heat, when 560 g of lime are obtained. How much quantity of carbon dioxide in grams, moles and litres at NTP is produced in the process?*

Solution: According to law of conservation of mass,

Mass of lime + Mass of carbon dioxide = Mass of calcium carbonate

$$560 \text{ g} + \text{Mass of } CO_2 = 1000 \text{ g}$$

$$\text{Mass of } CO_2 = 1000 - 560 = 440 \text{ g}$$

$$\text{Molecular mass of } CO_2 = 12 + 32 = 44 \text{ g (1 mole)}$$

$$\text{No. of moles in 440 g of } CO_2 = \frac{440}{44} = 10$$

1 mole of CO_2 occupies volume at NTP = 22.4 litre

10 moles of CO_2 will occupy volume at NTP

$= 22.4 \times 10 = 224$ litre

Example 3. *10 mL of hydrogen combine with 5 mL of oxygen to yield water. When 200 mL of hydrogen at NTP are passed over heated CuO, the CuO loses 0.144 g of its mass. Do these results correspond to the law of constant proportions?*

Solution: 1st Case:

$$\text{Mass of 10 mL hydrogen at NTP} = \frac{2}{22400} \times 10 = 0.00089 \text{ g}$$

$$\text{Mass of 5 mL of oxygen at NTP} = \frac{32}{22400} \times 5 = 0.00714 \text{ g}$$

Mass of 22400 mL of Hg vapour at NTP $= \frac{8.923}{1000} \times 22400$

$= 199.87$ g

Hence, molecular mass of Hg $= 199.87$ g

(c) Approximate atomic mass $= \frac{6.4}{\text{Sp. heat}} = \frac{6.4}{0.033} \doteq 193.93$ g

Valency of Hg $= \frac{193.93}{100} \simeq 2$ (nearest whole number)

So, accurate atomic mass = Eq. mass × Valency

$= 100 \times 2 = 200$ g

Atomicity $= \frac{\text{Mol. mass}}{\text{At. mass}} = \frac{199.88}{200} \simeq 1$

Hence, mercury molecules are monoatomic.

Example 23. *How many grams of CaO are required to neutralise 852 g of P_4O_{10}?* **(IIT 2005)**

Solution: The reaction will be:

$$6CaO + P_4O_{10} \longrightarrow 2Ca_3(PO_4)_2$$

$$852 \text{ g } P_4O_{10} \equiv 3 \text{ mol } P_4O_{10}$$

1 mole of P_4O_{10} neutralises 6 moles of CaO.

∴ 3 moles of P_4O_{10} will neutralise 18 moles of CaO.

∴ Mass of CaO $= 18 \times 56 = 1008$ g

Example 24. *If 1 grain is equal to 64.8 mg, how many moles of aspirin (mol. wt. = 169) are present in a 5 grain aspirin tablet?*

Solution: Mass of aspirin in the tablet $= 64.8 \times 5 = 324$ mg

$= 0.324$ g

Number of moles $= \frac{\text{Mass}}{\text{Molar mass}} = \frac{0.324}{169}$

$= 1.92 \times 10^{-3}$

Example 25. *If the volume occupied in a crystal by a molecule of NaCl is 47×10^{-24} mL, calculate the volume of the crystal weighing 1 g.*

Solution: Number of molecules of NaCl

$= \frac{\text{Mass}}{\text{Molar mass}} \times 6.023 \times 10^{23}$

$= \frac{1}{58.5} \times 6.023 \times 10^{23} = 1.03 \times 10^{22}$

Volume of crystal $= 1.03 \times 10^{22} \times 47 \times 10^{-24} = 0.484$ mL

Example 26. *A plant virus is found to consist of uniform cylindrical particles of 150 Å in diameter and 5000 Å long. The specific volume of the virus is 0.75 cm^3/g. If the virus is considered to be a single particle, find its molecular mass.* **(IIT 1999)**

Solution: Volume of cylindrical virus $= \pi r^2 l$

$$= 3.14 \times \left(\frac{150}{2} \times 10^{-8}\right)^2 \times 5000 \times 10^{-8}$$

$= 0.884 \times 10^{-16}$ cm^3

Mass of virus $= \frac{\text{Volume}}{\text{Specific volume}} = \frac{0.884 \times 10^{-16}}{0.75}$

$= 1.178 \times 10^{-16}$ g

Molar mass of virus = Mass of single virus $\times 6.023 \times 10^{23}$

$= 1.178 \times 10^{-16} \times 6.023 \times 10^{23}$

$= 7.095 \times 10^7$

Example 27. *Weighing 3104 carats (1 carat = 200 mg), the Cullinan diamond was the largest natural diamond ever found. How many carbon atoms were present in the stone?*

Solution: Mass of the stone

$= 3104 \times 200 = 620800$ mg $= 620.8$ g

Number of atoms of carbon

$= \frac{\text{Mass in gram}}{\text{Gram-atomic mass}} \times 6.023 \times 10^{23}$

$= \frac{620.8}{12} \times 6.023 \times 10^{23} = 3.12 \times 10^{25}$

Example 28. *A cylinder of compressed gas contains nitrogen and oxygen in the ratio 3 : 1 by mole. If the cylinder is known to contain 2.5×10^4 g of oxygen, what is the total mass of the gas mixture?*

Solution: Number of moles of oxygen in the cylinder

$= \frac{\text{Mass in gram}}{\text{Molecular mass in gram}} = \frac{2.5 \times 10^4}{32}$

$= 781.25$

∴ Number of moles of $N_2 = 3 \times 781.25 = 2343.75$

Mass of nitrogen in the cylinder $= 2343.75 \times 28$

$= 65625$ g

$= 6.5625 \times 10^4$ g

Total mass of the gas in the cylinder

$= 2.5 \times 10^4 + 6.5625 \times 10^4 = 9.0625 \times 10^4$ g

Example 29. *Atmospheric air has 78% N_2; 21% O_2; 0.9 % Ar and 0.1 % CO_2 by volume. What is the molecular mass of air in the atmosphere?*

Solution: Molecular mass of mixture

$= \frac{\Sigma \text{ \% of each}}{100} \times$ Molar mass

$= \frac{78}{100} \times 28 + \frac{21}{100} \times 32 + \frac{0.9}{100} \times 40 + \frac{0.1}{100} \times 44 = 28.964$

Example 30. *The famous toothpaste Forhans contains 0.76 g of sodium per gram of sodium monofluoroorthophosphate (Na_3PO_4F) in 100 mL.*

(a) How many fluorine atoms are present?

(b) How much fluorine in milligrams is present?

Solution:

Molar mass of $Na_3PO_4F = 3 \times 23 + 31 + 16 \times 4 + 19 = 183$

183 g Na_3PO_4F contains = 19 g fluorine

$\therefore$ 0.76 g Na_3PO_4F contains $= \frac{19}{183} \times 0.76$ g fluorine

$= 0.0789$ g $= 78.9$ mg fluorine

Number of fluorine atoms

$$= \frac{\text{Mass in gram}}{\text{Gram-atomic mass}} \times 6.023 \times 10^{23}$$

$$= \frac{0.0789}{19} \times 6.023 \times 10^{23}$$

$$= 2.5 \times 10^{21} \text{ atoms}$$

Example 31. *An alloy of iron (54.7%), nickel (45%) and manganese (0.3%) has a density of 8.17 g/cm^3. How many iron atoms are there in a block of alloy measuring 10 cm × 20 cm × 15 cm?*

Solution:

Volume of the block of alloy $= 10 \times 20 \times 15 \text{ cm}^3$

$= 3000 \text{ cm}^3$

Mass of the block $= 3000 \times 8.17 \text{ g} = 24510 \text{ g}$

Mass of iron in the block $= \frac{54.7}{100} \times 24510 = 13406.97$ g

Number of iron atoms in the block $= \frac{\text{Mass}}{\text{Atomic mass}} \times 6.023 \times 10^{23}$

$$= \frac{13406.97}{56} \times 6.023 \times 10^{23}$$

$$= 1.442 \times 10^{26}$$

Example 32. *An analysis of pyrex glass showed 12.9% B_2O_3, 2.2% Al_2O_3, 3.8% Na_2O, 0.4% K_2O and remaining is SiO_2. What is the ratio of silicon to boron atoms in the glass?* **(BCECE 2007)**

Solution:

Percentage composition of $B_2O_3 = 12.9\%$

Percentage composition of

$SiO_2 = 100 - [12.9 + 2.2 + 3.8 + 0.4]$

$= 80.7\%$

Number of moles of $B_2O_3 = \frac{\text{Mass}}{\text{Molar mass}} = \frac{12.9}{70} = 0.184$

Number of moles of boron atoms $= 2 \times 0.184$

Number of moles of $SiO_2 = \frac{\text{Mass}}{\text{Molar mass}} = \frac{80.7}{60} = 1.345$

Number of moles of silicon atoms = 1.345

$$\frac{\text{Number of atoms of silicon}}{\text{Number of atoms of boron}} = \frac{N_A \times 1.345}{N_A \times 0.184} = \frac{7.3}{1}$$

Where, N_A = Avogadro's number

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

22. x gram of $CaCO_3$ was completely burnt in air. The mass of the solid residue formed is 28 g. What is the value of 'x' in gram? **(EAMCET 2005)**

(a) 44 (b) 200 (c) 150 (d) 50

[Ans. (d)]

[Hint: $\underset{100\,g}{CaCO_3(s)} \longrightarrow \underset{56\,g}{CaO(s)} + CO_2(g)$

56 g residue ≡ 100 g $CaCO_3$

$\therefore$ 28 g residue ≡ 50 g $CaCO_3$]

23. The mass of carbon anode consumed (giving only carbon dioxide) in the production of 270 kg of Al metal from bauxite by Hall process is:

(a) 270 kg (b) 540 kg (c) 90 kg (d) 180 kg

[Ans. (c)]

[Hint: $\underset{3 \times 12\,g}{3C} + 2Al_2O_3 \longrightarrow \underset{4 \times 27 = 108\,g}{4Al} + 3CO_2$

$\because$ 108 g Al is produced by consuming = 36 g carbon

$\therefore$ 270×10^3 g Al will be produced by consuming

$= \frac{36}{108} \times 270 \times 10^3$ g carbon

$= 90 \times 10^3$ g = 90 kg carbon]

24. The equivalent mass of an element is 4. Its chloride has vapour density 59.25. Then the valency of the element is:

(a) 4 (b) 3 (c) 2 (d) 1

[Ans. (b)]

[Hint: Molecular mass of $MCl_n = 59.25 \times 2 = 118.5$

$a + 35.5 \times n = 118.5$ ... (i)

Equivalent mass $\times n + 35.5 \times n = 118.5$

$4n + 35.5n = 118.5$ ... (ii)

$\therefore$ $n = 3$]

25. Sulphur trioxide is prepared by the following two reactions:

$$S_8(s) + 8O_2(g) \longrightarrow 8SO_2(g)$$

$$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$$

How many grams of SO_3 are produced from 1 mole S_8?

(a) 1280 (b) 640

(c) 960 (d) 320

[Ans. (b)]

[Hint: From the given reaction, it is clear that, 1 mole S_8 will give 8 moles of SO_3.

$\therefore$ Mass of SO_3 formed will be $= 80 \times 8 = 640$ g.]

26. Calculate the number of millilitres at STP of H_2S gas needed to precipitate cupric sulphide completely from 100 mL of a solution containing 0.75 g of $CuCl_2$ in 1 L.

(a) 21.4 (b) 14.2

(c) 41.2 (d) 124

[Ans. (d)]

[**Hint:** $CuCl_2 + H_2S \longrightarrow CuS + 2HCl$

Number of moles of H_2S = Number of moles of $CuCl_2$

$$= \frac{0.75}{134.5} = 0.00557$$

Volume of $H_2S = 0.00557 \times 22400 = 124.8$ mL]

27. In the reaction,

$$As_2S_5 + xHNO_3 \longrightarrow 5H_2SO_4 + yNO_2 + 2H_3AsO_4 + 12H_2O$$

the values of x and y are: **[JEE (Orissa) 2006]**

(a) 40, 40 (b) 10, 10
(c) 30, 30 (d) 20, 20

[**Ans.** (a)]

[**Hint:** In RHS, there are 40 hydrogen atoms, hence only option (a) will be suitable.]

SUMMARY AND IMPORTANT POINTS TO REMEMBER

1. **Chemistry:** Branch of physical science which deals with the properties, composition and changes of matter. It has several branches. Main branches are (i) organic (ii) inorganic (iii) physical and (iv) analytical. It is wide in its scope and touches almost every aspect of our lives.
2. **Matter:** It is anything which has mass and occupies space. Matter exists in three physical states (i) solid (ii) liquid and (iii) gas. It is chemically classified into (a) elements (b) compounds and (c) mixtures.
3. **Energy:** The capacity of doing work. It is of various forms. One form can be converted into another but cannot be created or destroyed. The total amount of matter and energy available in the universe is constant. The relationship between mass and energy is given by Einstein equation, $E = mc^2$ (where, E = energy, m = mass, c = velocity of light).
4. **Intensive properties:** Do not depend on the quantity of matter, *e.g.*, colour, density, melting point, boiling point, etc.
5. **Extensive properties:** Depend on the quantity of matter, *e.g.*, volume, mass, weight, etc.
6. **Substance:** A variety of matter, all samples of which have the same composition and properties. Pure substances are divided into (i) elements and (ii) compounds.
7. **Element:** A substance which cannot be decomposed into anything more simpler by ordinary physical or chemical means. 117 elements are known. 88 elements have been isolated from natural sources and remaining 29 have been prepared by artificial means. Every element is represented by a **symbol** which is a small abbreviation of its full and lengthy name. Oxygen is the most abundant element. Silicon, aluminium, iron are second, third and fourth most abundant elements. Elements are classified as (i) metals (ii) non-metals and (iii) metalloids.
8. **Metals:** Generally solids (Hg—exception). They have properties such as lustre, hardness, malleable, ductile, good conductors of heat and electricity. Copper, zinc, iron, aluminium are metals.
9. **Non-metals:** Usually non-lustrous, brittle and poor conductors of electricity. Oxygen, carbon, nitrogen, chlorine, helium, etc., are non-metals.
10. **Metalloids:** Possess mixed properties of metals and non-metals both (*e.g.*, As, Sb, Sn).
11. **Compound:** Pure substance composed of two or more different elements in a fixed proportion of mass. The properties of a compound are altogether different from the properties of elements from which it has been constituted.
12. **Mixture:** A material containing two or more substances (elements or compounds) in any proportion, in which components do not lose their identity. Homogeneous mixture has a single phase while heterogeneous has more than one phase. Mixture can be separated into components by physical methods.
13. **Alloy:** A homogeneous mixture of two or more elements—metal and metal, metal and non-metal or non-metal and non-metal. They have unique properties.
14. **Physical change:** A temporary change, no change in chemical composition and mass. Physical properties alter. It can be reversed easily.
15. **Chemical change:** A permanent change, new substance is formed which possesses different composition and properties. It cannot be reversed easily. Chemical changes are of various types. The important ones are decomposition, synthesis, substitution, addition, internal rearrangement, polymerisation, double decomposition, etc.
16. **Law of conservation of mass:** (Lavoisier—1774) In a chemical change, mass is neither created nor destroyed. In chemical reactions:
    Total masses of reactants = Total masses of products.
17. **Law of constant proportions:** (Proust—1799) A chemical compound always contains the same element combined together in fixed proportion by mass.
18. **Law of multiple proportions:** (Dalton—1808) When two elements combine to form two or more compounds, the different masses of one element which combine with a fixed mass of the other element, bear a simple ratio to one another.
19. **Law of reciprocal proportions:** (Richter—1794) When two different elements combine with the same mass of a third element, the ratio in which they do so will be the same or simple multiple if both directly combine with each other. In all chemical reactions, substances react in the ratio of their equivalent masses.
20. **Law of gaseous volumes:** (Gay-Lussac—1808) Gases react with each other in simple ratio of their volumes and if product is also in gaseous state, its volume also bears a simple ratio with the volumes of gaseous reactants under similar conditions of temperature and pressure.

21. **Dalton's atomic theory:** Every element is composed of small indivisible, indestructible particles called atoms. Atoms of the same element are identical but differ in properties, mass and size of atoms of other elements. Atoms of different elements combine in simple ratio to form compounds. The relative number and kind of atoms are always the same in a given compound. Atoms cannot be created or destroyed.
22. **Atom:** The smallest particle of an element that takes part in a chemical reaction.
23. **Molecule:** The smallest particle of an element or compound that can have a stable existence.
24. **Formula:** Group of symbols of elements which represents one molecule of a substance. It represents also the chemical composition.
25. **Atomic mass:** Atomic mass of an element is the ratio of mass of one atom of an element to $\frac{1}{12}$th part of the mass of carbon-12.

Atomic mass of an element

$$= \frac{\text{Mass of one atom of the element}}{\text{Mass of one atom of carbon-12}} \times 12$$

26. **Atomic mass unit (amu):** $\frac{1}{12}$th mass of carbon-12. It is equal to 1.66×10^{-24} g.

Atomic mass of an element

$$= \frac{\text{Mass of one atom of the element}}{1\,\text{amu}}$$

The actual mass of an atom of element = Atomic mass in amu $\times 1.66 \times 10^{-24}$ g.

The atomic masses of elements are actually average relative masses because elements occur as mixture of isotopes.

27. **Gram-atomic mass or Gram atom:** Atomic mass expressed in grams. It is the absolute mass in grams of 6.02×10^{23} atoms of any element.

$$\text{No. of gram atoms} = \frac{\text{Mass of element in grams}}{\text{Atomic mass of the element in grams}}$$

28. **Molecular mass:** It indicates how many times one molecule of a substance is heavier in comparison to $\frac{1}{12}$th of mass of one atom of carbon-12. Mass of a molecule is equal to sum of masses of the atoms present in a molecule.
29. **Gram-molecular mass or Gram molecule:** Molecular mass expressed in gram. It is the absolute mass in gram of 6.02×10^{23} molecules of any substance.

No. of gram molecules

$$= \frac{\text{Mass of a substance in gram}}{\text{Molecular mass of the substance in gram}}$$

30. **Avogadro's hypothesis:** Under similar conditions of temperature and pressure, equal volumes of all gases contain same number of molecules.
31. **Gram molar volume:** The volume occupied by one gram-molecular mass of any gas at NTP (0°C or 273 K and one atm or 76 cm of Hg as pressure). Its value is 22.4 litre.
32. **Vapour density:**

$$\text{V.D.} = \frac{\text{Density of a gas}}{\text{Density of hydrogen}}$$

$$= \frac{\text{Mass of a certain volume of a gas}}{\text{Mass of same volume of hydrogen under same temperature and pressure}}$$

$$2\,\text{V.D.} = \text{Molecular mass}$$

33. **Mole:** A mole (mol) is defined as the number of atoms in 12.0 g of carbon-12. The number of atoms is 6.02×10^{23}. This number is called Avogadro's number.

$$\text{No. of moles} = \frac{\text{Mass of substance in gram}}{\text{Mass of one mole of the substance in gram}}$$

$$= \frac{\text{No. of particles}}{6.02 \times 10^{23}}$$

$$= \frac{\text{Volume of gas in litres at NTP}}{22.4}$$

Mass of one atom of an element

$$= \frac{\text{Gram atom of an element}}{6.02 \times 10^{23}}$$

Mass of one molecule of a substance

$$= \frac{\text{Gram-molecular mass of a substance}}{6.02 \times 10^{23}}$$

34. **Equivalent mass:** The number of parts by mass of the substance which combine or displace 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 or 108 parts by mass of silver.

The equivalent mass of an element may vary with change of valency.

Eq. mass of an element

$$= \frac{\text{Mass of element}}{\text{Mass of hydrogen}} \times 1.008$$

$$= \frac{\text{Mass of element} \times 11200}{\text{Volume in mL of hydrogen displaced at NTP}}$$

$$= \frac{\text{Mass of element}}{\text{Mass of oxygen}} \times 8$$

$$= \frac{\text{Mass of element}}{\text{Mass of chlorine}} \times 35.5$$

35. **Metal to metal displacement:** $\frac{m_1}{m_2} = \frac{E_1}{E_2}$
36. **Double decomposition:** $AB + CD \rightarrow \underset{\text{ppt.}}{AD} + CB$

$$\frac{\text{Mass of } AB}{\text{Mass of } AD} = \frac{\text{Eq. mass of } A + \text{Eq. mass of } B}{\text{Eq. mass of } A + \text{Eq. mass of } D}$$

Atomic mass of an element

$$= \text{Eq. mass of the element} \times \text{Valency}$$

37. Dulong and Petit's law:

$$\text{Atomic mass (approximate)} = \frac{6.4}{\text{Specific heat}}$$

38. Cannizzaro's method: Atomic mass of an element is the smallest mass of the element present in the molecular mass of any one of its compounds.

39. Law of isomorphism: Isomorphous compounds form crystals which have same size and shape and can grow in the saturated solution of each other.

Masses of two elements that combine with same mass of other elements in their respective compounds are in the ratio of their atomic masses.

40. Atomic mass from vapour density of a chloride:

$$\text{Valency of an element} = \frac{2\,\text{V.D. of a volatile chloride}}{\text{Eq. mass} + 35.5}$$

41. Types of formulae:

(i) **Empirical:** It represents the simplest relative whole number ratio of atoms of each element present in the molecule of a substance.

(ii) **Molecular:** It represents the actual number of atoms of each element present in one molecule of a substance.

Molecular formula = $n \times$ Empirical formula

$$n = \frac{\text{Molecular formula mass}}{\text{Empirical formula mass}}$$

(iii) **Structural:** It represents the way in which atoms of various elements are linked with each other.

42. Percentage of element:

$$\text{Percentage of element} = \frac{\text{Mass of element}}{\text{Molecular mass}} \times 100$$

43. Chemical equation: It is a symbolic representation of a chemical change. The equation becomes balanced when total number of atoms of various elements are made equal on both the sides of equation. Chemical equations are of two types (i) molecular and (ii) ionic. Chemical equation is based on law of conservation of mass.

44. Unit: It is the primary standard chosen to measure any physical quantity.

The seven units of measurement, namely mass, length, time, temperature, electric current, luminous intensity and amount of substance are taken as basic units. All other units can be derived from them and are, therefore, called derived units. SI units are used these days in all branches of science.

45. Significant figure: It is the total number of certain digits plus one doubtful digit.

Questions

1. Match the following, choosing one item from Column-X and the appropriate related item from Column-Y.

[A]

Column-X	**Column-Y**
(a) Efflorescence	(i) Homogeneous mixture
(b) Malleability	(ii) Heterogeneous mixture
(c) Alloy	(iii) Mole
(d) 1 amu	(iv) (1/12)th mass of carbon-12
(e) Sulphur and sand	(v) Tendency to lose water of crystallisation
(f) Amount of substance	(vi) Property of metal being hammered into thin sheets

[B]

Column-X	**Column-Y**
(a) Equal volumes of all gases contain equal number of molecules at NTP.	(i) Dalton's atomic theory
(b) The atom is indestructible.	(ii) Law of conservation of mass
(c) All pure samples of the same compound contain the same elements combined in the same proportion by mass.	(iii) Avogadro's law
(d) Total mass before and after the chemical reaction is same.	(iv) Dulong and Petit's law
(e) Atomic mass $= \dfrac{6.4}{\text{Specific heat}}$	(v) Gay-Lussac's law
(f) Gases react in simple ratio of their volumes.	(vi) Law of constant proportions

[C]

Column-X	**Column-Y**
(a) Most abundant element	(i) Platinum
(b) Most abundant metal	(ii) Diamond
(c) Liquid at room temp.	(iii) Aluminium
(d) Hardest substance	(iv) Plutonium
(e) Most ductile metal	(v) Mercury
(f) Transuranic element	(vi) Oxygen

2. **Matrix Matching Problems:**
(According to the new pattern of IIT Screening)

[A] Match the Column-X and Column-Y:

Column-X	**Column-Y**
(a) Vapour density	(i) Unitless
(b) Mole	(ii) 1 mol electrons
(c) 12 g carbon	(iii) Collection of 6.023×10^{23} atoms
(d) 96500 C	(iv) Molecular mass $\times \frac{1}{2}$

[B] Match the Column-X and Column-Y:

Column-X	**Column-Y**
(a) 1.6 g CH_4	(i) 0.1 mol
(b) 1.7 g NH_3	(ii) 6.023×10^{23} electrons
(c) HCHO	(iii) 40% carbon
(d) $C_6H_{12}O_6$	(iv) Vapour density = 15

[C] Match the Column-X and Column-Y:

Column-X	**Column-Y**
(a) 1 amu	(i) Heaviest particle of atom
(b) Proton	(ii) 1.66×10^{-27} kg
(c) Neutron	(iii) 931.5 MeV
(d) α-particle	(iv) Positively charged

[D] Match the Column-X with Column-Y for the reaction:

$$A + B_2 \rightarrow AB_2$$

Column-X	**Column-Y**
(a) 300 atoms of A + 200 molecules of B_2	(p) B_2 is limiting reagent
(b) 100 atoms of A + 100 molecules of B_2	(q) A is limiting reagent
(c) 5 mol of A + 2.5 mol of B_2	(r) None of the reactant is in excess
(d) 2.5 mol of A + 5 mol of B_2	(s) 200 molecules of AB_2 will be formed

Answers

1. [A] (a–v); (b–vi); (c–i); (d–iv); (e–ii); (f–iii)
   [B] (a–iii); (b–i); (c–vi); (d–ii); (e–iv); (f–v)
   [C] (a–vi); (b–iii); (c–v); (d–ii); (e–i); (f–iv).
2. [A] (a–i, iv); (b–iii); (c–iii); (d–ii)
   [B] (a–i, ii); (b–i, ii); (c–iii, iv); (d–iii)
   [C] (a–ii, iii); (b–ii, iv); (c–i); (d–iv)
   [D] (a–p, s); (b–r); (c–p); (d-q).

• PRACTICE PROBLEMS •

1. The density of mercury is 13.6 g/mL. Calculate the diameter of an atom of mercury assuming that each atom of mercury is occupying a cube of edge-length equal to the diameter of mercury atom.
   (Atomic mass of mercury = 200)
   [**Ans.** 2.9×10^{-8} cm]
2. A metal M of atomic mass 54.94 has a density of 7.42 g/cc. Calculate the apparent volume occupied by one atom of the metal.
   [**Ans.** 1.23×10^{-23} cc]
3. Find the charge of 1 g ion of N^{3-} in coulomb.
   [**Ans.** 2.894×10^5 coulomb]
4. Calculate the volume at NTP occupied by 6.25 g of nitrogen.
   [**Ans.** 5.0 litre]
5. 10 mL of hydrogen contains 2×10^3 molecules of hydrogen at certain pressure and temperature. Calculate the number of molecules of oxygen whose volume is 200 mL at the same temperature and pressure.
   [**Ans.** 4×10^4 molecules]
6. The masses of equal volumes of a gas and hydrogen are 25.6 g and 0.8 g respectively under same conditions of temperature and pressure. Find the molecular mass of the gas.
   [**Hint:** V.D. of the gas $= \frac{25.6}{0.8} = 32.0$
   Molecular mass $= 2$ V.D. $= 2 \times 32.0 = 64.0$]
7. One litre of a gas at NTP weighs 1.97 g. Find the molecular mass of gas.
   [**Ans.** 44.128]
8. How many moles of water are present in one litre of water?
   [**Ans.** 55.5 moles]
9. Calculate the mass of 6.02×10^{21} molecules of nitrogen.
   [**Ans.** 0.28 g]
10. 1.5276 g of $CdCl_2$ was found to contain 0.9367 g of cadmium. Calculate the atomic mass of cadmium.
    [**Ans.** 112.54]
    [**Hint:** Equivalent mass of cadmium $= \frac{\text{Mass of Cd}}{\text{Mass of Cl}} \times 35.5$
    $= \frac{0.9367}{0.5909} \times 35.5 = 56.27$
    Atomic mass = Equivalent mass × Valency]
11. Calculate how many methane molecules and how many hydrogen and carbon atoms are there in 25.0 g of methane? **(MLNR 1990; Dhanbad 1992)**
    [**Ans.** 9.41×10^{23} CH_4 molecules, 9.41×10^{23} carbon atoms and 37.64×10^{23} hydrogen atoms.]
    [**Hint:** No. of moles of methane $= \frac{25.0}{16.0}$
    One molecule of methane contains one carbon atom and four hydrogen atoms.]
12. How much sugar ($C_{12}H_{22}O_{11}$) will be required if each person on the earth is given 100 molecules of sugar? The population of the earth is 3×10^{10}.
    [**Ans.** 170.43×10^{-11} g]
13. A mixture of hydrogen and oxygen contains 20% by mass of hydrogen. What is the total number of molecules present per gram of the mixture?
    [**Ans.** 7.528×10^{22}]
    [**Hint:** In 1 gram of the mixture, 0.2 g of hydrogen and 0.8 g of oxygen are present. Moles of $H_2 = \frac{0.2}{2} = 0.1$, moles of oxygen $= \frac{0.8}{32} = 0.025$. Calculate the number of molecules of hydrogen and oxygen and then add.]
14. How many electrons are present in 18 mL of water? **(MLNR 1995)**
    [**Hint:** 18 mL water = 18 g water = 1 mole water $= 6.02 \times 10^{23}$ molecules, each molecule consists 10 electrons (8 electrons per oxygen atom, 2 electrons for two hydrogen atoms). Total electrons $= 10 \times 6.02 \times 10^{23} = 6.02 \times 10^{24}$]
15. Sulphur molecule is known to be composed of 8 atoms of the element. In a sample of 192 g of pure sulphur, calculate (i) number of g-atoms of sulphur; (ii) number of atoms of sulphur; (iii) number of moles of sulphur; (iv) number of molecules of sulphur.
    [**Ans.** g-atoms = 6; No. of atoms $= 6 \times 6.02 \times 10^{23}$; No. of moles = 0.75; No. of molecules $= 4.52 \times 10^{23}$]
    [**Hint:** The atomic mass of sulphur is 32.]
16. The vapour density of a mixture containing NO_2 and N_2O_4 is 38.3 at 27°C. Calculate the moles of NO_2 in 100 g of the mixture. **(MLNR 1993)**
    [**Hint:** Mol. mass of mixture $= 2 \times 38.3 = 76.6$
    No. of moles in 100 g of mixture $= \frac{100}{76.6}$
    Let a g of NO_2 is present in mixture.

Moles of NO_2 + Moles of N_2O_4 = Moles of mixture

$$\frac{a}{46}+\frac{100-a}{92}=\frac{100}{76.6} \quad \text{or} \quad a = 20.10 \text{ g}$$

Moles of NO_2 in mixture $=\frac{20.10}{46}=0.437$]

17. Calculate the number of oxygen atoms in 88 g CO_2. What would be the mass of CO having the same number of oxygen atoms? (BITS 1990)

[**Hint:** 88 g CO_2 = 2 moles of CO_2. One molecule consists of 2 oxygen atoms.

No. of oxygen atoms $= 2 \times 2 \times 6.02 \times 10^{23} = 24.08 \times 10^{23}$

CO molecule has one oxygen atom.

Mass of CO containing 24.08×10^{23} oxygen atoms

$$=\frac{28}{6.02 \times 10^{23}} \times 24.08 \times 10^{23} = 112 \text{ g}$$]

18. Density of water at room temperature is 1.0 g cm^{-3}. How many molecules are there in one drop of water if its volume is 0.1 cm^3?

[**Ans.** 3.34×10^{21} molecules]

[**Hint:** Mass of one drop = Vol. $\times d = 0.1 \times 1 = 0.1$ g

No. of moles $=\frac{0.1}{18}$; No. of molecules $= 6.02 \times 10^{23} \times \frac{0.1}{18}$]

19. Naturally occurring boron consists of two isotopes, whose atomic masses are 10.01 and 11.01. The atomic mass of natural boron is 10.81. Calculate the percentage of each isotope in natural boron. (MLNR 1994)

[**Ans.** % of isotope with atomic mass 10.01 = 20; % of isotope with atomic mass 11.01 = 80]

[**Hint:** Let x be the percentage of the isotope with atomic mass 10.01.

$$\frac{10.01 \times x}{100}+\frac{11.01(100-x)}{100}=10.81 \quad \text{or} \quad x = 20$$]

20. Chlorine has isotopes ^{35}Cl and ^{37}Cl. There are three ^{35}Cl isotopes for every ^{37}Cl isotope in a sample of chlorine. Calculate the atomic mass of chlorine.

[**Ans.** $A=\frac{3 \times 35 + 37 \times 1}{4}=35.5$]

21. Natural hydrogen gas is a mixture of ^1H and ^2H in the ratio of 5000 : 1. Calculate the atomic mass of the hydrogen.

[**Ans.** 1.000199]

22. Chromium has the following isotopic composition:

Mass number	Isotopic mass	Fractional abundance
50	49.9461	x
52	51.9405	0.8379
53	52.9407	0.0950
54	53.9389	0.0236

Calculate the value of x.

[**Ans.** 0.0435]

23. Use the data given in the following table to calculate the molar mass of naturally occurring argon:

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}	99.6%

[**Ans.** 39.947]

24. Density of oxygen at NTP is 1.429 g/litre. Calculate the standard molar volume of the gas.

[**Ans.** 22.39 litre mol^{-1}]

25. How many iron atoms are present in a stainless steel ball bearing having a radius of 0.254 cm? The stainless steel contains 85.6% Fe by weight and has density of 7.75 g/cm^3.

[**Ans.** 4.91×10^{21}]

26. The nucleus of an atom X is supposed to be a sphere with a radius of 5×10^{-13} cm. Find the density of the matter in the atomic nucleus if the atomic weight of X is 19.

[**Ans.** 6.02×10^{13} g/mL]

27. Calculate the number of atoms of each element present in 122.5 g of $KClO_3$.

[**Ans.** Number of atoms of 'K' $= 1 \times 6.023 \times 10^{23}$

Number of atoms of 'Cl' $= 1 \times 6.023 \times 10^{23}$

Number of atoms of 'O' $= 3 \times 6.023 \times 10^{23}$]

28. In an experiment, 1.0 g $CaCO_3$ on heating evolved 224 mL of CO_2 at NTP. What mass of CaO (calcium oxide) is formed?

[**Ans.** Mass of CaO = 0.56 g]

[**Hint:** Mass of 224 mL of $CO_2=\frac{44}{22400} \times 224 = 0.44$ g]

29. What mass of potassium chlorate ($KClO_3$) on heating gives 1.491 g of potassium chloride (KCl) and 0.672 litres of oxygen at NTP?

[**Ans.** Mass of $KClO_3$ = 2.451 g]

[**Hint:** Mass of 22.4 litre of oxygen at NTP = 32 g]

30. A compound AB completely decomposes into A and B on heating. 50 g of AB, on strong heating, gave 40 g of A. How much quantity of AB should be decomposed by heating to obtain 2.5 g of B? How much quantity of A will be produced in the process?

[**Ans.** 12.5 g AB is to be decomposed, 10.0 g of A will be produced.]

[**Hint:** $\underset{50\text{ g}}{AB} \rightarrow \underset{40\text{ g}}{A} + \underset{10\text{ g}}{B}$]

31. If 12.6 g of $NaHCO_3$ is added to 20.0 g of HCl solution, the residue solution is found to weigh 24.0 g. What is the mass and volume of CO_2 released at NTP in the reaction?

[**Ans.** 8.6 g CO_2 released. Volume at NTP $=\frac{22.4}{44} \times 8.6 = 4.378$ litre]

32. (i) 5.06 g of pure cupric oxide (CuO), on complete reduction by heating in a current of hydrogen, gave 4.04 g of metallic copper.

(ii) 1.3 g of pure metallic copper was completely dissolved in nitric acid and the resultant solution was carefully dried and ignited. 1.63 g CuO was produced in the process. Show that these results illustrate the law of constant proportions.

[**Ans.** In both cases, the ratio of copper and oxygen is 1 : 0.25. Hence, the law of constant proportions is illustrated.]

33. Metal M and chlorine combine in different proportions to form two compounds A and B. The mass ratio M : Cl is 0.895 : 1 in A and 1.791 : 1 in B. What law of chemical combination is illustrated?

[**Ans.** Masses of metal which combine with 1 part of chlorine are in the ratio of 1 : 2, which is a simple ratio. Hence, law of multiple proportions is illustrated.]

34. 2.8 g of calcium oxide (CaO) prepared by heating limestone were found to contain 0.8 g of oxygen. When one gram of oxygen was treated with calcium, 3.5 g of calcium oxide was obtained. Show that the results illustrate the law of definite proportions.

35. By means of the given analytical results show that law of multiple proportions is true:

Mercurous chloride	Mercuric chloride
Mercury = 84.92 %	Mercury = 73.80%
Chlorine = 15.08 %	Chlorine = 26.20%

[**Ans.** The masses of mercury which combine with 1 part of chlorine are in the ratio of 2 : 1, which is a simple ratio. Hence, law of multiple proportions is illustrated.]

36. 1 g of a metal, having no variable valency, produces 1.67 g of its oxide when heated in air. Its carbonate contains 28.57% of the metal. How much oxide will be obtained by heating 1 g of the carbonate?

[**Ans.** 0.477 g]

[**Hint:** $\dfrac{\text{Mass of metal}}{\text{Mass of oxygen}} = \dfrac{\text{Mass of metal in 1g of carbonate}}{x}$

i.e., $x = 0.1914$ g of oxygen

Mass of oxide $= 0.2857 + 0.1914 = 0.4771$ g]

37. 0.36 g of Mg combines with chlorine to produce 1.425 g of magnesium chloride. 9.50 g of another sample of anhydrous magnesium chloride gave, on electrolysis 2.24 litre of chlorine at NTP. Show that these data agree with the law of constant proportions.

[**Hint:** Mass of 2.24 litre of chlorine at NTP $= \dfrac{71}{22.4} \times 2.24$ $= 7.1$ g. In both cases, the ratio of masses of Mg and Cl is 1 : 3. Hence, law of constant proportions is followed.]

38. Carbon dioxide contains 27.27% carbon, carbon disulphide contains 15.97% carbon and sulphur dioxide contains 50% sulphur. Show that these figures illustrate the law of reciprocal proportions.

[**Hint:** The masses of oxygen and sulphur which combine with 1 part of carbon are in the ratio of 2.667 : 5.25, *i.e.*, 1 : 2. In sulphur dioxide, the masses of sulphur and oxygen are in the ratio of 1 : 1 which is a simple multiple of first. Hence, law of reciprocal proportions is illustrated.]

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

[**Hint:** The ratio of the masses of chlorine which combines with a fixed mass of phosphorus in two compounds is 3 : 5 which is a simple whole number ratio. Thus, the data illustrate law of multiple proportions.]

40. A and B are two hydrocarbons. A and B are heated separately in excess of oxygen when 0.028 g of A gave 44.8 mL CO_2 and 0.044 g of B gave 67.2 mL CO_2 at NTP. Show that the results are in agreement with law of multiple proportions.

[**Hint:** Determine the masses of CO_2 at NTP and then masses of carbon.

(A) Mass of $CO_2 = \dfrac{44}{22400} \times 44.8 = 0.088$ g,

Mass of carbon = 0.024 g, mass of hydrogen = 0.004 g.

(B) Mass of $CO_2 = \dfrac{44}{22400} \times 67.2 = 0.132$ g,

Mass of carbon = 0.036 g, mass of hydrogen = 0.008 g.

Thus, the masses of carbon combining with same mass of hydrogen are in the ratio of 4 : 3 which is a simple ratio. Hence, law of multiple proportions is followed.]

41. Aluminium oxide contains 52.9% aluminium and carbon dioxide contains 27.27% carbon. Assuming the validity of the law of reciprocal proportions, calculate the percentage of aluminium in aluminium carbide.

[**Hint:** From the data, it is observed that the ratio of masses of aluminium and carbon in aluminium carbide should be 3 : 1 or its simple multiple. Hence, percentage of aluminium in aluminium carbide $= \dfrac{3}{4} \times 100 = 75.0$]

42. Two volumes of ammonia, on dissociation gave one volume of nitrogen and three volumes of hydrogen. How much hydrogen will be obtained from dissociation of 40 mL of NH_3?

[**Ans.** 60 mL]

43. The following results were obtained by heating different oxides of lead in a current of hydrogen:

(a) 1.393 g of litharge gave 1.293 g of lead.

(b) 2.173 g of lead peroxide gave 1.882 g of lead.

(c) 1.721 g of red lead gave 1.552 g of lead.

Show that these results are in accordance with the law of multiple proportions.

[**Ans.** Masses of lead that combine with same mass of oxygen are in the ratio of 4 : 2 : 3 which is a simple ratio. So, the results are in accordance with the law of multiple proportions.]

44. Calculate the number of g-moles of CaO that could be obtained from 42.54 g of $CaCO_3$ and convert the number of g-moles to grams.

[No. of g-moles $= \dfrac{42.54}{100} = 0.4254$,

Mass of CaO $= 0.4254 \times 56 = 23.8$ g]

45. 1 g of a metal M which has specific heat of 0.06 combines with oxygen to form 1.08 g of oxide. What is the atomic mass of M?

[**Hint:** Approximate atomic mass $= \frac{6.4}{0.06} = 106.6$

$$\text{Equivalent mass of } M = \frac{1}{0.08} \times 8 = 100$$

$$\text{Valency} = \frac{106.6}{100} \approx 1$$

$$\text{Exact atomic mass} = 100 \times 1 = 100]$$

46. A compound contains 28% of nitrogen and 72% metal by mass. 3 atoms of the metal combine with 2 atoms of the nitrogen. Find the atomic mass of the metal.

[**Hint:** Valency of metal = 2 and valency of nitrogen = 3

Equivalent mass of nitrogen $= \frac{14}{3}$; $\frac{\text{Eq. mass of metal}}{14/3} = \frac{72}{28}$

Equivalent mass of metal = 12

Atomic mass of metal $= 12 \times 2 = 24$]

47. The chloride of a solid metallic element contains 57.89% by mass of the element. The specific heat of the element is 0.0324 cal deg^{-1} g^{-1}. Calculate the exact atomic mass of the element.

[**Hint:** Equivalent mass of the element $= \frac{57.89}{42.11} \times 35.5 = 48.8$

$$\text{Approximate atomic mass} = \frac{6.4}{0.0324} = 200$$

$$\text{Valency} = \frac{200}{48.8} \approx 4$$

$$\text{Exact atomic mass} = 48.8 \times 4 = 195.2\ ]$$

48. Two oxides of a metal contain 63.2% and 69.62% of the metal. The specific heat of the metal is 0.117. What are the formulae of the two oxides?

[**Ans.** MO_2 and M_2O_3]

49. White vitriol (hydrated zinc sulphate) is isomorphous with $MgSO_4 \cdot 7H_2O$. White vitriol contains 22.95% zinc and 43.9% of water of crystallisation. Find the atomic mass of zinc.

[**Hint:** The formula of white vitriol should be $ZnSO_4 \cdot 7H_2O$ as it is isomorphous to $MgSO_4 \cdot 7H_2O$, *i.e.*, 7 water molecules are associated with one zinc atom. $7H_2O = 7 \times 18 = 126$. Mass of Zn with which 126 parts of water by mass are associated $= \frac{22.95}{43.90} \times 126 = 65.87$. Atomic mass of zinc.]

50. A solid element burns in oxygen without any change in volume (of gas) under similar conditions of temperature and pressure. If the vapour density of pure gaseous product is 32, what is the equivalent mass of the element?

[**Hint:** One vol. of oxide contains 1 vol. of O_2.

One mole of oxide contains one mole of O_2.

$$\text{Mol. mass of oxide} = A + 32 = 2\ \text{V.D.} = 64$$

So, $A = 32$

32 parts of element combine with 32 parts of oxygen.

So, Equivalent mass of element $= \frac{32}{32} \times 8 = 8$]

51. If the equivalent mass of a metal (M) is x and the formula of its oxide is $M_m\,O_n$, then show that the atomic mass of M is $\frac{2xn}{m}$.

[**Hint:** m atoms of M combine with n atoms of oxygen.

1 atom of M combines with $\frac{n}{m}$ atoms of oxygen.

Hence, Valency $= \frac{2n}{m}$

Atomic mass = Equivalent mass × Valency

$$= x \times \frac{2n}{m} = \frac{2nx}{m}\ ]$$

52. Two oxides of metals A and B are isomorphous. The metal A whose atomic mass is 52, forms a chloride whose vapour density is 79. The oxide of the metal B contains 47.1% oxygen. Calculate the atomic mass of B.

[**Hint:** Let the valency of A be x. The formula of chloride $= ACl_x$

$$2\ \text{V.D.} = A + x \times 35.5 \qquad \text{or} \qquad x \approx 3$$

As the two oxides are isomorphous, the valency of B is also 3. Equivalent mass of $B = \frac{52.9}{47.1} \times 8 = 8.99$, atomic mass of $B = 8.99 \times 3 = 26.97$]

53. A mixture of 1.65×10^{21} molecules of X and 1.85×10^{21} molecules of Y weighs 0.688 g. If molecular mass of Y is 187, what is the molecular mass of X ?

[**Hint:** $\frac{A \times 1.65 \times 10^{21}}{6.02 \times 10^{23}} + \frac{187 \times 1.85 \times 10^{21}}{6.02 \times 10^{23}} = 0.688, A = 41.35$]

54. The equivalent mass of a metal is 29.73 and the vapour density of its chloride is 130.4. Find out the atomic mass of the metal.

[**Ans.** Atomic mass = 118.92]

[**Hint:** Valency $= \frac{2 \times \text{V.D.}}{\text{Eq. mass} + 35.5} = \frac{2 \times (130.4)}{(29.73 + 35.5)} \approx 4$]

55. Calculate the percentage of aluminium, sulphate radical and water in potash alum.

[**Ans.** Al = 5.69%; SO_4^{2-} = 40.51% ; Water = 45.57%]

56. Carbohydrates are represented by the general formula $C_m(H_2O)_n$. On heating, in absence of air, they decompose into steam (H_2O) and carbon. 3.1 g of a carbohydrate, on complete decomposition by heating in absence of air, leave a residue of 1.24 g of carbon. If the molecular mass of the carbohydrate be 180, find its molecular formula.

[**Hint:** Determine % of carbon in carbohydrate. It is 40%. Water is 60%. Empirical formula = CH_2O. Molecular formula $= 6 \times CH_2O = C_6H_{12}O_6$.]

57. A gaseous hydrocarbon contains 85.7% carbon and 14.3% hydrogen. 1 litre of the hydrocarbon weighs 1.26 g at NTP. Determine the molecular formula of the hydrocarbon.

[**Ans.** C_2H_4]

58. Equal masses of oxygen, hydrogen and methane are taken in a container under identical conditions. Find the ratio of their volumes.

[**Ans.** 1 : 16 : 2]

59. How many moles are there in 1 m^3 of any gas at NTP?
[**Ans.** 44.6 moles]

60. A hydrated chloride of metal contains 18.26% metal and 32.42% chloride ion by mass. The specific heat of metal is 0.16. What is hydrated chloride?
[**Ans.** $CaCl_2 \cdot 6H_2O$]

61. 1.878 g of MBr_x when heated in a stream of HCl gas was completely converted to chloride MCl_x which weighed 1.0 g. The specific heat of metal is 0.14 cal g^{-1}. Calculate the molecular masses of metal bromide and metal chloride.
[**Ans.** Mol. mass of metal bromide = 285.54;
Mol. mass of metal chloride = 152.2]

62. An automobile antifreeze consists of 38.7% C; 9.7% H and remaining oxygen by weight. When 0.93 g of it are vaporised at 200°C and 1 atm pressure, 582 mL of vapour are formed. Find the molecular formula of the antifreeze.
[**Ans.** $C_2H_6O_2$]

63. A mineral contained MgO = 31.88 %; SiO_2 = 63.37% and H_2O = 4.75%. Show that the simplest formula for the mineral is $H_2Mg_3Si_4O_{12}$.
(H = 1; Mg = 24; Si = 28; O = 16)

64. How many moles of NH_3 are there in 250 cm^3 of a 30% solution, the specific gravity of which is 0.90?
[**Ans.** 3.97 moles]

65. Haemoglobin contains 0.25% iron by mass. The molecular mass of haemoglobin is 89600. Calculate the number of iron atoms per molecule of haemoglobin.
[Atomic mass of Fe = 56]
[**Ans.** 4 atoms]

66. A sample of potato-starch was ground to give a starch like molecule. The product analysed 0.086% phosphorus. If each molecule is assumed to contain one atom of phosphorus, what is the average molecular mass of the material?
[**Ans.** 36000 amu]

67. Insulin contains 3.4% sulphur. Calculate minimum molecular mass of the insulin.
[**Ans.** 941.176 amu]
[**Hint:** For minimum molecular mass, one molecule of insulin must have atleast one sulphur atom.]

68. Calculate the number of carbon, hydrogen and oxygen atoms in 18 g of glucose.
[**Ans.** 3.61×10^{23} carbon atoms, 7.22×10^{23} hydrogen atoms, 3.61×10^{23} oxygen atoms]

69. Hydrated sulphate of a divalent metal of atomic weight 65.4 loses 43.85% of its weight on dehydration. Find the number of molecules of water of crystallisation in the formula of hydrated salt. **[JEE (West Bengal) 2005]**

[**Hint:** Formula of divalent hydrated metal sulphate will be

$$MSO_4 \cdot xH_2O$$

Molecular mass of salt = 65.4 + 96 + 18x
= (161.4 + 18x)

$$\% \text{ of water} = \frac{18x}{161.4 + 18x} \times 100 = 43.85$$

On solving, $x = 7$

$\therefore$ Molecular formula of hydrated salt = $MSO_4 \cdot 7H_2O$]

70. A person with fever has a temperature of 102.5°F. What is the temperature in degree celsius?
[**Hint:** Use $C = \frac{5(F-32)}{9}$]

71. An ornamental ring contains 275 carats of diamond. How many grams diamond does it have?
[**Hint:** 1 carat = 200 mg
$\therefore$ Mass of diamond = $275 \times 200 \times 10^{-3}$ g]

72. 1 volume of a gaseous compound consisting C, H, O on complete combustion in presence of 2.5 volume of O_2 gives 2 vol. of steam and 2 vol. of CO_2. What is the formula of the compound if all measurements are made at NTP?
[**Ans.** C_2H_4O]

73. 60 mL of a mixture of nitrous oxide and nitric oxide was exploded with excess of hydrogen. If 38 mL of N_2 was formed, calculate the volume of each gas in the mixture.
[**Ans.** NO = 44 mL and N_2O = 16 mL]

74. For a precious stone, 'carat' is used for specifying its mass. If 1 carat = 3.168 grains (a unit of mass) and 1 gram = 15.4 grains, find the total mass in kilogram of the ring that contains 0.5 carat diamond and 7 gram gold.
[**Ans.** 7.1×10^{-3} kg]

75. The density of a gaseous element is 5 times that of oxygen under similar conditions. If the molecule of the element is triatomic, what will be its atomic mass?
[**Ans.** 53.33]

76. Calculate the number of electrons, protons and neutrons in 1 mole of $^{18}O^{2-}$ ions.
[**Ans.** Electrons = $10 \times 6.023 \times 10^{23}$
Protons = $8 \times 6.023 \times 10^{23}$
Neutrons = $8 \times 6.023 \times 10^{23}$]

77. 600 mL of a mixture of O_2 and O_3 weighs 1 gm at NTP. Calculate the volume of ozone in the mixture.
[**Ans.** 200 mL]

OBJECTIVE QUESTIONS

Set-1: Questions with single correct answer

1. The father of modern chemistry is:
(a) Priestley (b) Lavoisier
(c) Dalton (d) Mendeleev
2. A pure substance can only be:
(a) a compound
(b) an element
(c) an element or a compound
(d) a heterogeneous mixture
3. A pure substance which contains only one type of atoms is called:
(a) an element (b) a compound
(c) a solid (d) a liquid
4. Which one of the phrases would be incorrect to use?
(a) A mole of an element (b) A mole of a compound
(c) An atom of an element (d) An atom of a compound
5. A symbol not only represents the name of the element but also its:
(a) atomic mass (b) atomic number
(c) atomicity (d) atomic volume
6. The credit for the discovery of transuranic elements goes to:
(a) Hahn (b) Rutherford
(c) Seaborg (d) Curie
7. The most abundant metal in earth's crust is:
(a) iron (b) magnesium
(c) calcium (d) aluminium
8. The most abundant element in earth's crust is:
(a) hydrogen (b) oxygen (c) nitrogen (d) silicon
9. Which one of the elements is not found in nature?
(a) Radium (b) Technetium
(c) Polonium (d) Helium
10. Which one of the following is not a compound?
(a) Marble (b) Ozone
(c) Carborundum (d) Quicklime
11. Which one of the following is not an element?
(a) Diamond (b) Ozone (c) Silica (d) Graphite
12. The direct change from solid to gaseous state is referred to as:
(a) dissociation (b) decomposition
(c) sublimation (d) deliquescence
13. Sulphur burns in oxygen to form sulphur dioxide. The properties of sulphur dioxide are:
(a) totally different from sulphur and oxygen
(b) similar to sulphur
(c) similar to oxygen
(d) more similar to sulphur than oxygen
14. Which one of the following statements is incorrect?
(a) All elements are homogeneous
(b) Compounds always contain two or more different elements
(c) A mixture is not always heterogeneous
(d) Air is a heterogeneous mixture
15. One sample of air is found to have 0.03% carbon dioxide and another sample 0.02%. This illustrates that:
(a) air is a compound
(b) air is an element
(c) air does not follow the law of constant proportions
(d) air is a mixture
16. Which one of the following is not a mixture?
(a) Distilled water
(b) Sugar dissolved in water
(c) Liquefied Petroleum Gas (LPG)
(d) Gasoline
17. Which of the following is a characteristic property of both mixtures and compounds?
(a) Their properties are same as those of their components
(b) Energy is released when they are formed
(c) Their masses are equal to the sum of the masses of their components
(d) They contain the components in fixed proportions
18. Name the scientist who stated that matter can be converted into energy:
(a) Boyle (b) Lavoisier (c) Avogadro (d) Einstein
19. Which one of the following is not an intensive property?
(a) Weight (b) Density
(c) Refractive index (d) Melting point
20. The metalloid among the following group of elements is: (CPMT 1993)
(a) P (b) As (c) Al (d) Po
21. Which of the following alloys does contain Cu and Zn? (IIT 1993)
(a) Bronze (b) Brass
(c) Type metal (d) Rolled gold
22. Which metal is present in german silver?
(a) Copper (b) Iron (c) Silver (d) Zinc
23. Which of the following processes results in the formation of a new compound?
(a) Dissolving common salt in water
(b) Heating water
(c) Heating platinum rod
(d) Heating iron rod
24. Which one of the following is not a chemical change?
(a) Sublimation (b) Combustion
(c) Electrolysis (d) Rusting
25. The law of multiple proportions is illustrated by the pair of compounds:
(a) sodium chloride and sodium bromide
(b) water and heavy water
(c) sulphur dioxide and sulphur trioxide
(d) magnesium hydroxide and magnesium oxide
26. In compound A, 1.0 g nitrogen combines with 0.57 g oxygen. In compound B, 2.0 g nitrogen unite with 2.24 g oxygen and in compound C, 3.0 g nitrogen combine with 5.11 g oxygen. These results obey the law of:

(a) multiple proportions (b) constant proportions
(c) reciprocal proportions (d) none of these

27. Which one is the best example of law of conservation of mass?
(a) 6 g of carbon is heated in vacuum, there is no change in mass
(b) 6 g of carbon combines with 16 g of oxygen to form 22 g of CO_2
(c) 6 g water is completely converted into steam
(d) A sample of air is heated at constant pressure when its volume increases but there is no change in mass

28. A chemical equation is balanced according to the law of:
(a) multiple proportions (b) constant proportions
(c) reciprocal proportions (d) conservation of mass

29. SO_2 gas was prepared by (i) burning sulphur in oxygen, (ii) reacting sodium sulphite with dilute H_2SO_4 and (iii) heating copper with conc. H_2SO_4. It was found that in each case sulphur and oxygen combined in the ratio of 1 : 1. The data illustrates the law of:
(a) conservation of mass (b) multiple proportions
(c) constant proportions (d) reciprocal proportions

30. A sample of $CaCO_3$ has Ca = 40%, C = 12% and O = 48%. If the law of constant proportions is true, then the mass of Ca in 5 g of $CaCO_3$ from another source will be:
(a) 2.0 g (b) 0.2 g (c) 0.02 g (d) 20.0 g

31. Potassium combines with two isotopes of chlorine (^{35}Cl and ^{37}Cl) respectively to form two samples of KCl. Their formation follows the law of:
(a) constant proportions (b) multiple proportions
(c) reciprocal proportions (d) none of these

32. Different proportions of oxygen in the various oxides of nitrogen, prove the law of:
(a) reciprocal proportions (b) multiple proportions
(c) constant proportions (d) conservation of mass

33. One part of an element *A* combines with two parts of *B* (another element). Six parts of element *C* combine with four parts of element *B*. If *A* and *C* combine together, the ratio of their masses will be governed by:
(a) law of definite proportions
(b) law of multiple proportions
(c) law of reciprocal proportions
(d) law of conservation of mass

34. H_2S contains 5.88% hydrogen, H_2O contains 11.11% hydrogen while SO_2 contains 50% sulphur. These figures illustrate the law of:
(a) conservation of mass (b) constant proportions
(c) multiple proportions (d) reciprocal proportions

35. Number of atoms in 4.25 g of NH_3 is: **(AFMC 2010)**
(a) 6.023×10^{23} (b) $4 \times 6.023 \times 10^{23}$
(c) 1.7×10^{24} (d) $4.5 \times 6.023 \times 10^{23}$

[**Hint:** Number of molecules of $NH_3 = \frac{w}{M} \times 6.023 \times 10^{23}$

$$= \frac{4.25}{17} \times 6.023 \times 10^{23}$$

$$\text{Number of atom} = 4 \times \frac{4.25}{17} \times 6.023 \times 10^{23}$$

$$= 6.023 \times 10^{23}$$]

36. Hydrogen combines with chlorine to form HCl. It also combines with sodium to form NaH. If sodium and chlorine also combine with each other, they will do so in the ratio of their masses as:
(a) 23 : 35.5 (b) 35.5 : 23
(c) 1 : 1 (d) 23 : 1

37. Zinc sulphate contains 22.65% Zn and 43.9% H_2O. If the law of constant proportions is true, then the mass of zinc required to give 40 g crystals will be:
(a) 90.6 g (b) 9.06 g (c) 0.906 g (d) 906 g

38. 3 g of a hydrocarbon on combustion in excess of oxygen produces 8.8 g of CO_2 and 5.4 g of H_2O. The data illustrates the law of:
(a) conservation of mass (b) multiple proportions
(c) constant proportions (d) reciprocal proportions

[**Hint:** Mass of carbon in 8.8 g $CO_2 = \frac{12}{44} \times 8.8 = 2.4$g;

Mass of hydrogen in 5.4 g $H_2O = \frac{2}{18} \times 5.4 = 0.6$ g

Total mass of (C + H) = 2.4 + 0.6 = 3.0 g]

39. In the reaction, $N_2 + 3H_2 \longrightarrow 2NH_3$, the ratio of volumes of nitrogen, hydrogen and ammonia is 1 : 3 : 2. These figures illustrate the law of:
(a) constant proportions (b) Gay-Lussac
(c) multiple proportions (d) reciprocal proportions

40. Two volumes of ammonia, on dissociation gave one volume of nitrogen and three volumes of hydrogen. How much hydrogen will be obtained from the dissociation of 10 litre of NH_3?
(a) 30 litre (b) 10 litre (c) 15 litre (d) 20 litre

41. If 6 litre of H_2 and 5.6 litre of Cl_2 are mixed and exploded in an eudiometer, the volume of HCl formed is:
(a) 6.0 litre (b) 5.6 litre (c) 11.2 litre (d) 11.6 litre

42. The law of constant proportions was enunciated by:
(a) Dalton (b) Berthelot (c) Avogadro (d) Proust

43. An important postulate of Dalton's atomic theory is:
(a) an atom contains electrons, protons and neutrons
(b) atom can neither be created nor destroyed nor divisible
(c) all the atoms of an element are not identical
(d) all the elements are available in nature in the form of atoms

44. The atomic masses of the elements are usually fractional because:
(a) elements consist of impurities
(b) these are mixtures of allotropes
(c) these are mixtures of isobars
(d) these are mixtures of isotopes

45. The chemical formula of a particular compound represents:
(a) the size of its molecule
(b) the shape of its molecule
(c) the total number of atoms in a molecule
(d) the number of different types of atoms in a molecule

46. Which one of the following properties of an element is not variable?
(a) Valency (b) Atomic mass
(c) Equivalent mass (d) All of these

47. 1 amu is equal to:
(a) 1.00758 g (b) 0.000549 g
(c) 1.66×10^{-24} g (d) 6.02×10^{-23} g

48. Which one of the following relationships is correct?
(a) At. mass = 6.4 × Sp. heat.
(b) At. mass × Sp. heat = 6.4
(c) At. mass × 6.4 = Sp. heat
(d) At. mass × Sp. heat × 6.4 = 1

49. A_1 g of an element gives A_2 g of its oxide. The equivalent mass of the element is:
(a) $\frac{A_2 - A_1}{A_1} \times 8$ (b) $\frac{A_2 - A_1}{A_2} \times 8$
(c) $\frac{A_1}{A_2 - A_1} \times 8$ (d) $(A_2 - A_1) \times 8$

50. A_1 g of an element gives A_2 g of its chloride; the equivalent mass of the element is:
(a) $\frac{A_1}{A_2 - A_1} \times 35.5$ (b) $\frac{A_2}{A_2 - A_1} \times 35.5$
(c) $\frac{A_2 - A_1}{A_1} \times 35.5$ (d) $\frac{A_2 - A_1}{A_2} \times 35.5$

51. Which one of the relationship is wrong?
(a) 2 V.D. = Mol. mass
(b) At. mass = Eq. mass × Valency
(c) At. mass = $\frac{6.4}{\text{Sp. heat}}$
(d) Valency = $\frac{\text{Mol. mass}}{\text{Eq. mass}}$

52. In m_1 g of a metal A displaces m_2 g of another metal B from its salt solution and if their equivalent masses are E_1 and E_2 respectively, then the equivalent mass of A can be expressed as:
(a) $E_1 = \frac{m_2 \times E_2}{m_1}$ (b) $E_1 = \frac{m_1}{m_2} \times E_2$
(c) $E_1 = \frac{m_1 \times m_2}{E_2}$ (d) $E_1 = \sqrt{\frac{m_1}{m_2} \times E_2}$

53. When the specific heat of a metallic element is 0.214 cal g^{-1}, the atomic mass will be closest to which one of the following?
(a) 1 (b) 12 (c) 30 (d) 66

54. Approximate atomic mass of an element is 26.89. If its equivalent mass is 8.9, the exact atomic mass of the element is:
(a) 26.89 (b) 8.9 (c) 17.8 (d) 26.7

55. When an element forms an oxide in which oxygen is 20% of the oxide by mass, the equivalent mass of the element will be:
(a) 32 (b) 40 (c) 60 (d) 128

56. 0.32 g of a metal gave on treatment with an acid 112 mL of hydrogen at NTP. Calculate equivalent mass of the metal: **[AMU (Engg.) 2010]**
(a) 58 (b 32 (c) 11.2 (d) 24
[**Hint:** Mass of metal that displaces 11200 mL hydrogen at STP will be its equivalent mass.
∴ Equivalent mass of metal = $\frac{0.32}{112} \times 11200 = 32$]

57. 74.5 g of a metallic chloride contains 35.5 g of chlorine. The equivalent mass of the metal is:
(a) 19.5 (b) 35.5 (c) 39.0 (d) 78.0

58. The product of atomic mass and specific heat of any element is a constant, approximately 6.4. This is known as:
(a) Dalton's law (b) Avogadro's law
(c) Gay-Lussac law (d) Dulong Petit's law

59. The molecular mass of chloride, MCl, is 74.5. The equivalent mass of the metal M will be:
(a) 39.0 (b) 74.5
(c) 110.0 (d) 35.5
[**Hint:** Mol. mass = At. mass + 35.5 = Eq. mass × valency + 35.5, Valency of M from the formula MCl is 1]

60. 1 g of hydrogen is found to combine with 80 g of bromine. 1 g of calcium combines with 4 g of bromine. The equivalent mass of calcium is:
(a) 10 (b) 20 (c) 40 (d) 80

61. 2.8 g of iron displaces 3.2 g of copper from a solution of copper sulphate. If the equivalent mass of iron is 28, the equivalent mass of copper will be:
(a) 16 (b) 32 (c) 48 (d) 64

62. The specific heat of a metal of atomic mass 32 is likely to be:
(a) 0.25 (b) 0.24 (c) 0.20 (d) 0.15

63. The equivalent mass of an element is 4. Its chloride has a vapour density 59.25. The valency of the element will be:
(a) 4 (b) 3 (c) 2 (d) 1

64. The equivalent mass of iron in the reaction,
$3Fe + 4H_2O = Fe_3O_4 + 4H_2$ would be:
(a) 21 (b) 56 (c) 42 (d) 10

65. The specific heat of a bivalent metal is 0.16. The approximate equivalent mass of the metal will be:
(a) 40 (b) 20 (c) 80 (d) 10

66. A sample of pure calcium weighing 1.35 g was quantitatively converted to 1.88 g of pure calcium oxide. Atomic mass of calcium would be:
(a) 20 (b) 40 (c) 16 (d) 35.5

67. A metal oxide is reduced by heating it in a stream of hydrogen. It is found that after complete reduction, 3.15 g of the oxide have yielded 1.05 g of the metal. We may conclude that: **(MLNR 1991)**
(a) atomic mass of the metal is 4
(b) atomic mass of the metal is 8
(c) equivalent mass of the metal is 4
(d) equivalent mass of the metal is 8

68. Compounds with identical crystal structure and analogous chemical formula are called:
(a) isomers (b) isotones
(c) allotropes (d) isomorphous

69. Which pair of the following substances is said to be isomorphous?
(a) White vitriol and blue vitriol
(b) Epsom salt and Glauber's salt
(c) Blue vitriol and Glauber's salt
(d) White vitriol and epsom salt

c mass of chlorine is 35.5. It has two isotopes of atomic
5 and 37. The percentage of heavier isotope is:
(b) 15 (c) 20 (d) 25
c mass of boron is 10.81. It has two isotopes with 80%
0% abundance respectively. The atomic mass of the
e having 80% abundance is 11.01. The atomic mass of
ier isotope is:
.81 (b) 11.01 (c) 10.01 (d) 21.82
of chlorine combines with a metal giving 111 g of its
le. The chloride is isomorphous with $MgCl_2 \cdot 6H_2O$. The
mass of the metal is:
(b) 30 (c) 40 (d) 69
pour density of a volatile chloride of a metal is 59.5 and
uivalent mass of the metal is 24. The atomic mass of the
nt will be:
(b) 48 (c) 24 (d) 12
xide of an element possesses the molecular formula,
. If the equivalent mass of the metal is 9, the atomic
of the metal will be:
(b) 18 (c) 9 (d) 4.5
olecular mass of a compound having formula MO and
lent mass 20 is:
(b) 40 (c) 28 (d) 20
ensity of air is 0.001293 g mL^{-1}. Its vapour density is:
3 (b) 14.3 (c) 1.43 (d) 0.143
Divide with the density of hydrogen, *i.e.*, 0.00009 g

of a substance when vaporised occupy a volume of 5.6
NTP. The molecular mass of the substance will be:
(b) $2M$ (c) $3M$ (d) $4M$
apour densities of two gases are in the ratio of 1 : 3.
molecular masses are in the ratio of:
3 (b) 1 : 2 (c) 2 : 3 (d) 3 : 1
ganic compound on analysis was found to contain
% of sulphur. The molecular mass of the compound, if its
ule contains two sulphur atoms, is:
0 (b) 2000
000 (d) 200000
tomic mass of an element is 27. If valency is 3, the
r density of the volatile chloride will be:
.75 (b) 6.675 (c) 667.5 (d) 81
ensity of a gas 'A' is three times that of a gas 'B'. If the
ular mass of A is M, the molecular mass of B is:
(b) $M/3$ (c) $\sqrt{3}M$ (d) $\frac{M}{\sqrt{3}}$
r density of a volatile substance is 4 in comparison to
ne ($CH_4 = 1$). Its molecular mass will be:
(b) 2 (c) 64 (d) 128
e the wrong statement:
mole means 6.02×10^{23} particles
olar mass is mass of one molecule
olar mass is mass of one mole of a substance
olar mass is molecular mass expressed in grams
le of CO_2 contains: **(MLNR 1990; CBSE 1993)**
02×10^{23} atoms of C

(c) 18.1×10^{23} molecules of CO_2
(d) 3 g-atoms of CO_2

85. Which among the following is the heaviest? **[PMT (Kerala) 2006]**
(a) One mole of oxygen
(b) One molecule of sulphur trioxide
(c) 100 amu of uranium
(d) 10 moles of hydrogen
(e) 44 g of carbon dioxide

86. The largest number of molecules is in:
(a) 28 g of CO (b) 46 g of C_2H_5OH
(c) 36 g of H_2O (d) 54 g of N_2O_5

87. Which of the following has the smallest number of molecules?
(a) 22.4×10^3 mL of CO_2 gas
(b) 22 g of CO_2 gas
(c) 11.2 litre of CO_2 gas
(d) 0.1 mole of CO_2 gas

88. The number of grams of H_2SO_4 present in 0.25 mole of H_2SO_4 is:
(a) 0.245 (b) 2.45 (c) 24.5 (d) 49.0

89. Number of molecules in 1 litre of oxygen at NTP is:
(a) $\frac{6.02 \times 10^{23}}{32}$ (b) $\frac{6.02 \times 10^{23}}{22.4}$
(c) 32×22.4 (d) $\frac{32}{22.4}$

90. 4.6×10^{22} atoms of an element weigh 13.8 g. The atomic mass of the element is:
(a) 290 (b) 180 (c) 34.4 (d) 10.4

91. The number of molecules in 89.6 litre of a gas at NTP are: **(BHU 1992)**
(a) 6.02×10^{23} (b) $2 \times 6.02 \times 10^{23}$
(c) $3 \times 6.02 \times 10^{23}$ (d) $4 \times 6.02 \times 10^{23}$

92. The total number of protons in 10 g of calcium carbonate is: **(CPMT 1992)**
(a) 3.0115×10^{24} (b) 1.5057×10^{24}
(c) 2.0478×10^{24} (d) 4.0956×10^{24}

93. 19.7 kg of gold was recovered from a smuggler. The atoms of gold recovered are: (Au = 197)
(a) 100 (b) 6.02×10^{23}
(c) 6.02×10^{24} (d) 6.02×10^{25}

94. The molecular mass of CO_2 is 44 amu and Avogadro's number is 6.02×10^{23}. Therefore, the mass of one molecule of CO_2 is:
(a) 7.31×10^{-23} (b) 3.65×10^{-23}
(c) 1.01×10^{-23} (d) 2.01×10^{-23}

95. Equal volumes of different gases at any definite temperature and pressure have:
(a) equal weights (b) equal masses
(c) equal densities (d) equal number of moles

96. A gaseous mixture contains oxygen and nitrogen in the ratio of 1 : 4 by mass. Therefore, the ratio of their number of molecules is:

m will react with ethyl alcohol to give:
of hydrogen (b) one mole of oxygen
of NaOH (d) 1/2 mole of hydrogen
an atom of carbon is:
(b) 1.99×10^{-23} g
(d) 1.99×10^{23} g
of moles of H_2 in 0.224 litre of hydrogen gas at
(MLNR 1994)
(b) 0.1 (c) 0.01 (d) 0.001
of 1 g helium at NTP in litres is:
(b) 0.56 (c) 2.8 (d) 0.28
s to contain 1.0×10^{24} particles, the mass of one
gen is:
(b) 5.32 g
(d) 16.0 g
of 1 litre capacity each are separately filled with
He, O_2 and O_3. At the same temperature and
ratio of the number of atoms of these gases
fferent flasks would be:
1 (b) 2 : 1 : 2 : 3
3 (d) 3 : 2 : 2 : 1
mperature and pressure, two flasks of equal
filled with H_2 and SO_2 separately. Particles which
number, in the two flaskes are:
(b) electrons
es (d) neutrons
a mixture of 6.02×10^{23} oxygen atoms and
ydrogen molecules at NTP is:
(b) 33.6 litre
(d) 22.4 litre
of molecules present in a drop of water, if its
05 mL, are:
^{21} (b) 1.66×10^{22}
^{23} (d) 1.66×10^{24}
was found to contain nitrogen and oxygen in the
n 28 g and oxygen 80 g. The formula of the
:
(b) N_2O_3
(d) N_2O_4
t formula of a compound containing 50% of
At. mass = 10) and 50% of the element Y (At.
s: **(IIT 1994)**
(b) X_2Y
(d) X_2Y_3
ts X (At. mass = 75) and Y (At. mass = 16)
ive a compound having 75.8% of X. The formula
und is: **(MLNR 1991)**
(b) X_2Y
(d) X_2Y_3
a certain compound was found to contain iodine
in the ratio of 254 : 80. The formula of the
: (At. mass of I = 127, O = 16)
(b) I_2O
(d) I_2O_5

110. A compound of aluminium and chlorine is composed of 9.0 g Al for every 35.5 g of chlorine. The empirical formula of the compound is:
(a) $AlCl$ (b) $AlCl_3$ (c) $AlCl_2$ (d) $AlCl_4$

111. The haemoglobin from red corpuscles of most mammels contain approximately 0.33% of iron by mass. The molecular mass of haemoglobin is 67200. The number of iron atoms in each molecule of haemoglobin is:
(a) 4 (b) 3 (c) 2 (d) 1

112. The percentage of P_2O_5 in diammonium hydrogen phosphate, $[(NH_4)_2HPO_4]$ is: **(CPMT 1992)**
(a) 23.48 (b) 46.96
(c) 53.78 (d) 71.00

113. The percentage of nitrogen in urea (NH_2CONH_2), is:
(a) 38.4 (b) 46.6 (c) 59.1 (d) 61.3

114. The chloride of a metal has the formula MCl_3. The formula of its phosphate is:
(a) M_2PO_4 (b) MPO_4 (c) M_3PO_4 (d) $M(PO_4)_2$

115. 10 g of hydrofluoric acid gas occupies 5.6 litre of volume at NTP. If the empirical formula of the gas is HF, then its molecular formula will be: (At. mass of F = 19)
(a) HF (b) H_3F_3
(c) H_2F_2 (d) H_4F_4
[**Hint:** Molecular mass $= \frac{10}{5.6} \times 22.4 = 40$]

116. Calcium pyrophosphate is represented by the formula $Ca_2P_2O_7$. The molecular formula of ferric pyrophosphate is:
(a) $Fe_2P_2O_7$ (b) FeP_2O_7
(c) $Fe(P_2O_7)_3$ (d) $Fe_4(P_2O_7)_3$

117. The percentage of available chlorine in a sample of bleaching powder, $CaOCl_2 \cdot 2H_2O$, is:
(a) 30 (b) 50 (c) 43.5 (d) 59.9

118. SI unit of energy is:
(a) kg m^2 s^{-2} (b) kg m^{-1} s^2
(c) kg m^2 s^{-1} (d) kg m^2 s^2

119. One micro gram is equal to:
(a) 10^{-3} g (b) 10^3 g (c) 10^6 g (d) 10^{-6} g

120. Significant figures in 0.00051 are:
(a) 5 (b) 3 (c) 2 (d) 4

121. The number of significant figures in 6.02×10^{23} is:
(a) 23 (b) 3 (c) 4 (d) 26

122. Express 0.006006 into scientific notation in three significant digits:
(a) 6.01×10^{-3} (b) 6.006×10^{-3}
(c) 6.00×10^{-3} (d) 6.0×10^{-3}

123. The proper value of significant figures in 38.0 + 0.0035 + 0.00003 is:
(a) 38 (b) 38.0035
(c) 38.00353 (d) 38.0

124. Which of the following is the correct unit for measuring nuclear radii?
(a) Micron (b) Millimetre

er to prepare 1 litre normal solution of $KMnO_4$, how grams of $KMnO_4$ are required if the solution is to be acid medium for oxidation? **[PET (MP) 2002]**

8 g (b) 31.6 g (c) 62 g (d) 790 g

an oxide of a metal is converted to chloride completely yields 5 g chloride. The equivalent weight of metal is: **(KCET 2002)**

.25 (b) 3.325 (c) 12 (d) 20

er of atoms 558.5 g Fe (At. wt. of Fe = 55.85 g mol^{-1}) is: **(AIEEE 2002)**

ice that in 60 g carbon (b) 6.023×10^{23}

lf that of 8 g He (d) $558.5 \times 6.023 \times 10^{23}$

efix 10^{18} is: **[MEE (Kerala) 2002]**

ga (b) exa (c) kilo (d) nano

ega

ence in density is the basis of: **[MEE (Kerala) 2002]**

ra filtration (b) molecular sieving

avity separation (d) molecular attraction

mic absorption

ective of the source, pure sample of water always yields % mass of oxygen and 11.11% mass of h... This is **(Kerala) 2002]**

ned by the law of: ...tant composition

nservation of mass (b) constant volume

ltiple proportion

y-Lussac's l... of electron weigh one kilogram?

many ... **[IIT (Screening) 2002]**

3×10^{23} (b) $\frac{1}{9.108} \times 10^{31}$

$\frac{3 \times 10^{54}}{9.108}$ (d) $\frac{1}{9.108 \times 6.023} \times 10^{8}$

he numbers: 161 cm; 0.161 cm; 0.0161 cm. The of significant figure for three numbers is: **[AFMC (Pune) 2002]**

, 5 (b) 3, 3, 3 (c) 3, 3, 4 (d) 3, 4, 4

ne of the following laws directly explains the law of ation of mass? **(AFMC 2002)**

d's rule (b) Dalton's law

gadro's law (d) Berzelius hypothesis

as maximum number of atoms? **[IIT (Screening) 2003]**

C(12) (b) 56 g Fe(56)

Al(27) (d) 108 g Ag(108)

ich of sulphur is present in an organic compound, if ompound gave 1.158 g of $BaSO_4$ on analysis? **[PET (Kerala) 2005]**

(b) 15% (c) 20% (d) 25%

of H_2 and 20 mL of O_2 react to form water, what is e end of the reaction? **(AFMC 2005)**

137. The density of a liquid is 1.2 g/mL. There are 35 drops in 2 mL. The number of molecules in one drop (molar mass of liquid = 70) is:

(a) $\left(\frac{1.2}{35}\right) N_A$ (b) $\left(\frac{1}{35}\right)^2 N_A$

(c) $\frac{1.2}{(35)^2} N_A$ (d) $1.2 N_A$

138. A sample of PCl_3 contains 1.4 mole of the substance. How many atoms are there in the sample? **[CEE (Kerala) 2004]**

(a) 4 (b) 5.6

(c) 8.431×10^{23} (d) 3.372×10^{24}

(e) 2.409×10^{24}

139. The equivalent weight of phosphoric acid H_3PO_4 in the reaction, $NaOH + H_3PO_4 \longrightarrow NaH_2PO_4 + H_2O$ is: **[BHU (Pre.) 2005]**

(a) 59 (b) 49 (c) 25 (d) 98

[Hint: Only one hydrogen of H_3PO_4 is replaced, *i.e.*, its basicity = 1

$$\text{Equivalent mass} = \frac{\text{Molecular mass}}{\text{Basicity}} = \frac{98}{1} = 98]$$

140. 5.6 g of an organic compound on burning with excess of oxygen gave 17.6 g of CO_2 and 7.2 g H_2O. The organic compound is: **[PET (Kerala) 2006]**

(a) C_6H_6 (b) C_4H_8 (c) C_3H_8 (d) CH_3COOH

(e) CH_3CHO

141. The decomposition of a certain mass of $CaCO_3$ gave 11.2 dm^3 of CO_2 gas at STP. The mass of KOH required to completely neutralise the gas is: **(KCET 2006)**

(a) 56 g (b) 28 g (c) 42 g (d) 20 g

[**Hint:** 11.2 dm^3 of CO_2 at STP = $\frac{1}{2}$ mole CO_2

$$KOH + CO_2 \longrightarrow KHCO_3$$

$\frac{1}{2}$ mole CO_2 will be neutralised by $\frac{1}{2}$ mole KOH, *i.e.*, 28 g KOH.]

142. How many moles of magnesium phosphate, $Mg_3(PO_4)_2$, will contain 0.25 mole of oxygen atoms? **(AIEEE 2006)**

(a) 0.02 (b) 3.125×10^{-2}

(c) 1.25×10^{-2} (d) 2.5×10^{-2}

[**Hint:**

$\because$ 8 mole oxygen atoms are present in 1 mole $Mg_3(PO_4)_2$

$\therefore$ 0.25 mole oxygen atoms will be present in $\frac{1}{8} \times 0.25$ mole

$Mg_3(PO_4)_2$, *i.e.*, 3.125×10^{-2} mole $Mg_3(PO_4)_2$]

143. An element, X has the following isotopic composition,

^{200}X : 90%

^{199}X : 8%

^{202}X : 2%

the weighted average atomic mass of the naturally- occurring element 'X' is closest to : **[CBSE (Med.) 2007]**

(a) 201 amu (b) 202 amu (c) 199 amu (d) 200 amu age atomic mass of

$$\times 200\Big] + \Big[\frac{8}{100}\times 199\Big] + \Big[\frac{2}{100}\times 202\Big]$$

5 amu $\approx$ 200 amu]

statement for 14 g CO is : **(VMMC 2007)**

s 2.24 L at NTP

po ds to $\frac{1}{2}$ mole of CO

pond to same mole of CO and N_2

ponds 3.01×10^{23} molecules of CO

r of hydrogen atoms present in 25.6 g of sucrose

) which ha molar mass of 342.3 g, is:

(VITEEE 2008)

3 (b) 9.91×10^{23}

3 (d) 44 ... 23

nber of moles of sucrose

$$= \frac{\text{Mass}}{\text{Molar mass}} = \frac{2...}{342.3}$$

$= 0.075$

noles of hydrogen atom $= 0.075\times 22$

toms of hydrogen $= 0.075\times 22\times 6.023\times 10^{23}$

$= 9.9\times 10^{23}$]

upied by one molecule of water (density = 1 g/cm^3)

[CBSE-PMT (Pre.) 2008]

^{3} cm^3 (b) 5.5×10^{-23} cm^3

^{3} cm^3 (d) 6.023×10^{-23} cm^3

ass of one molecule $= \dfrac{18}{6.023\times 10^{23}}$ g

$= 2.98\times 10^{-23}$ g

ne molecule $= \dfrac{M}{\text{Density}} = \dfrac{2.98\times 10^{-23}}{1}$ cm^3

$\approx 3\times 10^{-23}$ cm^3]

gen contains as many atoms as in: **(KCET 2008)**

hydrogen (b) 5 g of hydrogen

hydrogen (d) 1 g of hydrogen

s consists of uniform cylindrical particles of 150Å and 5000Å long. The specific volume of virus is If the virus is considered to be a single particle, its ass is:

)^7 g mol^{-1} (b) 7.90×10^7 g mol^{-1}

)^7 g mol^{-1} (d) 9.70×10^7 g mol^{-1}

ume of single virus $= \pi r^2 h$

$= 3.14\times (75\times 10^{-8})^2\times (5000\times 10^{-8})$

$= 8.836\times 10^{-17}$ cm^3

ingle virus $= \dfrac{\text{Volume}}{\text{Specific volume}} = \dfrac{8.836\times 10^{-17}\text{ cm}^3}{0.75\text{ cm}^3/\text{g}}$

$= 1.178\times 10^{-16}$ g

ss of virus $= 1.178\times 10^{-16}\times 6.023\times 10^{23}$

$= 7.09\times 10^7$ g mol^{-1}]

149. Common salt obtained from sea-water contains 95% NaCl by mass. The approximate number of molecules present in 10 g salt is: **(DPMT 2009)**

(a) 10^{21} (b) 10^{22}

(c) 10^{23} (d) 10^{24}

[**Hint :** Mass of NaCl in 10 g salt $= 10\times\dfrac{95}{100} = 9.5$ g

Number of molecules of NaCl $= \dfrac{9.5}{58.5}\times 6.023\times 10^{23}$

$= 9.78\times 10^{22}\approx 10^{23}$]

150. 10 g hydrogen and 64 g oxygen were filled in a steel vessel and exploded. Amount of water produced in this reaction will be: **[CBSE (PMT) 2009]**

(a) 3 mol (b) 4 mol

(c) 1 mol (d) 2 mol

[**Hint :** $n_{H_2} = \dfrac{10}{2} = 5$ $\quad n_{O_2} = \dfrac{64}{32} = 2$

$$2H_2 + O_2 \longrightarrow 2H_2O$$

Case I : If H_2 is completely consumed then :

$n_{H_2O} = 5$ mol

Case II : If O_2 is completely consumed then

$n_{H_2O} = \dfrac{2}{1}\times 2 = 4$ mol

Since, O_2 gives le... amount of product on complete consumption hence, it is limiting ... number of moles of water formed will be equal to 4.]

151. An organic compound made of C, H and N contains 20% nitrogen. What will be its molecular mass if it contains only one nitrogen atom in it? **[J... (...B) 2009]**

(a) 70 (b) 140

(c) 100 (d) 65

[**Hint :** $\% N = \dfrac{\text{Mass of nitrogen}}{\text{Molecular mass}}\times 100$

$20 = \dfrac{14}{m}\times 100$

$m = 70$]

152. 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: **(IIT 2009)**

(a) 55.85 u (b) 55.95 u

(c) 55.75 u (d) 56.05 u

[**Hint :** Atomic mass of Fe $= \dfrac{5}{100}\times 54 + \dfrac{90}{100}\times 56 + \dfrac{5}{100}\times 57$

$= 55.95$ amu]

153. The number of atoms in 0.1 mol triatomic gas is: ($N_A = 6.02\times 10^{23}$ mol^{-1}) **[CBSE (PMT) 2010]**

(a) 6.026×10^{22} (b) 1.806×10^{23}

(c) 3.6×10^{23} (d) 1.8×10^{22}

[**Hint :** No. of atoms $= 0.1\times 3\times 6.02\times 10^{23}$

$= 1.806\times 10^{23}$]

questions given below may have more than one correct answers

the following relationships are wrong?
$\approx$ 0.1 bar (b) 1 litre = 1 dm^3
0.239 cal (d) 1 eV = 9.11 $\times 10^{-4}$ J

of the following numbers have same significant

) (b) 0.60 (c) 6.0 (d) 60

f the following have the same mass?
nole of O_2 gas
nole of SO_2 gas
3 $\times 10^{22}$ molecules of SO_2 gas
4 $\times 10^{23}$ molecules of O_2 gas

4. Select the numbers with same significant figures:
(a) 6.02×10^{23} (b) 0.25
(c) 6.60×10^{-34} (d) 1.50
5. Which are isomorphic to each other?
(a) $CuSO_4 \cdot 5H_2O$ (b) $ZnSO_4 \cdot 7H_2O$
(c) $FeSO_4 \cdot 7H_2O$ (d) $FeSO_4 \cdot 8H_2O$
6. 11.2 L of a gas at STP weighs 14 g. The gas could be:
(a) N_2 (b) CO
(c) NO_2 (d) N_2O
7. 8 g O_2 has same number of molecules as that in:
(a) 14 g CO (b) 7 g CO
(c) 11 g CO_2 (d) 22 g CO_2

Assertion-Reason TYPE QUESTIONS

e following questions, two statements are given as A) and **'Reason'** (R). Answer the questions by er codes given below:
oth (A) and (R) are correct and (R) is the correct anation of (A).
oth (A) and (R) are correct but (R) is not the correct anation of (A).
) is correct and (R) is wrong.
) is wrong but (R) is correct.
oth (A) and (R) are wrong.

O_2 and 1 g O_3 have equal number of atoms.
ss of 1 mole atom is equal to its gram-atomic mass.
our density of sulphur vapour relative to oxygen is 2 ause sulphur atom is twice as heavy as that of oxygen n.
our density depends upon the molecular state of the stance in vapour state.
vogram is equal to 1 amu.
ogram is reciprocal of Avogadro's number.
ole H_2SO_4 contains same mass of oxygen and sulphur.
ole H_2SO_4 represents 98 g mass.
nole oxygen and N_2 have same volume at same perature and pressure.
ole gas at NTP occupies 22.4 litre volume at STP.

6. (A) Empirical formula of glucose is HCHO.
(R) Molecular formula of glucose will also be equal to HCHO.
7. (A) The volume of 1 mole of an ideal gas at 1 bar pressure at 25°C is 24.78 litre.
(R) 1 bar = 0.987 atm.
8. (A) Atomic weight = Specific heat (cal/mol) $\times$ 6.4
(R) The formula is valid for metals only.
9. (A) Number of moles of H_2 in 0.224 L of H_2 is 0.01 mol.
(R) 22.4 litres of H_2 at STP contains 6.023×10^{23} mol. **(AIIMS 1996)**
10. (A) The equivalent weight of an element is variable.
(R) The valency of an element is variable. **(AIIMS 1995)**
11. (A) The number of significant figures in 507000 is three.
(R) In 507000, all the zeros are significant.
12. (A) Law of conservation of mass is invalid for nuclear fission, fusion and disintegration.
(R) The law proposes that mass is neither created nor destroyed in a reaction.
13. (A) Mass spectrometer is used for determination of atomic mass of isotopes.
(R) Isotopes are the atoms of same element having same atomic number but different mass numbers.

E QUESTIONS

3. (a)	4. (d)	5. (a)	6. (c)	7. (d)	8. (b)
1. (c)	12. (c)	13. (a)	14. (d)	15. (d)	16. (a)
9. (a)	20. (b)	21. (b)	22. (a)	23. (d)	24. (a)
7. (b)	28. (d)	29. (c)	30. (a)	31. (d)	32. (b)
5. (a)	36. (a)	37. (b)	38. (a)	39. (b)	40. (c)
3. (b)	44. (d)	45. (d)	46. (b)	47. (c)	48. (b)
1. (d)	52. (b)	53. (c)	54. (d)	55. (a)	56. (b)
9. (a)	60. (b)	61. (b)	62. (c)	63. (b)	64. (a)
7. (c)	68. (d)	69. (d)	70. (d)	71. (c)	72. (c)
5. (a)	76. (b)	77. (d)	78. (a)	79. (d)	80. (a)
3. (b)	84. (a)	85. (e)	86. (c)	87. (d)	88. (c)
1. (d)	92. (a)	93. (d)	94. (a)	95. (d)	96. (c)
9. (c)	100. (a)	101. (a)	102. (b)	103. (c)	104. (d)
7. (b)	108. (d)	109. (d)	110. (b)	111. (a)	112. (c)
5. (c)	116. (d)	117. (c)	118. (a)	119. (d)	120. (c)
3. (d)	124. (d)	125. (b)	126. (a)	127. (a)	128. (b)
1. (d)	132. (b)	133. (c)	134. (a)	135. (e)	136. (d)
. (d)	140. (b)	141. (b)	142. (b)	143. (d)	144. (a)
7. (b)	148. (a)	149. (c)	150. (b)	151. (a)	152. (b)

3. (b, c)	4. (a, c, d)	5. (b, c)	6. (a, b)	7. (b, c)

I-REASON TYPE QUESTIONS

(c)	4. (d)	5. (b)	6. (c)	7. (b)	8. (b)
(e)	12. (b)	13. (a)			

BRAIN STORMING PROBLEMS

OBJECTIVE QUESTIONS for IIT ASPIRANTS

ng questions contain single correct option:

he following table:

npound l. mass)	Mass of the compound (in grams) taken
CO_2 (44)	4.4
NO_2 (46)	2.3
H_2O_2 (34)	6.8
SO_2 (64)	1.6

two compounds have least mass of oxygen?
ular masses of compounds are given in brackets.) **(EAMCET 2004)**

nd IV (b) I and III (c) I and II (d) III and IV

I. Mass of oxygen present $= \frac{4.4}{44} \times 32 = 3.2$ g

II. Mass of oxygen present $= \frac{2.3}{46} \times 32 = 1.6$ g

II. Mass of oxygen present $= \frac{6.8}{34} \times 32 = 6.4$ g

V. Mass of oxygen present $= \frac{1.6}{64} \times 32 = 0.8$ g

l IV have least mass of oxygen.]

osphate of a certain metal M is $M_3(PO_4)_2$. The correct a of metal sulphate would be:

$(SO_4)_3$ (b) MSO_4

$(SO_4)_2$ (d) M_2SO_4

rcentage of Se in peroxidase enzyme is 0.5% by mass mass of Se = 78.4 amu). Then, the minimum lar mass of enzyme which contains not more than one n is:

68×10^4 amu (b) 1.568×10^7 amu

68×10^3 amu (d) 1.568×10^6 amu

$\because$ 0.5 g Se is present in 100 g enzyme.

Se will be present in $\frac{100}{0.5} \times 78.4$ g enzyme

$= 15680$ amu

$= 1.568 \times 10^4$ amu]

mber of moles of a gas in 1 m^3 of volume at NTP is:

(b) 0.446 (c) 1.464 (d) 44.6

$1\ m^3 = 1000$ L

nber of moles $= \frac{1000}{22.4} = 44.6$]

al number of electrons present in 18 mL water (density) is:

3×10^{23} (b) 6.023×10^{24}

3×10^{25} (d) 6.023×10^{21}

[**Hint:** Mass = 18 g

Number of molecules of H_2O in 18 g mass $= 6.023 \times 10^{23}$

Number of electrons in 18 g water $= 6.023 \times 10^{23} \times 10$

$= 6.023 \times 10^{24}$

$\because$ Each molecule of water contains 10 electrons.]

6. What is the empirical formula of vanadium oxide if 2.74 g of metal oxide contains 1.53 g of metal?

(a) V_2O_3 (b) VO (c) V_2O_5 (d) V_2O_7

[Hint: % of V $= \frac{1.53}{2.74} \times 100 = 55.83$

$\therefore$ % of O = 44.17

Element	%	Atomic ratio	Simplest ratio
V	55.83	$\frac{55.83}{52} = 1.1$	$\frac{1.1}{1.1} = 1$
O	44.17	$\frac{44.17}{16} = 2.76$	$\frac{2.76}{1.1} = 2.5$

V : O = 2 : 5

Thus, empirical formula = V_2O_5]

7. Number of moles of electrons in 4.2 g of N^{3-} ion (nitride ion) is:

(a) 3 (b) 2 (c) 1.5 (d) 4.2

8. The ratio of volumes occupied by 1 mole O_2 and 1 mole CO_2 under identical conditions of temperature and pressure is:

(a) 1 : 1 (b) 1 : 2 (c) 1 : 3 (d) 2 : 1

9. The maximum amount of $BaSO_4$ that can be obtained on mixing 0.5 mole $BaCl_2$ with 1 mole H_2SO_4 is:

(a) 0.5 mol (b) 0.1 mol (c) 0.15 mol (d) 0.2 mol

[**Hint:** $H_2SO_4 + BaCl_2 \longrightarrow BaSO_4 + 2HCl$

0.5 mole $BaCl_2$ will react with 0.5 mole H_2SO_4 to give 0.5 mole $BaSO_4$]

10. If 10^{21} molecules are removed from 100 mg CO_2, then number of moles of CO_2 left are:

(a) 6.10×10^{-4} (b) 2.8×10^{-3}

(c) 2.28×10^{-3} (d) 1.36×10^{-2}

[**Hint:** Number of molecules in 100 mg CO_2

$$= \frac{\text{Mass}}{\text{Molar mass}} \times 6.023 \times 10^{23}$$

$$= \frac{0.1}{44} \times 6.023 \times 10^{23}$$

$$= 1.368 \times 10^{21}$$

Molecules remaining $= 1.368 \times 10^{21} - 10^{21} = 0.368 \times 10^{21}$

Number of moles remaining $= \frac{0.368 \times 10^{21}}{6.023 \times 10^{23}} = 6.1 \times 10^{-4}$]

on one gram ion of Al^{3+} ion is:

: e coulomb (b) $\frac{1}{3} \times N_A \times e$ coulomb

e coulomb (d) $3 \times N_A \times e$ coulomb

: gram ion of Al^{3+} means one mole ion of Al^{3+}. mole $Al^{3+} = 3 \times e \times N_A$ coulomb.]

nass of N_2O as well as CO_2 is 44 g mol^{-1}. At 25°C essure, 1 L N_2O contains n molecules of gas. The CO_2 molecules in 2 L under same conditions will

(b) $2n$ (c) $\frac{n}{2}$ (d) $\frac{n}{4}$

is dissolved in 1 L water. The number of ions of $^-$ in 1 mL of this solution will be:

^{19} (b) 1.2×10^{22} (c) 1.2×10^{20} (d) 6.02×10^{20}

nber of moles of NaCl

$$= \frac{\text{Mass}}{\text{Molar mass}} = \frac{5.85}{58.5} = 0.1$$

ons ($Na^+ + Cl^-$) in 1 L

$$= 2 \times 0.1 \times 6.023 \times 10^{23}$$

$$= 12.046 \times 10^{22}$$

ons in 1 mL $= \frac{12.046 \times 10^{22}}{1000} = 1.2 \times 10^{20}$]

de has the formula X_2O_3. It can be reduced by give free metal and water. 0.1596 g of metal oxide g of hydrogen for complete reduction. The atomic al in amu is:

(b) 155.8 (c) 5.58 (d) 55.8

$$_2O_3 + 3H_2 \longrightarrow 2X + 3H_2O$$

nol 3 mol

48) g 6g

H_2 is required by 0.1596 g oxide

ill be required by 159.6 g oxide

$$2a + 48 = 159.6$$

$$a = 55.8$$

tomic mass of metal M.]

(PH_3) decomposes to produce vapours of and H_2 gas. What will be the change in volume L of phosphine is decomposed?

(b) 500 mL

(d) – 500 mL

$PH_3(g) \longrightarrow P_4(g) + 6H_2(g)$

4 mL $\longrightarrow$ 1 mL 6 mL

0 mL $\longrightarrow \frac{100}{4}$ $\frac{6}{4} \times 100$

0 mL $\longrightarrow$ 25 mL 150 mL

eases by 75 mL.]

neutron is assumed to half of its original value, t of proton is assumed to be twice of its original he atomic mass of $^{14}_{6}C$ will be:

[**Hint:** In the isotope $^{14}_{6}C$: Number of protons = 6

Number of neutrons = 8

New atomic mass will be $= 2 \times 6 + \frac{1}{2} \times 8 = 16$

% Increase in mass $= \frac{16-14}{14} \times 100 = 14.28\%$]

17. The mass and charge of 1 mole electrons will be:
(a) 1 kg; 96500 C (b) 0.55 mg; 96500 C
(c) 1.55 mg; 96500 C (d) 5.5 mg; 96500 C

18. The simplest formula of the compound containing 50% X (atomic mass 10 amu) and 50% Y (atomic mass 20 amu) is:
(a) XY_2 (b) X_2Y (c) X_2Y_3 (d) XY

[**Hint:**

Element	%	Atomic ratio	Simplest ratio
X	50	$\frac{50}{10} = 5$	$\frac{5}{2.5} = 2$
Y	50	$\frac{50}{20} = 2.5$	$\frac{2.5}{2.5} = 1$

Formula = X_2Y]

19. Rest mass of 1 mole neutrons ($m_n = 1.675 \times 10^{-27}$ kg) is:
(a) 1.8×10^{-3} kg (b) 1.008×10^{-4} kg
(c) 1.08×10^{-3} kg (d) 1.008×10^{-3} kg

[**Hint:** Mass of 1 mole neutrons

$$= 1.675 \times 10^{-27} \times 6.023 \times 10^{23}$$

$$= 1.008 \times 10^{-3} \text{ kg}]$$

20. Loschmidt number is the number of:
(a) molecules present in 1 mL of a gas at STP
(b) molecules present in 1 gram mole of a gas at STP
(c) atoms present in 1 mL of a gas at STP
(d) atoms present in 1 gram mole of a gas at STP

21. Which of the following statements is incorrect?
(a) One gram mole of silver equals $\frac{108}{6.023} \times 10^{-23}$ g
(b) One mole of CH_4 and 17 g of NH_3 at NTP occupy same volume
(c) One mole Ag weighs more than that of two moles of Ca
(d) One gram mole of CO_2 is 6.023×10^{23} times heavier than one molecule of CO_2

22. One atom of an element 'X' weighs 6.664×10^{-23} gm. The number of gram atoms in 40 kg of it is:
(a) 10 (b) 100 (c) 10000 (d) 1000

23. The density of a liquid is 1.2 g/mL. There are 35 drops in 2 mL. The number of molecules in 1 drop is (molecular weight of liquid = 70):
(a) $\frac{1.2}{35} N_A$ (b) $\left(\frac{1}{35}\right)^2 N_A$
(c) $\frac{1.2}{(35)^2} N_A$ (d) $1.2 N_A$

[**Hint:** Volume of one drop $= \frac{2}{35}$ mL

Number of moles in one drop $= \frac{2 \times 1.2}{35 \times 70} = \frac{1.2}{(35)^2}$

nolecules in one drop $= \frac{12}{(35)^2} \times N_A$]

ne of a liquid will contain 4 mole? Molar mass of 0 and its density is 1.4 g/mL:

(b) 1.6 L (c) 0.8 L (d) 4.8×10^{23} L

t x L liquid contain 4 mole of it.

ber of moles $= \frac{\text{Mass}}{\text{Molar mass}}$

$$4 = \frac{x \times 1000 \times 1.4}{280}$$

$$x = \frac{4 \times 280}{1.4 \times 1000} = 0.8 \text{ L}]$$

ratio of Fe^{2+} to Fe^{3+} in a mixture of $FeSO_4$ and having equal number of sulphate ions in both l ferric sulphates is:

(b) 3 : 2

(d) none of these

$SO_4 \longrightarrow Fe^{2+} + SO_4^{2-}$

$O_4)_3 \longrightarrow 2Fe^{3+} + 3SO_4^{2-}$

mole SO_4^{2-} ions are furnished by both $FeSO_4$ and

moles of $Fe^{2+} = x$

moles of $Fe^{3+} = \frac{2}{3}x$

$$Fe^{2+} : Fe^{3+} :: x : \frac{2}{3}x$$

$$= 3 : 2]$$

f electrons present in 3.6 mg of NH_4^+ are:

0^{21} (b) 1.2×10^{20} (c) 1.2×10^{22} (d) 2×10^{-3}

umber of electrons in one ion of $NH_4^+ = 10$

ions in 3.6 mg NH_4^+

$$= \frac{3.6 \times 10^{-3}}{18} \times 6.023 \times 10^{23} = 1.2 \times 10^{20}$$

ber of electrons in 3.6 mg $NH_4^+ = 1.2 \times 10^{20} \times 10$

$$= 1.2 \times 10^{21}]$$

ction $4A + 2B + 3C \longrightarrow A_4B_2C_3$, what will be the moles of product formed, starting from one mole of le of B and 0.72 mole of C?

(b) 0.3 (c) 0.24 (d) 2.32

$\underset{\text{mol}}{A} + \underset{0.6\text{ mol}}{2B} + \underset{0.72\text{ mol}}{3C} \longrightarrow A_4B_2C_3$

sent case, reactant 'C' will be the limiting reactant will give least amount of product on being completely

gives 1 mol product,

'C' will give 0.24 mol of product.]

f $Na_2SO_4 \cdot nH_2O$ contains 12.6 gm of water. The ' is:

[**Hint:** $Na_2SO_4 \cdot nH_2O$

Molar mass = (142 + 18n)

Mass of water $= \frac{12.6}{26.8} \times (142 + 18n)$

$$18n = \frac{12.6}{26.8} \times (142 + 18n)$$

$$n = 7]$$

29. Consider the following data:

Element	Atomic weight
A	12.01
B	35.5

A and B combine to form a new substance X. If 4 moles of B combine with 1 mole of A to give 1 mole of X, then the weight of 1 mole of X is:

(a) 154 g (b) 74 g (c) 47.5 g (d) 160 g

30. How many moles of Na^+ ions are in 20 mL of 0.4 M Na_3PO_4?

(a) 0.008 (b) 0.024 (c) 0.05 (d) 0.20

[**Hint:** No. of moles of $Na_3PO_4 = \frac{MV}{1000} = \frac{0.4 \times 20}{1000}$

$= 0.008$

Number of moles of $Na^+ = 3 \times$ Number of moles of Na_3PO_4

$= 3 \times 0.008 = 0.024$]

31. The element whose one atom has mass of 10.86×10^{-26} kg is:

(a) boron (b) calcium (c) silver (d) zinc

32. An electric discharge is passed through a mixture containing 50 cc of O_2 and 50 cc of H_2. The volume of the gases formed (i) at room temperature, (ii) at 110°C will be:

(a) (i) 25 cc (ii) 50 cc (b) (i) 50 cc (ii) 75 cc

(c) (i) 25 cc (ii) 75 cc (d) (i) 75 cc (ii) 75 cc

[**Hint:** $2H_2(g) + O_2(g) \longrightarrow 2H_2O$

50 cc H_2 will combine with 25 cc O_2 to form 50 cc H_2O

$\therefore$ O_2 left = 25 cc

At room temperature, H_2O will be in liquid state but at 110°C, it will be gaseous. Thus, volume of gases at 25°C and 110°C will be 25 cc and 75 cc respectively.]

33. The mass of carbon present in 0.5 mole of $K_4[Fe(CN)_6]$ is:

(a) 1.8 g (b) 18 g (c) 3.6 g (d) 36 g

[**Hint:** 1 mole of $K_4[Fe(CN)_6]$ contains 6 mole carbon, *i.e.*, 72 g carbon.]

34. Caffeine has a molecular weight of 194. If it contains 28.9% by mass of nitrogen, number of atoms of nitrogen in one molecule of caffeine is:

(a) 4 (b) 6 (c) 2 (d) 3

[**Hint:** Mass of nitrogen in 194 amu caffeine $= \frac{28.9}{100} \times 194$

$= 56$ amu

$\therefore$ One molecule of caffeine will contain 4 atoms of nitrogen.]

35. Chlorine can be prepared by reacting HCl with MnO_2. The reaction is represented by the equation,

$$MnO_2(s) + 4HCl(aq.) \longrightarrow Cl_2(g) + MnCl_2(aq.) + 2H_2O(l)$$

Assuming that the reaction goes to completion, what mass of conc. HCl solution (36% HCl by mass) is needed to produce 2.5 g

36. What is the mass per cent of oxygen in $Al_2(SO_4)_3 \cdot 18H_2O$? The molar mass of this substance is 666.43 g/mol :
(a) 9.60 (b) 28.8 (c) 43.2 (d) 72

37. 0.25 g of an element '*M*' reacts with excess fluorine to produce 0.547 g of the hexafluoride MF_6. What is the element?
(a) Cr (b) Mo (c) S (d) Te

38. How many electrons are present in 2×10^{-3} moles of ${}^{18}_{8}O^{2-}$?
(a) 1.2×10^{21} (b) 9.6×10^{21} (c) 1.2×10^{22} (d) 1.9×10^{22}

39. Fluorine reacts with uranium to form UF_6.

$$U(s) + 3F_2(g) \longrightarrow UF_6(g)$$

How many fluorine molecules are required to produce 2 mg of UF_6 from an excess of uranium? The molar mass of UF_6 is 352 g mol^{-1}.
(a) 3.4×10^{18} (b) 1×10^{19} (c) 2×10^{19} (d) 3.4×10^{21}

40. What is the formula of a substance with mass percentages of 35.79% for S, 62.92% for O and 1.13% for H?
(a) H_2SO_3 (b) H_2SO_4 (c) $H_2S_2O_7$ (d) $H_2S_2O_8$

41. In 1811, Avogadro calculated the formula of camphor by means of elemental chemical analysis and by measuring the density of its vapour. Avogadro found the density to be 3.84 g/L when he made the measurement at 210°C at 1 atm pressure. Which of the following is the correct formula of camphor?
(a) $C_{10}H_{14}O$ (b) $C_{10}H_{16}O$ (c) $C_{10}H_{16}O_2$ (d) $C_{10}H_{18}O$
(e) None of these

[**Hint:** $Pm = dRT$

$$m = \frac{dRT}{P} = \frac{3.84 \times 0.0821 \times 483}{1} = 152.27$$

$\therefore C_{10}H_{16}O$ will be the correct formula.]

42. A quantity of aluminium has a mass of 54 g. What is the mass of same number of magnesium atoms?
(a) 12.1 g (b) 24.3 g (c) 48.6 g (d) 97.2 g

43. When 1 L of CO_2 is heated with graphite, the volume of the gases collected is 1.5 L. Calculate the number of moles of CO produced at STP:
(a) $\frac{1}{11.2}$ (b) $\frac{28}{22.4}$ (c) $\frac{1}{22.4}$ (d) $\frac{14}{22.4}$

[**Hint:** $\underset{1-x}{CO_2(g)} + C(s) \longrightarrow \underset{2x}{2CO(g)}$

$$\text{Total volume} = 1 - x + 2x = 1 + x = 1.5$$
$$x = 0.5 \text{ L}$$
$$\therefore \text{ Volume of CO} = 2 \times 0.5 = 1 \text{ L}$$
$$\text{Number of moles of CO} = \frac{1}{22.4}\ ]$$

44. Which of the following has greatest number of atoms?
(a) 1 g of butane (C_4H_{10}) (b) 1 g of nitrogen (N_2)
(c) 1 g of silver (Ag) (d) 1 g of water (H_2O)

45. A metal oxide has the formula M_2O_3. It can be reduced by H_2 to give free metal and water. 0.1596 g of M_2O_3 required 6 mg of H_2 for complete reduction. The atomic mass of the metal is:
(a) 27.9 (b) 79.8
(c) 55.8 (d) 159.8

[**Hint:** $\underset{(2x+48)\text{ g}}{M_2O_3} + \underset{6\text{ g}}{3H_2} \longrightarrow 2M + 3H_2O$

x = Atomic mass of metal

$\because$ 0.006 g H_2 reduces 0.1596 g M_2O_3

$\therefore$ 6 g H_2 will reduce $\frac{0.1596}{0.006} \times 6$ g $M_2O_3 = 159.6\ M_2O_3$

$$2x + 48 = 159.6$$
$$2x = 111.6$$
$$x = 55.8]$$

46. In a compound of molecular formula A_mB_n:
(a) number of equivalents of A, B and A_mB_n are same
(b) number of moles of A, B and A_mB_n are same
(c) $m \times$ moles of $A = n \times$ moles of $B = (m+n) \times$ moles of A_mB_n
(d) $n \times$ moles of $A = m \times$ moles of $B = (m+n) \times$ moles of A_mB_n

47. 4.4 g of CO_2 and 2.24 litre of H_2 at STP are mixed in a container. The total number of molecules present in the container will be:
(a) 6.022×10^{23} (b) 1.2044×10^{23}
(c) 6.023×10^{26} (d) 6.023×10^{24}

48. A partially dried clay mineral contains 8% water. The original sample contained 12% water and 45% silica. The % of silica in the partially dried sample is nearly:
(a) 50% (b) 49%
(c) 55% (d) 47%

[**Hint:**

	Clay	Silica	Water
Initial stage:	43%	45%	12%
Final stage:	$(92 - x)$	x	8%

Ratio of silica and clay will remain constant, before and after drying.

$$\therefore \quad \frac{45}{43} = \frac{x}{92 - x}$$
$$x = 47\%]$$

49. Which of the following is isomorphous with $MgSO_4 \cdot 7H_2O$?
(a) Green vitriol (b) Blue vitriol
(c) Red vitriol (d) Vitriol of mass

50. In the reaction;

$$CO + \frac{1}{2}O_2 \longrightarrow CO_2; \quad N_2 + O_2 \longrightarrow 2NO$$

10 mL of mixture containing carbon monoxide and nitrogen required 7 mL oxygen to form CO_2 and NO, on combustion. The volume of N_2 in the mixture will be:
(a) 7/2 mL (b) 17/2 mL (c) 4 mL (d) 7 mL

[**Hint:** $\underset{x}{CO(g)} + \underset{x/2}{\frac{1}{2}O_2(g)} \longrightarrow CO_2(g)$

$\underset{y}{N_2(g)} + \underset{y}{O_2(g)} \longrightarrow 2NO(g)$

$$x + y = 10 \quad \ldots (i)$$
$$\frac{x}{2} + y = 7 \quad \ldots (ii)$$

Solving eqs. (i) and (ii),

$$x = 6 \text{ and } y = 4]$$

51. 1.44 g of titanium (Ti) reacted with excess of O_2 and produced x gm of a nonstoichiometric compound $Ti_{1.44}O_1$. The value of x is:

(b) 1.77
(d) None of these

he reaction:

$$Ti + O_2 \longrightarrow Ti_{1.44}O_1$$

noles of titanium = Number of moles of $Ti_{1.44}O_1$

$$\frac{1.44}{48} = \frac{x}{48 \times 1.44 + 16}$$

$x = 1.77$ g]

e of 75% alcohol by mass ($d = 0.8$ g/cm^3) must be pare 150 cc of 30% alcohol by mass (d = 0.9

L (b) 56.25 mL
(d) 33.56 mL

V mL of alcohol was used.

$$\frac{75}{100} \times V \times 0.8 = \frac{30}{100} \times 150 \times 0.9$$

$V = 67.5$ mL]

Following questions may have more than one correct options:

1. 11.2 L of a gas at STP weighs 14 g. The gas could be:
(a) N_2O (b) NO_2 (c) N_2 (d) CO
2. In which of the following pairs do 1 g of each have an equal number of molecules?
(a) N_2O and CO (b) N_2 and C_3O_2
(c) N_2 and CO (d) N_2O and CO_2
3. 8 g of oxygen has the same number of molecules as in:
(a) 11 g CO_2 (b) 22 g CO_2 (c) 7 g CO (d) 14 g CO
4. Which of the following has three significant figures?
(a) 6.02×10^{23} (b) 0.25
(c) 6.60×10^{-34} (d) 1.50
5. 1 mole of $^{14}_{7}N^{3-}$ ions contains:
(a) $7 \times 6.023 \times 10^{23}$ electrons (b) $7 \times 6.023 \times 10^{23}$ protons
(c) $7 \times 6.023 \times 10^{23}$ neutrons (d) $14 \times 6.023 \times 10^{23}$ protons
6. 1 g atom of nitrogen represents:
(a) 14 g nitrogen
(b) 11.2 litre of N_2 at NTP
(c) 22.4 litre of N_2 at NTP
(d) 6.023×10^{23} molecules of N_2

wers

ect option

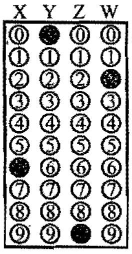
2. (b)	3. (a)	4. (d)	5. (b)	6. (c)	7. (a)	8. (a)
10. (a)	11. (d)	12. (b)	13. (c)	14. (d)	15. (c)	16. (b)
18. (b)	19. (d)	20. (a)	21. (a)	22. (d)	23. (c)	24. (c)
26. (a)	27. (c)	28. (d)	29. (a)	30. (b)	31. (d)	32. (c)
34. (a)	35. (b)	36. (d)	37. (b)	38. (c)	39. (b)	40. (c)
42. (c)	43. (c)	44. (a)	45. (c)	46. (a)	47. (b)	48. (d)
50. (c)	51. (b)	52. (c)				

e than one correct options

2. (c, d)	3. (a, c)	4. (a, c, d)	5. (b, c)	6. (a)

Integer Answer TYPE QUESTIONS

o each of the following questions is a teger, ranging from 0 to 9. If the correct estion numbers X, Y, Z and W (say) are respectively, then the correct darkening ll look like the given figure.



ous alkane (C_nH_{2n+2}) is exploded with oxygen. The of O_2 used and CO_2 formed are in the ratio of 7 : 4. the value of n.

$$C_nH_{2n+2} + \left[n + \frac{n+1}{2}\right]O_2 \longrightarrow nCO_2 + (n+1)\,H_2O$$

$$\frac{n + \frac{(n+1)}{2}}{n} = \frac{7}{4}$$

$$n = 2\,]$$

nany atoms do a mercury vapour molecule consist of, if pour density of mercury vapour relative to air is 6.92? ic mass of mercury is 200). The average molar mass of 9 g/mol.

$$\frac{\text{Vapour density of Hg vapour}}{\text{Vapour density of air}} = \frac{\text{Molar mass of Hg}}{\text{Molar mass of Air}}$$

$$\frac{6.92}{1} = \frac{m}{29}$$

$$m = 200\text{ g/mol}$$

mass is same as that of atomic mass hence mercury vapour ns monoatomic mercury.]

nole of an element contains 4.2×10^{24} electrons. What is omic number of the element?

romolecule of iron has molar mass 2800 amu, it contains on by mass. The number of iron atom in one formula unit macromolecule is:

: Number of iron atoms in one formula unit of compound

$$= \frac{\%}{100}\times\frac{\text{Molecular mass}}{\text{Atomic mass}} = \frac{8}{100}\times\frac{2800}{56} = 4\,]$$

es of 'A' and 10 moles of 'B' are mixed and allowed to ccording to the equation :

$$A + 3B \longrightarrow 2C$$

nany moles of C are present when there are 4 moles of A container?

nany water molecules will be there in 3×10^{-23} g sample er?

7. 5 g H_2 is allowed to react with 14 g N_2 for the following reaction:

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

What mass of H_2 will be left unreacted at the end of reaction?

[**Hint** :N_2 is limiting reactant, thus 14 g N_2 will give 17 g NH_3 and xg H_2 remains unreacted.

Mass before reaction = Mass after reaction

$$(5+14) \quad = \quad (17+x)$$

$$x = 2\text{g}\,]$$

8. Calculate the number of moles of water in 976 g $BaCl_2\cdot2H_2O$.
9. If Avogadro's number be 3.01×10^{23} then the atomic mass of carbon will be:
10. How many moles of R will be produced when 8 mol of P and 5 mol of Q are allowed to react according to the equation :

$$2P + Q \longrightarrow R$$

11. The mass of 1×10^{22} molecules of blue vitriol ($CuSO_4\cdot xH_2O$) is 4.144 g. The value of 'x' will be:
12. What will be the mass (in kg) of 7.298×10^6 mol electrons?
13. Silver (Atomic weight = 108 g mol^{-1}) has a density of 10.5 g cm^{-3}. The number of silver atoms on a surface area of $10^{-12}m^2$ can be expressed in scientific notation as $y\times10^x$. The value of x is: **(IIT 2010)**

[**Hint** : Mass of 1 cm^3 Ag = 1 × 10.5 g

$$\text{Number of atoms} = \frac{10.5}{108}\times6.023\times10^{23}$$

$$\text{Number of atoms in 1 cm} = \left[\frac{10.5}{108}\times6.023\times10^{23}\right]^{1/3}$$

$$\text{Number of atoms in 1 cm}^2 = \left[\frac{10.5}{108}\times6.023\times10^{23}\right]^{2/3}$$

Number of atoms in $10^{-2}m^2$ or $10^{-8}\,cm^2$

$$= \left[\frac{10.5}{108}\times6.023\times10^{23}\right]^{2/3}\times10^{-8} = 1.5\times10^7\,]$$

14. A student performs a titration with different burettes and finds titre values of 25.2 mL, 25.25 mL and 25.0 mL. The number of significant figures in average titre value is: **(IIT 2010)**

[**Hint** : Average titre value $= \dfrac{25.2 + 25.25 + 25.0}{3}$

$$= \frac{75.45}{3} = \frac{75.4}{3} = 25.1$$

(In addition, result is reported upto least place of decimal)]

nswers

2. (1) 3. (7) 4. (4) 5. (4) 6. (1) 7. (2) 8. (8)

10. (4) 11. (5) 12. (4) 13. (7) 14. (3)

LINKED COMPREHENSION TYPE QUESTIONS

1

y, 'mole' is an essential tool for the chemical t is a basic SI unit adopted by the 14th general weights and measurements in 1971. A mole contains ntary particles as the number of atoms present in 12 g of a gas at STP occupies 22.4 litre volume. Molar ids and liquids is not definite. Molar mass of a o called gram-atomic mass or gram molar mass. The g of mole is plenty, heap or the collection of large le of a substance contains 6.023×10^{23} elementary om or molecule. Atomic mass unit (amu) is the unit of g., atomic mass of single carbon is 12 amu.

lowing questions:

s of one amu is approximately:
(b) 0.5 g
10^{-24} g (d) 3.2×10^{-24} g

of a gas at STP are found to have a mass of 22 g. The r mass of the gas is:
(b) 44 (c) 88 (d) 33

s of one molecule of water is approximately:
(b) 0.5 g
10^{-24} g (d) 3.2×10^{-23} g

ny atoms are present in 49 g of H_2SO_4?
023×10^{23} (b) $5 \times 6.023 \times 10^{23}$
023×10^{23} (d) $7 \times 3.02 \times 10^{23}$

s at STP contains 3×10^{22} molecules. The number of s in x L ozone at STP will be:
22 (b) 4×10^{23} (c) 6.02×10^{23} (d) 3×10^{24}

lro's number is 1×10^{23} mol^{-1} then the mass of one xygen would be:
nu (b) 16×6.02 amu
1 (d) 16×10^{-23} amu

the Avogadro's number then number of valence n 4.8 g of O^{2-} is:
(b) 4.2 N_A (c) 1.6 N_A (d) 3.2 N_A

2

he atoms of same element; they have same atomic rent mass numbers. Isotopes have different number heir nucleus. If an element exists in two isotopes asses 'a' and 'b' in the ratio $m:n$, then average be $\dfrac{m \times a + n \times b}{m+n}$.

opes of same element have same position in the he elements which have single isotope are called nents. Greater is the percentage composition of an

Answer the following questions:

1. The isotopes of chlorine with mass number 35 and 37 exist in the ratio of..... Its average atomic mass is 35.5.
(a) 1:1 (b) 2:1 (c) 3:1 (d) 3:2
2. Which of the following isotopes is/are used to decide the scale of atomic mass?
(a) $^{12}_{6}C$ (b) $^{14}_{6}C$ (c) $^{16}_{8}O$ (d) $^{14}_{7}N$
3. Atomic mass of boron is 10.81. It has two isotopes namely $^{11}_{5}B$ and $^{x}_{5}B$ with their relative abundance of 80% and 20% respectively. The value of x is:
(a) 10.05 (b) 10 (c) 10.01 (d) 10.02
4. The ratio of the mass of ^{12}C atom to that of an atom of element X (whose atomicity is four) is 1:9. The molecular mass of element X is:
(a) 480 g mol^{-1} (b) 432 g mol^{-1}
(c) 36 g mol^{-1} (d) 84 g mol^{-1}
5. ^{12}C and ^{14}C isotopes are found as 98% and 2% respectively in any sample. Then, the number of ^{14}C atoms in 12 g of the sample will be:
(a) 1.5 mole atoms (b) 1.032×10^{22} atoms
(c) 2.06×10^{21} atoms (d) 2 g atoms

• Passage 3

Empirical formula is the simplest formula of the compound which gives the atomic ratio of various elements present in one molecule of the compound. However, the molecular formula of the compound gives the number of atoms of various elements present in one molecule of the compound.

$$\text{Molecular formula} = (\text{Empirical formula}) \times n$$

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

A compound may have same empirical and molecular formulae. Both these formulae are calculated by using percentage composition of constituent elements.

Answer the following questions:

1. Two metallic oxides contain 27.6% and 30% oxygen respectively. If the formula of first oxide is M_3O_4, that of second will be:
(a) MO (b) MO_2 (c) M_2O_5 (d) M_2O_3
2. Which of the following compounds have same empirical formula?
(a) Formaldehyde (b) Glucose
(c) Sucrose (d) Acetic acid
3. Which of the following represents the formula of a substance which contains 50% oxygen?
(a) N_2O (b) CO_2
(c) NO_2 (d) CH_3OH
4. An oxide of iodine (I = 127) contains 25.4 g of iodine and 8 g of oxygen. Its formula could be:
(a) I_2O_3 (b) I_2O
(c) I_2O_5 (d) I_2O_7

lrofluoric acid gas occupies 5.6 litres of volume at e empirical formula of the gas is HF, then its formula in the gaseous state will be:

(b) H_2F_2 (c) H_3F_3 (d) H_4F_4

species having different percentage composition of :

)OH and $C_6H_{12}O_6$ (b) CH_3COOH and C_2H_5OH CH_3 and HCOOH (d) C_2H_5OH and CH_3OCH_3

nd of Na, C and O contains 0.0887 mol Na, 0.132 2.65×10^{22} atoms of carbon. The empirical formula npound is:

(b) $Na_3C_5O_2$

)$_3$ (d) $Na_{0.0887}C_{2.65 \times 10^{22}}O_{0.132}$

4

loids are extracted from the extracts of the plants na. Marijuana owes its activity to tetrahydro iich contains 70% as many as carbon atoms as s and 15 times as many hydrogen atoms as oxygen m of tetrahydro cannabinol is 0.00318.

lowing questions:

ır mass of the compound is:

mu (b) 314 amu (c) 143 amu (d) 341 amu

ar formula of the compound is:

$_{30}O_2$ (b) $C_{21}H_{14}O_3$

$_{46}O$ (d) none of these

of oxygen atoms in 1 mol of the tetrahydro ol is:

(b) N_A (c) $3\,N_A$ (d) $4\,N_A$

$_A = 6.023 \times 10^{23}$

ge composition of carbon in the compound is:

% (b) 70.85% (c) 80.25% (d) 59.64%

5

sity of a compound is defined as the ratio of mass of a of gas to the mass of the same volume of hydrogen tical conditions of temperature and pressure.

$$= \frac{\textit{Mass of certain volume of gas (22.4 L) at STP}}{\textit{Mass of same volume of } H_2 \textit{ gas (22.4 L) at STP}}$$

$$= \frac{Mw}{2}$$

lecular mass of gas = Vapour density × 2

is a unitless quantity; it is unaffected by variation of nd pressure.

llowing questions:

density of a metal chloride is 66. Its oxide contains tal. The atomic mass of the metal is:

(b) 54 (c) 27.06 (d) 2.086

Number of equivalents of metal = Number of equivalents of oxygen

$$\frac{53}{E} = \frac{47}{8}$$

$$E = 9.02$$

Molecular formula of metal chloride = MCl_n

Molecular mass = $[n \times 9.02 + n \times 35.5] = 132$

$\therefore \quad n = 3$

$\therefore$ Atomic mass of metal = $3 \times 9.02 = 27.06$]

2. The vapour density of a mixture containing NO_2 and N_2O_4 is 38.3 at 27°C. The moles of NO_2 in 100 moles of mixture are:
   (a) 33.48 (b) 53.52 (c) 38.3 (d) 76.6
3. At STP, 5.6 litre of a gas weighs 60 g. The vapour density of gas is:
   (a) 60 (b) 120 (c) 30 (d) 240
4. Which of the following two substances have same vapour density?
   (a) Glucose (b) Fructose (c) Sucrose (d) Starch
5. Let $NH_4HS(s)$ is heated in a closed vessel to decompose.

$$NH_4HS(s) \rightleftharpoons NH_3(g) + H_2S(g)$$

The vapour density of the mixture will be:
(a) equal to that of NH_4HS
(b) lesser than that of NH_4HS
(c) greater than that of NH_4HS
(d) cannot be predicted

• Passage 6

Precision refers to the closeness of a set of values obtained for identical measurement of a quantity. Precision depends on the limitations of measuring devices and the skills with which it is used. However, accuracy refers to the closeness of a single measurement to its true value.

The digits in a properly recorded measurement are known as significant figures. These are meaningful digits in a measured or calculated quantity. The greater the number of significant figures in a reported result, smaller is the uncertainty and greater is the precision. The zeros at the beginning are not counted. The zeros to the right of a decimal point are counted. In the numbers that do not contain a decimal point, "trailing" zeros may or may not be significant. The purpose of zeros at the end of a number is to convey the correct range of uncertainty.

Answer the following questions:

1. If repeated measurements give values close to one another, the number is:
   (a) surely precise (b) surely accurate
   (c) surely precise and accurate (d) all of these are correct
2. The number of significant figures in a measured number contains how many uncertain number of digits?
   (a) Zero (b) 1
   (c) 2 (d) Cannot be predicted
3. In the number 2.4560, there are 5 significant digits. Which one is the least significant digit?
   (a) 2 (b) 4 (c) 0 (d) 6
4. If we add 296.2 and 2.256, we get the answer as 298.456 g. The number of significant figures in the result are:
   (a) 6 (b) 5 (c) 4 (d) 3
5. In which of the following numbers, all the zeros are not significant?
   (a) 0.0010 (b) 0.00100 (c) 0.001000 (d) 0.001

Answers

Passage 1.	1. (c)	2. (c)	3. (d)	4. (d)	5. (a)	6. (c)	7. (a)
Passage 2.	1. (c)	2. (a, c)	3. (b)	4. (b)	5. (b)		
Passage 3.	1. (d)	2. (a, b)	3. (d)	4. (c)	5. (b)	6. (b, c)	7. (a)
Passage 4.	1. (b)	2. (a)	3. (a)	4. (c)			
Passage 5.	1. (c)	2. (c)	3. (b)	4. (a, b)	5. (b)		
Passage 6.	1. (a)	2. (b)	3. (c)	4. (c)	5. (d)		

SELF ASSESSMENT

ASSIGNMENT NO. 1

SECTION-I

Straight Objective Type Questions

This section contains 10 multiple choice questions. Each question has 4 choices (a), (b), (c) and (d), out of which only one is correct.

1. Cartisone is a molecular substance containing 21 atoms of carbon per molecule. The mass percentage of carbon in cartisone is 69.98%. What is the molecular mass of cartisone?
(a) 360.4 (b) 176.5 (c) 287.6 (d) 312.8

2. Total number of atoms present in 25 mg of camphor, $C_{10}H_{16}O$ is:
(a) 2.57×10^{21} (b) 9.89×10^{19}
(c) 2.67×10^{21} (d) 6.02×10^{20}

3. The oxide of a metal contains 60% of the metal. What will be the percentage of bromine in the bromide of the metal, if the valency of the metal is the same in both, the oxide and the bromide?
(a) 93% (b) 87% (c) 70% (d) 77%

4. The radius of water molecule having density $1\,g\ mL^{-1}$ is:
(a) 1.925 Å (b) 73.46 Å (c) 19.25 Å (d) 7.346 Å

5. 3 g of an oxide of a metal is converted completely to 5 g chloride. Equivalent mass of metal is:
(a) 33.25 (b) 3.325 (c) 12 (d) 20

6. Quantitative analysis of a compound shows that it contains 0.110 mole of 'C', 0.055 mole of 'N' and 0.165 mole of 'O'. Its molecular mass is about 270. How many atoms of carbon are there in empirical and molecular formulae of the compound respectively?

	Empirical formula	Molecular formula
(a)	1	3
(b)	2	2
(c)	2	6
(d)	3	2

7. Total number of electrons present in 11.2 L of NH_3 at STP is:
(a) 6.02×10^{23} (b) 3.01×10^{23}
(c) 3.01×10^{24} (d) 5.1×10^{24}

8. Which one of the following is not a unit of length?
(a) Angstrom (b) Light-year
(c) Micron (d) Radian

9. Unit of $J\ pa^{-1}$ is equivalent to:
(a) m^3 (b) cm^3
(c) dm^3 (d) none of these

10. The relative abundance of two isotopes of atomic masses 85 and 87 are 75% and 25% respectively. The average atomic mass of element is:
(a) 86 (b) 40 (c) 85.5 (d) 75.5

SECTION-II

Multiple Answers Type Objective Questions

11. Mass of one atom of oxygen is/are:
(a) 16 amu (b) 32 amu
(c) 16 gm (d) 2.656×10^{-23} gm

12. Which of the following compounds have same percentage composition of carbon?
(a) $C_6H_{12}O_6$ (b) CH_3COOH
(c) $HCOOCH_3$ (d) $C_{12}H_{22}O_{11}$

13. Which of the following is/are correct about 1 mole electrons?
(a) 6.023×10^{23} electrons (b) 5.48×10^{-7} kg
(c) 96500 coulomb charge (d) None of these

14. In which of the following numbers, all zeros are significant?
(a) 5.0005 (b) 0.0030 (c) 30.000 (d) 0.5200

15. Which of the following are correct SI units?
(a) Amount of substance in $mol\ L^{-1}$
(b) Pressure of gas in pascal
(c) Density of a solid in $kg\ m^{-3}$
(d) Force in newton

SECTION-III

Assertion-Reason Type Questions

This section contains 4 questions. Each question contains **Statement-1** (Assertion) and **Statement-2** (Reason). Each question has following 4 choices (a), (b), (c) and (d), out of which only one is correct.

(a) Statement-1 is true; statement-2 is true; statement-2 is a correct explanation for statement-1.
(b) Statement-1 is true; statement-2 is true; statement-2 is not a correct explanation for statement-1.
(c) Statement-1 is true; statement-2 is false.
(d) Statement-1 is false; statement-2 is true.

16. **Statement-1:** Avogadro's number is a dimensionless quantity.

Because

Statement-2: It is a number of atoms or molecules in one gram mole.

17. **Statement-1:** An element has variable equivalent mass.

Because

Statement-2: The valency of element is variable.

18. **Statement-1:** Vapour density of CH_4 is half of O_2.

Because

Statement-2: 1.6 g of CH_4 contains same number of electrons as 3.2 g of O_2.

19. **Statement-1:** Specific gravity is dimensionless quantity.

Because

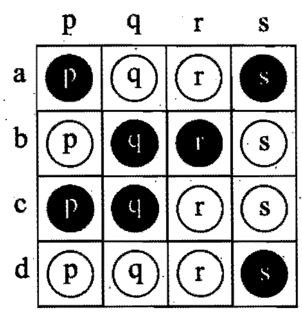
Statement-2: Specific gravity is relative density of a substance, measured with respect to density of water at 4°C.

SECTION-IV

Matrix-Matching Type Questions

This section contains 3 questions. Each question contains statement given in two columns which have to be matched. Statements (a, b, c and d) in Column-I have to be matched with statements (p, q, r and s) in Column-II. The answers to these questions have to be appropriately bubbled as illustrated in the following examples:

If the correct matches are (a-p,s); (b-q,r); (c-p,q) and (d-s); then the correctly bubbled 4 × 4 matrix should be as follows:



20. Match the Column-I with Column-II:

Column-I	**Column-II**
(a) N_2	(p) 40% carbon by mass
(b) CO	(q) Empirical formula CH_2O
(c) $C_6H_{12}O_6$	(r) Vapour density = 14
(d) CH_3COOH	(s) $14N_A$ ($N_A = 6.023 \times 10^{23}$) electrons in a mole

21. Match the Column-I with Column-II:

Column-I	**Column-II**
(a) 1 L	(p) 10^{-5} N
(b) 1 J	(q) 0.2389 cal
(c) 9.9×10^6 erg	(r) 10^{-3} m^3
(d) 1 Dyne	(s) 6.25×10^{18} eV

22. Match the Column-I with Column-II:

Column-I	**Column-II**
(a) 1 g mole of $O_2(g)$	(p) mass, 32 g
(b) 0.5 mole of $SO_2(g)$	(q) mass, 24 g
(c) 1 g of $H_2(g)$	(r) volume, 11.2 L at STP
(d) 0.5 mole of $O_3(g)$	(s) $1.5 \times 6.023 \times 10^{23}$ atoms

Answers

1. (a) **2.** (c) **3.** (b) **4.** (a) **5.** (a) **6.** (c) **7.** (c) **8.** (d)
9. (a) **10.** (c) **11.** (a, d) **12.** (a, b, c) **13.** (a, b, c) **14.** (a, c) **15.** (b, c, d) **16.** (a)
17. (b) **18.** (c) **19.** (a) **20.** (a-r, s) (b-r, s) (c-p, q) (d-p, q) **21.** (a-r) (b-q, s) (c-q) (d-p)
22. (a-p) (b-p, r, s) (c-r) (d-q, s)