

Fundamentals of Chemistry

Major Concepts

- 1.1 Branches of Chemistry
- 1.2 Basic Definitions
- 1.3 Chemical species
- 1.4 Avogadro's Number and Mole
- 1.5 Chemical Calculations

Time allocation

Teaching periods	12
Assessment periods	03
Weightage	10%

Students Learning Outcomes

Students will be able to:

- Identify and provide examples of different branches of chemistry.
- Differentiate among branches of chemistry.
- Distinguish between matter and a substance.
- Define ions, molecular ions, formula units and free radicals.
- Define atomic number, atomic mass, atomic mass unit.
- Differentiate among elements, compounds and mixtures.
- Define relative atomic mass based on C-12 scale.
- Differentiate between empirical and molecular formula.
- Distinguish between atoms and ions.
- Differentiate between molecules and molecular ions.
- Distinguish between ion and free radicals.
- Classify the chemical species from given examples.
- Identify the representative particles of elements and compounds.
- Relate gram atomic mass, gram molecular mass and gram formula mass to mole.
- Describe how Avogadro's number is related to a mole of any substance.
- Distinguish among the terms gram atomic mass, gram molecular mass and gram formula mass.
- Change atomic mass, molecular mass and formula mass into gram atomic mass, gram molecular mass and gram formula mass.

Introduction

The knowledge that provides understanding of this world and how it works, is science. *The branch of science which deals with the composition, structure, properties and reactions of matter is called chemistry.* It deals with every aspect of our life.

The development of science and technology has provided us a lot of facilities in daily life. Imagine the role and importance of petrochemical products, medicines and drugs, soap, detergents, paper, plastics, paints and pigments, insecticides, pesticides which all are fruit of the efforts of chemists. The development of chemical industry has also generated toxic wastes, contaminated water and polluted air around us. On the other hand, chemistry also provides knowledge and techniques to improve our health and environment and to explore and to conserve the natural resources.

In this chapter, we will study about different branches of chemistry, basic definitions and concepts of chemistry.

1.1 BRANCHES OF CHEMISTRY

It is a fact that we live in the world of chemicals. We all depend upon different living organisms which require water, oxygen or carbon dioxide for their survival. Today chemistry has a wide scope in all aspects of life and is serving the humanity day and night. Chemistry is divided into following main branches: physical chemistry, organic chemistry, inorganic chemistry, biochemistry, industrial chemistry, nuclear chemistry, environmental chemistry and analytical chemistry.

1.1.1 Physical Chemistry

Physical Chemistry is defined as the *branch of chemistry that deals with the relationship between the composition and physical properties of matter along with the changes in them.* The properties such as structure of atoms or formation of molecules behavior of gases, liquids and solids and the study of the effect of temperature or radiation on matter are studied under this branch.

1.1.2 Organic Chemistry

Organic Chemistry is the study of *covalent compounds of carbon and hydrogen (hydrocarbons) and their derivatives.* Organic compounds occur naturally and are also synthesized in the laboratories. Organic chemists determine the structure and properties of these naturally occurring as well as synthesized compounds. Scope of this branch covers petroleum, petrochemicals and pharmaceutical industries.

1.1.3 Inorganic Chemistry

Inorganic chemistry *deals with the study of all elements and their compounds except those of compounds of carbon and hydrogen (hydrocarbons) and their derivatives.* It has applications in every aspect of the chemical industry such as glass, cement, ceramics and metallurgy (extraction of metals from ores).

1.1.4 Biochemistry

It is *the branch of chemistry in which we study the structure, composition, and chemical reactions of substances found in living organisms*. It covers all chemical processes taking place in living organisms, such as synthesis and metabolism of biomolecules like carbohydrates, proteins and fats. Biochemistry emerged as a separate discipline when scientists began to study how living things obtain energy from food or how the fundamental biological changes occur during a disease. Examples of applications of biochemistry are in the fields of medicine, food science and agriculture, etc.

1.1.5 Industrial Chemistry

The *branch of chemistry that deals with the manufacturing of chemical compounds on commercial scale, is called industrial chemistry*. It deals with the manufacturing of basic chemicals such as oxygen, chlorine, ammonia, caustic soda, nitric acid and sulphuric acid. These chemicals provide the raw materials for many other industries such as fertilizers, soap, textiles, agricultural products, paints and paper, etc.

1.1.6 Nuclear Chemistry

Nuclear Chemistry is the *branch of chemistry that deals with the radioactivity, nuclear processes and properties*. The main concern of this branch is with the atomic energy and its uses in daily life. It also includes the study of the chemical effects resulting from the absorption of radiation within animals, plants and other materials. It has vast applications in medical treatment (radiotherapy), preservation of food and generation of electrical power through nuclear reactors, etc.

1.1.7 Environmental Chemistry

It is the branch of chemistry in which we study about components of the environment and the effects of human activities on the environment. Environmental chemistry is related to other branches like biology, geology, ecology, soil and water. The knowledge of chemical processes taking place in environment is necessary for its improvement and protection against pollution.

1.1.8 Analytical Chemistry

Analytical chemistry is *the branch of chemistry that deals with separation and analysis of a sample to identify its components*. The separation is carried out prior to qualitative and quantitative analysis. Qualitative analysis provides the identity of a substance (composition of chemical species). On the other hand, quantitative analysis determines the amount of each component present in the sample. Hence, in this branch different techniques and instruments used for analysis are studied. The scope of this branch covers food, water, environmental and clinical analysis.



- i. In which branch of chemistry behaviour of gases and liquids is studied?
- ii. Define biochemistry?
- iii. Which branch of chemistry deals with preparation of paints and paper?
- iv. In which branch of chemistry the metabolic processes of carbohydrates and proteins are studied?
- v. Which branch of chemistry deals with atomic energy and its uses in daily life?
- vi. Which branch of chemistry deals with the structure and properties of naturally occurring molecules?

1.2 BASIC DEFINITIONS

Matter is simply defined as *anything that has mass and occupies space*. Our bodies as well as all the things around us are examples of matter. In chemistry, we study all types of matters that can exist in any of three physical states: solid, liquid or gas.

A piece of matter in pure form is termed as a **substance**. Every substance has a fixed composition and specific properties or characteristics. Whereas, impure matter is called a **mixture**; which can be homogeneous or heterogeneous in its composition.

We know that every substance has physical as well as chemical properties. *The properties those are associated with the physical state of the substance are called **physical properties*** like colour, smell, taste, hardness, shape of crystal, solubility, melting or boiling points, etc. For example, when ice is heated, it melts to form water. When water is further heated, it boils to give steam. In this entire process only the physical states of water change whereas its chemical composition remains the same.

The **chemical properties** depend upon the composition of the substance. When a substance undergoes a chemical change, its composition changes and a new substances are formed. For example, decomposition of water is a chemical change as it produces hydrogen and oxygen gases. All materials are either a substance or a mixture. Figure 1.1 shows simple classification of the matter into different forms.

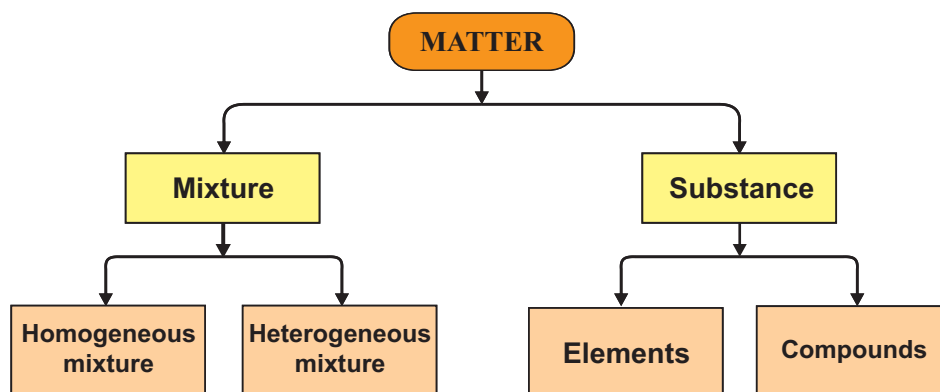


Fig. 1.1. Classification of matter

1.2.1 Elements, Compounds and Mixtures

1.2.1.1 Elements

In the early ages, only nine elements (carbon, gold, silver, tin, mercury, lead, copper, iron and sulphur) were known. At that time, it was considered that elements were the substances that could not be broken down into simpler units by ordinary chemical processes. Until the end of nineteenth century, sixty-three elements had been discovered. Now 118 elements have been discovered, out of which 92 are naturally occurring elements. Modern definition of **element** is that *it is a substance made up of same type of atoms, having same atomic number and cannot be decomposed into simple substances by ordinary chemical means*. It means that each element is made up of unique type of atoms that have very specific properties.

Elements occur in nature in free or combined form. All the naturally occurring elements found in the world have different percentages in the earth's crust, oceans and atmosphere. Table 1.1. shows natural occurrence in percentage by weight of some major elements around us. It shows concentrations of these major elements found in the three main systems of our environment.

Table 1.1 Natural Occurrences by Weight % of Some Major Elements

Earth's crust		Oceans		Atmosphere	
Oxygen	47 %	Oxygen	86 %	Nitrogen	78 %
Silicon	28 %	Hydrogen	11 %	Oxygen	21%
Aluminium	7.8 %	Chlorine	1.8 %	Argon	0.9%

Elements may be solids, liquids or gases. Majority of the elements exist as solids e.g. sodium, copper, zinc, gold, etc. There are very few elements which occur in liquid state e.g. mercury and bromine. A few elements exist as gases e.g. nitrogen, oxygen, chlorine and hydrogen.

On the basis of their properties, elements are divided into metals, non-metals and metalloids. About 80 percent of the elements are metals.



Do you know?

Major part of a living body is made up of water i.e. 65% to 80% by mass.

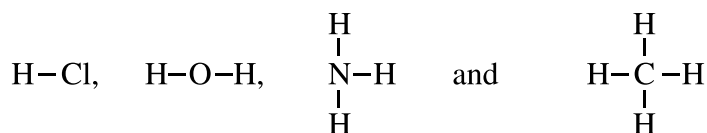
Six elements constitute about 99% of our body mass; namely: Oxygen 65 %, Carbon 18%, Hydrogen 10 %, Nitrogen 3%, Calcium 1.5% and Phosphorus 1.5%.

Potassium, Sulphur, Magnesium and Sodium constitute 0.8% of our body mass. Whereas Copper, Zinc, Fluorine, Chlorine, Iron, Cobalt and Manganese constitute only 0.2% of our body mass.

Elements are represented by **symbols**, which are abbreviations for the name of elements. A symbol is taken from the name of that element in English, Latin, Greek or German. If it is one letter, it will be capital as H for Hydrogen, N for Nitrogen and C for Carbon etc. In case of two letters symbol, only first letter is capital e.g. Ca for Calcium, Na for Sodium and Cl for Chlorine.

The unique property of an element is **valency**. It is combining capacity of an element with other elements. It depends upon the number of electrons in the outermost shell.

In simple covalent compounds, valency is the number of hydrogen atoms which combine with one atom of that element or the number of bonds formed by one atom of that element e.g. in the following compounds.



The valency of chlorine, oxygen, nitrogen and carbon is 1, 2, 3 and 4, respectively.

In simple ionic compounds valency is the number of electrons gained or lost by an atom of an element to complete its octet. Elements having less than four electrons in their valence shell; prefer to lose the electrons to complete their octet. For example, atoms of Na, Mg and Al have 1, 2 and 3 electrons in their valence shells respectively. They lose these electrons to have valency of 1, 2 and 3, respectively. On the other hand, elements having five or more than five electrons in their valence shells, gain electrons to complete their octet. For example, N, O and Cl have 5, 6 and 7 electrons in their valence shells respectively. They gain 3, 2 and 1 electrons respectively to complete their octet. Hence, they show valency of 3, 2 and 1, respectively. A radical is a group of atoms that have some charge. Valencies of some common elements and radicals are shown in Table 1.2.

Table 1.2 Some Elements and Radicals with their Symbols and Common Valencies

Element / Radical	Symbol	Valency	Element / Radical	Symbol	Valency
Sodium	Na	1	Hydrogen	H	1
Potassium	K	1	Chlorine	Cl	1
Silver	Ag	1	Bromine	Br	1
Magnesium	Mg	2	Iodine	I	1
Calcium	Ca	2	Oxygen	O	2
Barium	Ba	2	Sulphur	S	2
Zinc	Zn	2	Nitrogen	N	3
Copper	Cu	1,2	Phosphorus	P	3,5
Mercury	Hg	1,2	Boron	B	3
Iron	Fe	2,3	Arsenic	As	3
Aluminium	Al	3	Carbon	C	4
Chromium	Cr	3	Carbonate	CO ₃ ²⁻	2
Ammonium	NH ₄ ⁺	1	Sulphate	SO ₄ ²⁻	2
Hydronium	H ₃ O ⁺	1	Sulphite	SO ₃ ²⁻	2
Hydroxide	OH ⁻	1	Thiosulphate	S ₂ O ₃ ²⁻	2
Cyanide	CN ⁻	1	Nitride	N ³⁻	3
Bisulphate	HSO ₄ ⁻	1	Phosphate	PO ₄ ³⁻	3
Bicarbonate	HCO ₃ ⁻	1			

Some elements show more than one valency, i.e. they have variable valency. For example, in ferrous sulphate (FeSO₄) the valency of iron is 2. In ferric sulphate (Fe₂(SO₄)₃), the valency of iron is 3. Generally, the Latin or Greek name for the element (e.g., Ferrum) is modified to end in 'ous' for the lower valency (e.g. Ferrous) and to end in 'ic' for the higher valency (e.g. Ferric).

1.2.1.2 Compound

Compound is a *substance made up of two or more elements chemically combined together in a fixed ratio by mass*. As a result of this combination, elements lose their own properties and produce new substances (compounds) that have entirely different properties. Compounds can't be broken down into its constituent elements by simple physical methods. For example, carbon dioxide is formed when elements of carbon and oxygen combine chemically in a fixed ratio of 12:32 or 3:8 by mass. Similarly, water is a compound formed by a chemical combination between hydrogen and oxygen in a fixed ratio of 1:8 by mass.

Compounds can be classified as ionic or covalent. Ionic compounds do not exist in independent molecular form. They form a three dimensional crystal lattice, in which each ion is surrounded by oppositely charged ions. These oppositely charged ions attract each other very strongly, as a result ionic compounds have high melting and boiling points. These compounds are represented by **formula units** e.g. NaCl, KBr, CuSO₄.

The covalent compounds mostly exist in molecular form. A molecule is a true representative of the covalent compound and its formula is called **molecular formula** e.g. H₂O, HCl, H₂SO₄, CH₄.

Table 1.3 Some Common Compounds with their Formulae

Compound	Chemical Formula
Water	H ₂ O
Sodium chloride (Common salt)	NaCl
Silicon dioxide (Sand)	SiO ₂
Sodium hydroxide (Caustic Soda)	NaOH
Sodium carbonate (Washing Soda)	Na ₂ CO ₃ .10H ₂ O
Calcium oxide (Quick Lime)	CaO
Calcium carbonate (Limestone)	CaCO ₃
Sugar	C ₁₂ H ₂₂ O ₁₁
Sulphuric acid	H ₂ SO ₄
Ammonia	NH ₃

Remember

Always use:

- Standard symbols of elements
- Chemical formulae of compounds
- Proper abbreviations of scientific terms
- Standard values and SI units for constants

1.2.1.3 Mixture

When two or more elements or compounds mix up physically without any fixed ratio, they form a **mixture**. On mixing up, the component substances retain their own chemical identities and properties. The mixture can be separated into parent components by physical methods such as distillation, filtration, evaporation, crystallisation or magnetization. *Mixtures that have uniform composition throughout are called **homogeneous mixtures** e.g. air, gasoline, ice cream. Whereas, **heterogeneous mixtures** are those in which composition is not uniform throughout e.g. soil, rock and wood.*



Air is a mixture of nitrogen, oxygen, carbon dioxide, noble gases and water vapours.

Soil is a mixture of sand, clay, mineral salts, water and air.

Milk is a mixture of water, sugar, fat, proteins, mineral salts and vitamins.

Brass is a mixture of copper and zinc metals.

Table 1.4 Difference between a Compound and a Mixture

	Compound	Mixture
i.	It is formed by a chemical combination of atoms of the elements.	Mixture is formed by the simple mixing up of the substances.
ii.	The constituents lose their identity and form a new substance having entirely different properties from them.	Mixture shows the properties of the constituents.
iii.	Compounds always have fixed composition by mass.	Mixtures do not have fixed composition.
iv.	The components cannot be separated by physical means.	The components can be separated by simple physical methods.
v.	Every compound is represented by a chemical formula.	It consists of two or more components and does not have any chemical formula.
vi.	Compounds have homogeneous composition.	They may be homogeneous or heterogeneous in composition
vii.	Compounds have sharp and fixed melting points	Mixtures do not have sharp and fixed melting points.



- Can you identify mixture, element or compound out of the following:
Coca cola, petroleum, sugar, table salt, blood, gun powder, urine, aluminium, silicon, tin, lime and ice cream.
- How can you justify that air is a homogenous mixture. Identify substances present in it.
- Name the elements represented by the following symbols:
Hg, Au, Fe, Ni, Co, W, Sn, Na, Ba, Br, Bi.
- Name a solid, a liquid and a gaseous element that exists at the room temperature.
- Which elements do the following compounds contain?
Sugar, common salt, lime water and chalk.

1.2.1 Atomic Number and Mass Number

The **atomic number** of an element is equal to the number of protons present in the nucleus of its atoms. It is represented by symbol '**Z**'. As all atoms of an element have the same number of protons in their nuclei, they have the same atomic number.

Hence, each element has a specific atomic number termed as its identification number. For example, all hydrogen atoms have 1 proton, their atomic number is $Z=1$. All atoms in carbon have 6 protons, their atomic number is $Z=6$. Similarly, in oxygen all atoms have 8 protons having atomic number $Z=8$ and sulphur having 16 protons shows atomic number $Z=16$.

The **mass number** is the sum of number of protons and neutrons present in the nucleus of an atom. It is represented by symbol ' A '.

It is calculated as $A=Z+n$ where n is the number of neutrons.

Each proton and neutron has same mass. For example, hydrogen atom has one proton and no neutron in its nucleus, its mass number $A=1+0=1$. Carbon atom has 6 protons and 6 neutrons, hence its mass number $A=12$. Atomic numbers and mass numbers of a few elements are given in Table 1.5

Table 1.5 Some Elements along with their Atomic and Mass Numbers

Element	Number of Protons	Number of Neutrons	Atomic Number Z	Mass Number A
Hydrogen	1	0	1	1
Carbon	6	6	6	12
Nitrogen	7	7	7	14
Oxygen	8	8	8	16
Fluorine	9	10	9	19
Sodium	11	12	11	23
Magnesium	12	12	12	24
Potassium	19	20	19	39
Calcium	20	20	20	40

Example 1.1

How many protons and neutrons are there in an atom having $A = 238$ and $Z = 92$.

Solution:

First of all, develop data from the given statement of the example and then solve it with the help of data.

Data

$$A=238$$

$$Z=92$$

Number of protons ?

Number of neutrons?

$$\text{Number of protons} = Z = 92$$

$$\begin{aligned}\text{Number of Neutrons} &= A - Z \\ &= 238 - 92 \\ &= 146\end{aligned}$$

1.2.3 Relative Atomic Mass and Atomic Mass Unit

As we know that the mass of an atom is too small to be determined practically. However, certain instruments enable us to determine the ratio of the atomic masses of various elements to that of carbon-12 atoms. This ratio is known as the relative atomic mass of the element. The **relative atomic mass** of an element is the average mass of the atoms of that element as compared to $1/12^{\text{th}}$ (one-twelfth) the mass of an atom of carbon-12 isotope (an element having different mass number but same atomic number). Based on carbon-12 standard, the mass of an atom of carbon is 12 units and $1/12^{\text{th}}$ of it comes to be 1 unit. When we compare atomic masses of other elements with atomic mass of carbon-12 atom, they are expressed as relative atomic masses of those elements. The unit for relative atomic masses is called **atomic mass unit**, with symbol '**amu**'. One atomic mass unit is $1/12^{\text{th}}$ the mass of one atom of carbon-12th. When this atomic mass unit is expressed in **grams**, it is:

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

For example:

$$\begin{aligned}\text{Mass of a proton} &= 1.0073 \text{ amu} && \text{or } 1.672 \times 10^{-24} \text{ g} \\ \text{Mass of a neutron} &= 1.0087 \text{ amu} && \text{or } 1.674 \times 10^{-24} \text{ g} \\ \text{Mass of an electron} &= 5.486 \times 10^{-4} \text{ amu.} && \text{or } 9.106 \times 10^{-28} \text{ g}\end{aligned}$$

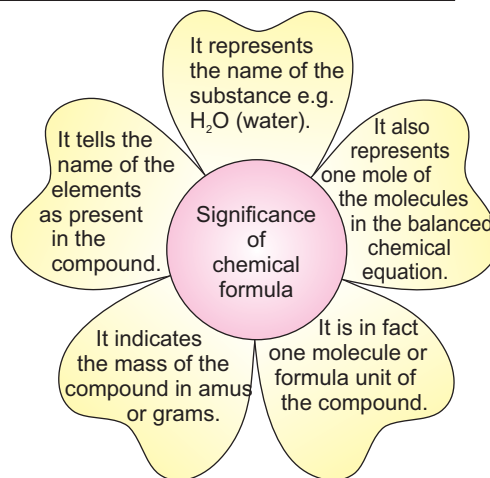


- i) How many amu 1 g of a substance has?
- ii) Is atomic mass unit a SI unit of an atomic mass?
- iii) What is the relationship between atomic number and atomic mass?
- iv) Define relative atomic mass.
- v) Why atomic mass of an atom is defined as relative atomic mass?

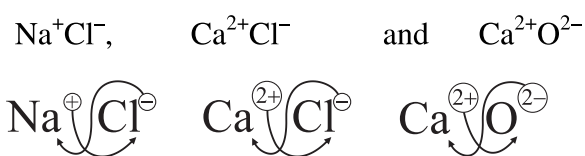
1.2.4 How to write a Chemical Formula

Compounds are represented by **chemical formulae** as elements are represented by symbols. Chemical formulae of compounds are written keeping the following steps in consideration.

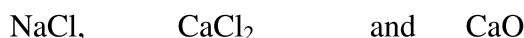
- i. Symbols of two elements are written side by side, in the order of positive ion first and negative ion later.
- ii. The valency of each ion is written on the right top corner of its symbol, e.g. Na^+ , Ca^{2+} , Cl^- and O^{2-} .



- iii. This valency of each ion is brought to the lower right corner of other ion by 'cross-exchange' method, e.g.



They are written as:



- iv. If the valencies are same, they are offset and are not written in the chemical formula. But if they are different, they are indicated as such at the same position, e.g. in case of sodium chloride both the valencies are offset and formula is written as NaCl, whereas, calcium chloride is represented by formula CaCl_2 .
- v. If an ion is a combination of two or more atoms which is called radical, bearing a net charge on it, e.g. SO_4^{2-} (sulphate) and PO_4^{3-} (phosphate), then the net charge represents the valency of the radical. The chemical formula of such compounds is written as explained in (iii) and (iv); writing the negative radical within the parenthesis. For example, chemical formula of aluminium sulphate is written as $\text{Al}_2(\text{SO}_4)_3$ and that of calcium phosphate as $\text{Ca}_3(\text{PO}_4)_2$.

1.2.4.1 Empirical formula

Chemical formulae are of two types. The simplest type of formula is **empirical formula**. It is the *simplest whole number ratio of atoms present in a compound*. The empirical formula of a compound is determined by knowing the percentage composition of a compound. However, here we will explain it with simple examples.

The covalent compound silica (sand) has simplest ratio of 1:2 of silicon and oxygen respectively. Therefore, its empirical formula is SiO_2 . Similarly, glucose has simplest ratio 1:2:1 of carbon, hydrogen and oxygen, respectively. Hence, its empirical formula is CH_2O .

As discussed earlier, the ionic compounds exist in three dimensional network forms. Each ion is surrounded by oppositely charged ions in such a way to form electrically neutral compound. Therefore, the simplest unit taken as a representative of an ionic compound is called **formula unit**. It is *defined as the simplest whole number ratio of ions, as present in the ionic compound*. In other words, ionic compounds have only empirical formulae. For example, formula unit of common salt consists of one Na^+ and one Cl^- ion and its empirical formula is NaCl. Similarly, formula unit of potassium bromide is KBr, which is also its empirical formula.

1.2.4.2 Molecular Formula

Molecules are formed by the combination of atoms. These molecules are represented by **molecular formulae** that show *actual number of atoms of each element present in a molecule of that compound*. Molecular formula is derived from empirical formula by the following relationship:

$$\text{Molecular formula} = (\text{Empirical formula})_n$$

Where n is 1,2,3 and so on.

For example, molecular formula of benzene is C_6H_6 which is derived from the empirical formula CH where the value of n is 6.

The molecular formula of a compound may be same or a multiple of the empirical formula. A few compounds having different empirical and molecular formulae are shown in Table 1.6.

Table 1.6 Some Compounds with their Empirical and Molecular Formulae

Compound	Empirical formula	Molecular formula
Hydrogen peroxide	HO	H_2O_2
Benzene	CH	C_6H_6
Glucose	CH_2O	$\text{C}_6\text{H}_{12}\text{O}_6$

Some compounds may have same empirical and molecular formula e.g. water (H_2O), hydrochloric acid (HCl), etc.

1.2.5 Molecular Mass and Formula Mass

The sum of atomic masses of all the atoms present in one molecule of a molecular substance, is its **molecular mass**. For example, molecular mass of chlorine (Cl_2) is 71.0 amu, of water (H_2O) is 18 amu and that of carbon oxide (CO_2) is 44 amu.

Example 1.2

Calculate the molecular mass of Nitric acid, HNO_3 .

Solution

Atomic mass of H = 1 amu

Atomic mass of N = 14 amu

Atomic mass of O = 16 amu

Molecular formula = HNO_3

$$\begin{aligned} \text{Molecular mass} &= 1 (\text{At. mass of H}) + 1 (\text{At. mass of N}) + 3 (\text{At. mass of O}) \\ &= 1 + 14 + 3(16) \\ &= 1 + 14 + 48 \\ &= 63 \text{ amu} \end{aligned}$$

Some ionic compounds that form three dimensional solid crystals, are represented by their formula units. **Formula mass** in such cases is the sum of atomic masses of all the atoms present in one formula unit of a substance. For example, formula mass of sodium chloride is 58.5 amu and that of CaCO_3 is 100 amu.

Example 1.3

Calculate the formula mass of Potassium sulphate K_2SO_4

Solution

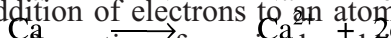
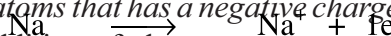
Atomic mass of K	=	39 amu
Atomic mass of S	=	32 amu
Atomic mass of O	=	16 amu
Formula unit	=	K_2SO_4
Formula mass of K_2SO_4	=	$2(39) + 1(32) + 4(16)$
	=	$78 + 32 + 64$
	=	174 amu

**Test yourself****1.3 CHEMISTRY****1.3.1 Ions (Cations and Anions), Molecular Ions and Free Radicals**

- What is the relationship between empirical formula and formula unit?
- How can you differentiate between molecular formula and empirical formula?
- Identify the following formulae as formulas or unit molecular formulae: H_2O_2 , CH_4 , $C_6H_{12}O_6$, $C_{12}H_{22}O_{11}$, $BaCO_3$, KBr
- What is empirical formula of acetic acid (CH_3COOH)? Find out its molecular mass.
- Calculate the formula masses of: Na_2SO_4 , $ZnSO_4$ and $CuCO_3$.

Ion is an atom or group of atoms having a charge on it. The charge may be positive or negative. There are two types of ions i.e. cations and anions. *An atom or group of atoms having positive charge on it is called cation.* The cations are formed when atoms lose electrons from their outermost shells. For example, Na^+ , K^+ are cations. The following equations show the formation of cations from atoms.

Atoms Cations



An atom or a group of atoms that has a negative charge on it, is called anion. Anion is formed by the gain or addition of electrons to an atom. For example, Cl^- and O^{2-} . Following examples show the formation of an anion by addition of electrons to an atom.

Table 1.7 Difference between Atoms and Ions

Atoms		Anions
Cl	+ $1e^-$	$\longrightarrow Cl^-$
O	+ $2e^-$	$\longrightarrow O^{2-}$

	Atom	Ion
i.	It is the smallest particle of an element.	It is the smallest unit of an ionic compound.
ii.	It can or cannot exist independently and can take part in a chemical reaction.	It cannot exist independently and is surrounded by oppositely charged ions.
iii.	It is electrically neutral. <i>When a molecule loses or gains an electron, it forms a molecular ion.</i>	It has a net charge (either negative or positive) on it.

1.3.1.1 Molecular Ion

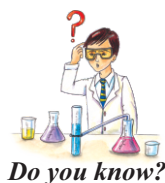
molecular ion or radical is a species having positive or negative charge on it. Like other ions they can be cationic molecular ions (if they carry positive charge) or anionic molecular ions (if they carry negative charge). Cationic molecular ions are more abundant than anionic molecular ions. For example, CH_4^+ , He^+ , N_2^+ . When gases are bombarded with high energy electrons in a discharge tube, they ionize to give molecular ions. Table 1.8 shows some differences between molecule and molecular ion.

Table 1.8 Difference between Molecule and Molecular Ion

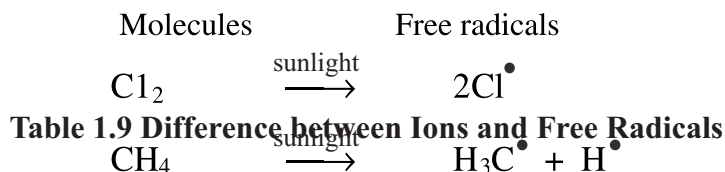
	Molecule	Molecular Ion
i.	It is the smallest particle of an element or compound which can exist independently and shows all the properties of that compound.	It is formed by gain or loss of electrons by a molecule.
ii.	It is always neutral.	It can have negative or positive charge
iii.	It is formed by the combination of atoms	It is formed by the ionization of a molecule.
iv.	It is a stable unit.	It is a reactive specie.

1.3.1.2 Free Radicals

Free radicals are atoms or group of atoms possessing odd number of (unpaired) electrons. It is represented by putting a dot over the symbol of an element e.g. H^\bullet , Cl^\bullet , $\text{H}_3\text{C}^\bullet$. Free radicals are generated by the homolytic (equal) breakage of the bond between two atoms when they absorb heat or light energy. A free radical is extremely reactive species as it has the tendency to complete its octet. Table 1.9 shows some of the differences between ions and free radicals.



Most of the universe exists in the form of plasma, the fourth state of matter. Both the cationic and anionic molecular ions are present in it.



	Ions	Free Radicals
i.	These are the atoms which bear some charge.	These are the atoms that have odd number of electrons.
ii.	They exist in solution or in crystal lattice.	They can exist in solutions as well in air.
iii.	Their formation is not affected by the presence of light.	They may form in the presence of light.

1.3.3. Types of Molecules

A molecule is formed by the chemical combinations of atoms. It is the smallest unit of a substance. It shows all the properties of the substance and can exist independently. There are different types of molecules depending upon the number and types of atoms combining. A few types are discussed here.

A molecule consisting of only one atom is called **monoatomic molecule**. For example, the inert gases helium, neon and argon all exist independently in atomic form and they are called monoatomic molecules.

If a molecule consists of two atoms, it is called **diatomic molecule**. For example: hydrogen (H_2), oxygen (O_2), chlorine (Cl_2) and hydrogen chloride (HCl).

If it consists of three atoms, it is called **triatomic molecule**. For example: H_2O and CO_2 . If a molecule consists of many atoms, it is called polyatomic. For example: methane (CH_4), sulphuric acid (H_2SO_4) and glucose ($\text{C}_6\text{H}_{12}\text{O}_6$).

A Molecule containing same type of atoms, is called **homoatomic molecule**. For example: hydrogen (H_2), ozone (O_3), sulphur (S_8) and phosphorus (P_4) are the examples of molecules formed by the same type of atoms. When a molecule consists of different kinds of atoms, it is called **heteroatomic molecule**. For example: CO_2 , H_2O and NH_3 .

mass of a substance. **Avogadro's Number** is a *collection of 6.02×10^{23} particles*. It is represented by symbol ' N_A '. Hence, the 6.02×10^{23} number of atoms, molecules or formula units is called Avogadro's number that is equivalent to one 'mole' of respective substance. In simple words, 6.02×10^{23} particles are equal to one mole as twelve eggs are equal to one dozen.

To understand **the relationship between the Avogadro's number and the mole** of a substance let us consider a few examples.

- i. 6.02×10^{23} atoms of carbon are equivalent to one mole of carbon.
- ii. 6.02×10^{23} molecules of H_2O are equivalent to one mole of water.
- iii. 6.02×10^{23} formula units of NaCl are equivalent to one mole of sodium chloride.

Thus, 6.02×10^{23} atoms of elements or 6.02×10^{23} molecules of molecular substance or 6.02×10^{23} formula units of ionic compounds are equivalent to 1 mole.

For further explanation about number of atoms in molecular compounds or number of ions in ionic compounds let us discuss two examples:

- i. One molecule of water is made up of 2 atoms of hydrogen and 1 atom of oxygen, hence $2 \times 6.02 \times 10^{23}$ atoms of hydrogen and 6.02×10^{23} atoms of oxygen constitute one mole of water.
- ii. One formula unit of sodium chloride consists of one sodium ion and one chloride ion. So there are 6.02×10^{23} number of Na ions and 6.02×10^{23} Cl ions in one mole of sodium chloride. Thus, the total number of ions in 1 mole of NaCl is 12.04×10^{23} or 1.204×10^{24} .

1.5.2 Mole (Chemist secret unit)

A **mole** is defined as the amount(mass) of a substance that contains 6.02×10^{23} number of particles (atoms, molecules or formula units). It establishes a link between mass of a substance and number of particles as shown in summary of molar calculations. It is abbreviated as 'mol'.

You know that a substance may be an element or compound (molecular or ionic). Mass of a substance is either one of the following: atomic mass, molecular mass or formula mass. These masses are expressed in atomic mass units (amu). But when these masses are expressed in grams, they are called as molar masses.

Scientists have agreed that Avogadro's number of particles are present in one molar mass of a substance. Thus, **quantitative definition of mole** is the atomic mass, molecular mass or formula mass of a substance expressed in grams is called mole.



Amedeo Avogadro (1776-1856) was an Italian scholar. He is famous for molecular theory commonly known as Avogadro's law. In tribute to him, the number of particles (atoms, molecules, ions) in mole of a substance 6.02×10^{23} is known as the Avogadro's constant.

For example:

Atomic mass of carbon expressed as 12 g = 1 mol of carbon

Molecular mass of H_2O expressed as 18 g = 1 mol of water

Molecular mass of H_2SO_4 expressed as 98 g = 1 mol of H_2SO_4

Formula mass of NaCl expressed as 58.5 g = 1 mol of NaCl

Thus, **the relationship between mole and mass** can be expressed as:

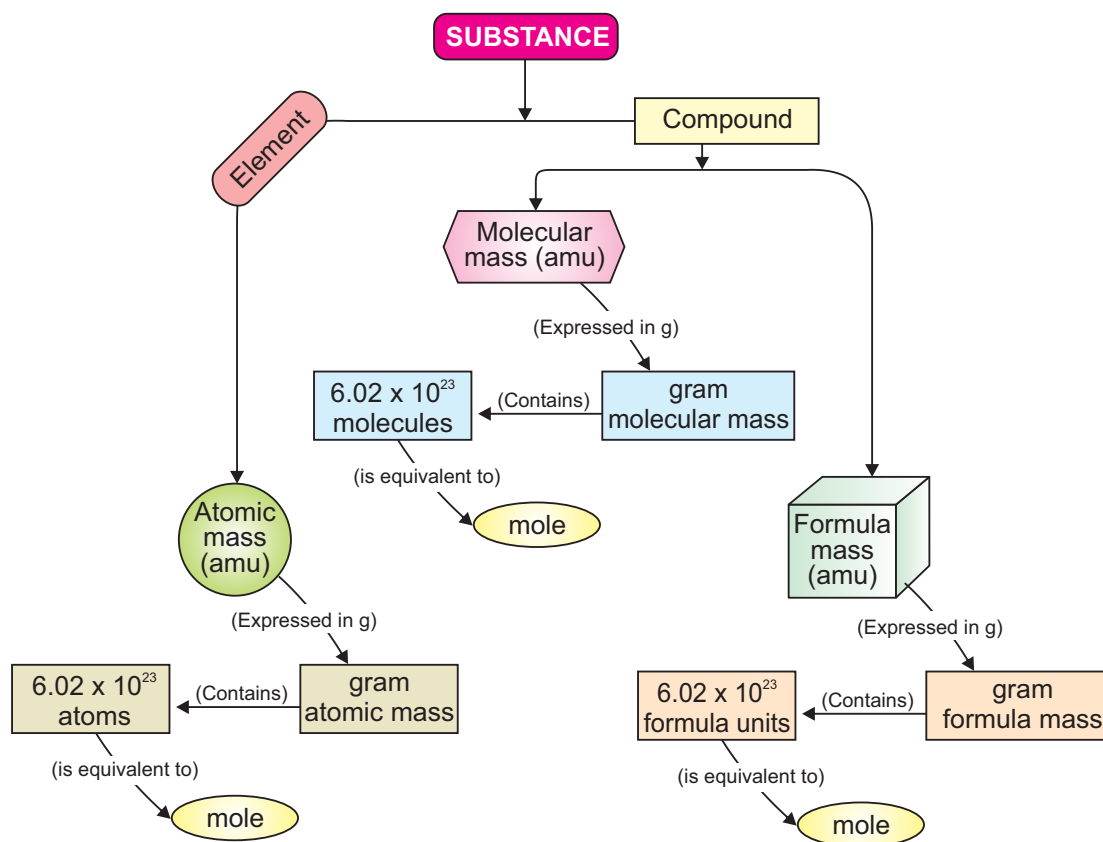
$$\text{Number of moles} = \frac{\text{known mass of a substance}}{\text{molar mass of the substance}}$$

Or,

$$\text{Mass of substance (g)} = \text{number of moles} \times \text{molar mass}$$

A detailed relationship between a substance and a mole through molar mass and number of particles is presented here.

Summary showing a relationship between a substance and a mole.





- i. Which term is used to represent the mass of 1 mole of molecules of a substance?
- ii. How many atoms are present in one gram atomic mass of a substance?
- iii. Explain the relationship between mass and mole of a substance.
- iv. Find out the mass of 3 moles of oxygen atoms.
- v. How many molecules of water are present in half mole of water?

Example 1.4

Calculate the gram molecule (number of moles) in 40 g of H_3PO_4 .

Solution

$$\begin{aligned} \text{Given mass of } H_3PO_4 &= 40 \text{ g} \\ \text{Molecular mass of } H_3PO_4 &= 98 \text{ g mol}^{-1} \end{aligned}$$

Putting these values in equation

$$\begin{aligned} \text{Number of gram molecule (mol)} &= \frac{\text{mass of a substance}}{\text{molar mass of the substance}} \\ &= \frac{40}{98} = 0.408 \end{aligned}$$

Therefore, 40 grams will contain 0.408 gram molecule (mol) of H_3PO_4 .

1.6 CHEMICAL CALCULATIONS

In chemical calculations, we calculate number of moles and number of particles of a given mass of a substance or vice versa. These calculations are based upon mole concept. Let us have a few examples of these calculations.

Calculating the number of moles and number of particles from known mass of a substance.

First calculate the number of moles from given mass by using equation

$$\text{Number of moles} = \frac{\text{known mass of a substance}}{\text{molar mass of the substance}}$$

Then calculate number of particles from the calculated number of moles with the help of following equation:

$$\text{Number of particles} = \text{number of moles} \times 6.02 \times 10^{23}$$

1.6.1 Mole-Mass Calculations

In these calculations, we calculate the number of moles of a substance from the known mass of the substance with the help of following equation:

$$\text{Number of moles} = \frac{\text{known mass of a substance}}{\text{molar mass of the substance}}$$

When we rearrange the equation to calculate mass of a substance from the number of moles of a substance we get,

$$\text{mass of substance (g)} = \text{number of moles} \times \text{molar mass(g)}$$

Example 1.5

You have a piece of coal (carbon) weighing 9.0 gram. Calculate the number of moles of coal in the given mass.

Solution

The mass is converted to the number of moles by the equation:

$$\begin{aligned} \text{Number of moles} &= \frac{\text{known mass of a substance}}{\text{molar mass of the substance}} \\ &= \frac{9.0}{12} = 0.75 \end{aligned}$$

So, 9.0 g of coal is equivalent to 0.75 mol.

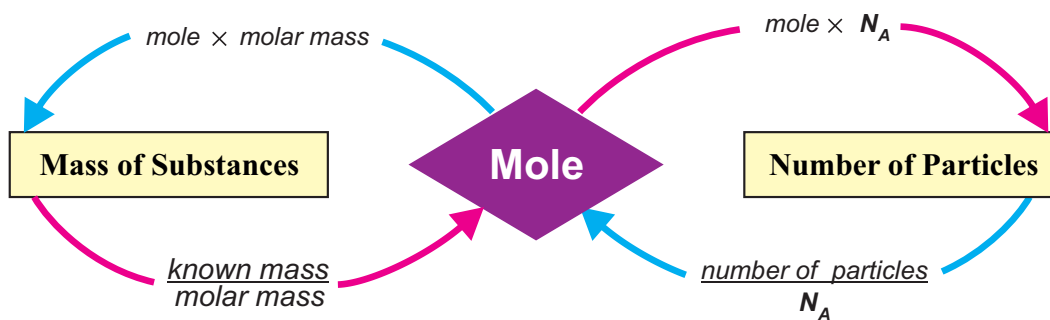
1.6.2 Mole-Particle Calculations

In these calculations, we can calculate the number of *moles* of a substance from the given number of particles. (These particles are the atoms, molecules or formula units).

$$\text{Number of moles} = \frac{\text{given number of particles}}{\text{Avogadro's number}} = \frac{\text{given number of particles}}{6.02 \times 10^{23}}$$

On rearranging above equation we get,

$$\text{Number of particles} = \text{number of moles} \times 6.02 \times 10^{23}$$

Summary of Molar Calculations:**Remember**

- Never calculate the number of particles from mass of the substance or vice versa. Always make calculations through moles.
- For calculations of the number of atoms in molecular compounds and the number of ions in ionic compounds; first calculate the number of molecules or formula units and then calculate the number of atoms or ions.

Example 1.6

Calculate the number of moles, number of molecules and number of atoms present in 6 grams of water.

Solution

$$\begin{aligned}
 \text{The known mass of water} &= 6 \text{ g} \\
 \text{Molar mass of H}_2\text{O} &= 18 \text{ g} \\
 \text{Number of moles of water} &= \frac{\text{mass of a substance}}{\text{molar mass of the substance}} \\
 &= 6/18 \\
 &= 0.33 \text{ mol} \\
 \text{Number of molecules} &= \text{number of moles} \times \text{Avogadro's number} \\
 &= 0.33 \times 6.02 \times 10^{23} \\
 &= 1.98 \times 10^{23} \text{ molecules}
 \end{aligned}$$

The number of molecules contained in 6 grams of water are 1.98×10^{23}

As we know 1 molecule of water consists of 3 atoms, therefore:

$$\begin{aligned}
 \text{Number of atoms} &= 3 \times 1.98 \times 10^{23} \\
 &= 5.94 \times 10^{23}
 \end{aligned}$$

Example 1.7

There are 3.01×10^{23} molecules of CO_2 present in a container. Calculate the number of moles and its mass.

Solution

We can calculate the number of molecules of CO_2 by putting the values in equation

$$\begin{aligned}
 \text{Number of moles of CO}_2 &= \frac{\text{known number of molecules}}{\text{Avogadro's number}} \\
 \text{Number of moles of CO}_2 &= \frac{3.01 \times 10^{23}}{6.02 \times 10^{23}} \\
 &= 0.5 \text{ mol}
 \end{aligned}$$

Then by putting this value in this equation we get

$$\begin{aligned}
 \text{Mass of substance} &= \text{number of moles} \times \text{molar mass (g)} \\
 \text{Mass of CO}_2 &= 0.5 \times 44 \\
 &= 22\text{g}
 \end{aligned}$$



- i. How many atoms of sodium are present in 3 moles of sodium and what is the mass of it?
- ii. How many atoms are in 1 amu and 1 g of hydrogen (H)?
- iii. How many atoms are present in 16 g of O and 8g of S?
- iv. Is the mass of 1 mole of O and 1 mole of S same?
- v. What do you mean by 1 atom of C and 1 gram atom of C?
- vi. If 16 g of oxygen contains 1 mole of oxygen atoms calculate the mass of one atom of oxygen in grams.
- vii. How many times is 1 mole of oxygen atom heavier than 1 mole of hydrogen atom?
- viii. Why does 10 g nitrogen gas contain the same number of molecules as 10 g of carbon monoxide?



THE MOLECULARITY OF THE PHYSICAL WORLD.

The nature of the physical world as perceived through men's senses has been investigated in depth. The biggest lesson we learnt in 20th century is that Chemistry has become central science. It leads to the discovery of every chemical reaction in any living and non-living thing based on formation of "molecules". A reaction in the smallest living organism or in the most developed species like man, always takes place through the process of molecule formation. Hence it provides basis of "molecularity" of the physical world.

CORPUSCULAR NATURE OF MATTER.

In 1924 de Broglie put forward the theory of dual nature of matter i.e. matter has both the properties of particles as well as waves. He explained the background of two ideas. He advocated that these two systems could not remain detached from each other. By mathematical evidences, he proved that every moving object is attached with waves and every wave has corpuscular nature as well. It formulated a basis to understand corpuscular nature of matter.

THE WORKS OF DIFFERENT SCIENTISTS AT THE SAME TIME HANDICAP OR PROMOTE THE GROWTH OF SCIENCE.

Over the course of human history, people have developed many interconnected and validated ideas about the physical, biological, psychological and social worlds. Those ideas have enabled successive generations to achieve an increasingly comprehensive and reliable understanding of the human species and its environment. The means used to develop these ideas are particular ways of observing, thinking, experimenting and validating. These ways represent a fundamental aspect of the nature of science and reflect how science tends to differ from other modes of knowing. It is the union of science, mathematics and technology that forms the scientific endeavor and that makes it so successful. Although, each of these human enterprises has a character and history of its own, each is dependent on and reinforces the others.

MOLE - A QUANTITY

A computer counting with a speed of 10 million atoms a second would take 2 billion years to count one mole of atoms.

If one mole of marbles were spread over the surface of the Earth, our planet would be covered by a 3 miles thick layer of marbles.

A glass of water, which contains about 10 moles of water, has more water molecules than the grains of sand in the Sahara desert.

Key Points

- Chemistry is study of composition and properties of matter. It has different branches.
- Substances are classified into elements and compounds.
- Elements consist of only one type of atoms.
- Compounds are formed by chemical combination of atoms of the elements in a fixed ratio.
- Mixtures are formed by mixing up elements or compounds in any ratio. They are classified as homogeneous and heterogeneous mixtures.

- Each atom of an element has a specific atomic number (Z) and a mass number or atomic mass (A).
- Atomic mass of an atom is measured relative to a standard mass of C-12.
- Relative atomic mass of an element is the mass of an element compared with 1/12 mass of an atom of C-12 isotope.
- Atomic mass unit is 1/12 of the mass of one atom of C-12, $1 \text{ amu} = 1.66 \times 10^{-24} \text{ g}$
- Empirical formula is the simplest type of chemical formula, which shows the relative number of atoms of each element in a compound.
- Molecular formula gives the actual number of atoms of each element in a molecule.
- Formula mass is the sum of atomic masses of all the atoms in one formula unit of a substance.
- An atom or group of atoms having a charge on it is called an ion. If it has positive charge it is called a cation and if it has negative charge it is called an anion.
- There are different types of molecules: monoatomic, polyatomic, homoatomic and heteroatomic.
- The number of particles in one mole of a substance is called Avogadro's number. The value of this number is 6.02×10^{23} It is represented as N_A .
- The amount of a substance having 6.02×10^{23} particles is called a mole. The quantitative definition of mole is atomic mass, molecular mass or formula mass expressed in grams.

EXERCISE

Multiple Choice Questions

Put a (✓) on the correct answer

1. **Industrial chemistry deals with the manufacturing of compounds:**
 (a) in the laboratory (b) on micro scale
 (c) on commercial scale (d) on economic scale
2. **Which one of the following compounds can be separated by physical means?**
 (a) mixture (b) element (c) compound (d) radical
3. **The most abundant element occurring in the oceans is:**
 (a) oxygen (b) hydrogen (c) nitrogen (d) silicon
4. **Which one of the following elements is found in most abundance in the Earth's crust?**
 (a) oxygen (b) aluminium (c) silicon (d) iron
5. **The third abundant gas found in the Earth's atmosphere is:**
 (a) carbon monoxide (b) oxygen. (c) nitrogen (d) argon
6. **One amu (atomic mass unit) is equivalent to:**
 (a) $1.66 \times 10^{-24} \text{ mg}$ (b) $1.66 \times 10^{-24} \text{ g}$
 (c) $1.66 \times 10^{-24} \text{ kg}$ (d) $1.66 \times 10^{-23} \text{ g}$

7. Which one of the following molecule is not tri-atomic?
(a) H_2 (b) O_3 (c) H_2O (d) CO_2
8. The mass of one molecule of water is:
(a) 18 amu (b) 18 g (c) 18 mg (d) 18 kg
9. The molar mass of H_2SO_4 is:
(a) 98g (b) 98 amu (c) 9.8 g (d) 9.8 amu
10. Which one of the following is a molecular mass of O_2 in amu?
(a) 32 amu (b) 53.12×10^{-24} amu
(c) 1.92×10^{-25} amu (d) 192.64×10^{-25} amu
11. How many number of moles are equivalent to 8 grams of CO_2 ?
(a) 0.15 (b) 0.18 (c) 0.21 (d) 0.24
12. In which one of the following pairs has the same number of ions?
(a) 1 mole of $NaCl$ and 1 mole of $MgCl_2$
(b) $1/2$ mole of $NaCl$ and $1/2$ mole of $MgCl_2$
(c) $1/2$ mole of $NaCl$ and $1/3$ mole of $MgCl_2$
(d) $1/3$ mole of $NaCl$ and $1/2$ mole of $MgCl_2$
- Which one of the following pairs has the same mass?
13. (a) 1 mole of CO and 1 mole of N_2
(b) 1 mole of CO and 1 mole of CO_2
(c) 1 mole of O_2 and 1 mole of N_2
(d) 1 mole of O_2 and 1 mole of CO_2

Short answer questions.

1. Define industrial chemistry and analytical chemistry.
2. How can you differentiate between organic and inorganic chemistry?
3. Give the scope of biochemistry.
4. How does homogeneous mixture differ from heterogeneous mixture?
5. What is the relative atomic mass? How is it related to gram?
6. Define empirical formula with an example.
7. State three reasons why do you think air is a mixture and water a compound?
8. Explain why are hydrogen and oxygen considered elements whereas water as a compound.
9. What is the significance of the symbol of an element?
10. State the reasons: soft drink is a mixture and water is a compound.
11. Classify the following into element, compound or mixture:
i. He and H_2 ii. CO and Co iii. Water and milk
iv. Gold and brass v. Iron and steel
12. Define atomic mass unit. Why is it needed?

13. State the nature and name of the substance formed by combining the following:
- i. Zinc + Copper ii. Water + Sugar
iii. Aluminium + Sulphur iv. Iron + Chromium + Nickel
14. Differentiate between molecular mass and formula mass, which of the followings have molecular formula?
 H_2O , $NaCl$, KI , H_2SO_4
15. Which one has more atoms: 10 g of Al or 10 g of Fe?
16. Which one has more molecules: 9 g of water or 9 g of sugar ($C_{12}H_{22}O_{11}$)?
17. Which one has more formula units: 1 g of $NaCl$ or 1 g of KCl ?
18. Differentiate between homoatomic and heteroatomic molecules with examples.
19. In which one of the followings the number of hydrogen atoms is more? 2 moles of HCl or 1 mole of NH_3 (Hint: 1 mole of a substance contains as much number of moles of atoms as are in 1 molecule of a substance)

Long Answer Questions.

1. Define element and classify the elements with examples.
2. List five characteristics by which compounds can be distinguished from mixtures.
3. Differentiate between the following with examples:
- i. Molecule and gram molecule ii. Atom and gram atom
iii. Molecular mass and molar mass iv. Chemical formula and
gram formula
4. Mole is SI unit for the amount of a substance. Define it with examples?

Numericals

1. Sulphuric acid is the king of chemicals. If you need 5 moles of sulphuric acid for a reaction, how many grams of it will you weigh?
2. Calcium carbonate is insoluble in water. If you have 40 g of it; how many Ca^{2+} and CO_3^{2-} ions are present in it?
3. If you have 6.02×10^{23} ions of aluminium; how many sulphate ions will be required to prepare $Al_2(SO_4)_3$?
4. Calculate the number of molecules in the following compounds:
a. 16 g of H_2CO_3 b. 20 g of HNO_3 c. 30 g of $C_6H_{12}O_6$
5. Calculate the number of ions in the following compounds:
a. 10 g of $AlCl_3$ b. 30 g of $BaCl_2$ c. 58 g of $H_2SO_{4(aq)}$
6. What will be the mass of 2.05×10^{16} molecules of H_2SO_4
7. How many atoms are required to prepare 60 g of HNO_3 ?
8. How many ions of Na^+ and Cl^- will be present in 30 g of $NaCl$?
9. How many molecules of HCl will be required to have 10 grams of it?
10. How many grams of Mg will have the same number of atoms as 6 grams of C have?