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ATOMS, MOLECULES AND CHEMICAL ARITHMETIC

Chemistry is the study of matter and the changes it undergoes. Chemistry is often called the central science, because a basic knowledge of chemistry is essential for the study of biology, physics, geology, ecology, and many other subjects.

Although chemistry is an ancient science, its modern foundation was laid in the nineteenth century, when intellectual and technological advances enabled scientists to break down substances into ever smaller components and consequently to explain many of their physical and chemical characteristics.

Chemistry plays a pivotal role in many areas of science and technology e.g. in health, medicine, energy and environment, food, agriculture and new materails.

As you are aware, atoms and molecules are so small that we cannot see them with our naked eyes or even with the help of a microscope. Any sample of matter which can be studied consists of extremely large number of atoms or molecules. In chemical reactions, atoms or molecules combine with one another in a *definite number ratio.* Therefore, it would be pertinent if we could specify the total number of atoms or molecules in a given sample of a substance. We use many *number units* in our daily life. For example, we express the number of bananas or eggs in terms of 'dozen'. In chemistry we use a number unit called **mole** which is very large.

With the help of **mole concept** it is possible to take a desired number of atoms/ molecules by weighing. Now, in order to study chemical compounds and reactions in the laboratory, it is necessary to have adequate knowledge of the quantitative relationship among the amounts of the reacting substances that take part and products formed in the chemical reaction. This relationship is knows as stoichiometry. **Stoichiometry** (derived from the Greek *Stoicheion* = element and *metron* = measure) is the term we use to refer to all the quntatitative aspects of chemical compounds and reactions. In the present lesson, you will see how chemical formulae are determined and how chemical equations prove useful in predicting the proper amounts of the reactants that must be mixed to carry out a complete reaction. In other words we can take reactants for a reaction in such a way that none of the reacting substances is in excess. This aspect is very vital in chemistry and has wide application in industries.

OBJECTIVES

After reading this lesson you will be able to :

- explain the scope of chemistry;
- explain the atomic theory of matter;
- state the laws of chemical combinaton;
- explain Dalton's atomic theory;
- define the terms element, atoms and molecules.
- state the need of SI units;
- list base SI units;
- explain the relationship between mass and number of particles;
- define Avogadro's constant and state its significance;
- calculate the molar mass of different elements and compounds;
- define molar volume of gases at STP.
- define empirical and molecular formulae;
- differentiate between empirical and molecular formulae;
- calculate precentage by mass of an element in a compound and also work out empirical formula from the percentage composition;
- establish relationship between mole, mass and volume;
- calculate the amount of substances consumed or formed in a chemical reaction using a balanced equation and mole concept, and
- explain the role of limiting reagent in limiting the amount of the products formed.

1.1 SCOPE OF CHEMISTRY

Chemistry plays an important role in all aspects of our life. Let us discuss role of chemistry in some such areas.

1.1.1 Health and Medicine

Three major advances in this century have enabled us to prevent and treat diseases. Public health measures establishing sanitation systems to protect vast numbers of people from infectious diseases; surgery with anesthesia, enabling physicians to cure potentially fatal conditions, such as an inflamed appendix; and the introduction of vaccines and antibiotics that made it possible to prevent diseases spread by microbes. Gene therapy promises to be the fourth revolution in medicine. (A gene is the basic unit of inheritance.) Several thousand known conditions, including cystic fibrosis and hemophilia, are carried by inborn damage to a single gene. Many other ailments, such as cancer, heart disease, AIDS, and arthritis, result to an extent from impairment of one or more genes involved in the body's defenses. In gene therapy, a selected healthy gene is delivered to a patient's cell to cure or ease such disorders. To carry out such a procedure, a doctor must have a sound knowledge of the chemical properties of the molecular components involved.

Chemists in the pharmaceutical industry are researching potent drugs with few or no side effects to treat cancer, AIDS, and many other diseases as well as drugs to increase the number of successful organ transplants. On a broader scale, improved understanding of the mechanism of ageing will lead to a longer and healthier lifespan for the world's population.

11.2 Energy and the Environment

Energy is a by-product of many chemical processes, and as the demand for energy continues to increase, both in technologically advanced countries like the United States and in developing ones like India. Chemists are actively trying to find new energy sources. Currently the major sources of energy are fossil fuels (coal, petroleum, and natural gas). The estimated reserves of these fuels will last us another 50-100 years at the present rate of consumption, so it is urgent that we find alternatives.

Solar energy promises to be a viable source of energy for the future. Every year earth's surface receives about 10 times as much energy from sunlight as is contained in all of the known reserves of coal, oil, natural gas, and uranium combined. But much of this energy is "wasted" because it is reflected back into space. For the past thirty years, intense research efforts have shown that solar energy can be harnessed effectively in two ways. One is the conversion of sunlight directly to electricity using devices called *photovoltaic cells.* The other is to use sunlight to obtain hydrogen from water. The hydrogen can then be fed into *a fuel cell* to generate electricity. Although our understanding of the scientific process of converting solar energy to electricity has advanced, the technology has not yet improved to the point where we can produce electricity on a large scale at an economically acceptable cost. By 2050, however, it has been predicted that solar energy will supply over 50 percent of our power needs.

Another potential source of energy is nuclear fission, but because of environmental concerns about the radioactive wastes from fission processes, the future of the nuclear industry is uncertain. Chemists can help to devise better ways to dispose of nuclear waste. Nuclear fusion, the process that occurs in the sun and other stars, generates huge amounts of energy without producing much dangerous radioactive waste. In another 50 years, nuclear fusion will likely be a significant source of energy.

Energy production and energy utilization are closely tied to the quality of our environment. A major disadvantage of burning fossil fuels is that they give off carbon dioxide, which is *a greenhouse gas* (that is, it promotes the heating of Earth's atmosphere), along with sulfur dioxide and nitrogen oxides, which result in acid rain and smog. Harnessing solar energy has no such detrimental effects on the environment. By using fuel-efficient automobiles and more effective catalytic converters, we should be able to drastically reduce harmful auto emissions and improve the air quality in areas with heavy traffic. In addition, electric cars, powered by durable, long-lasting batteries, should be more prevalent in the next century, and their use will help to minimize air pollution.

1.1.3 Materials and Technology

Chemical research and development in the twentieth century have provided us with new materials that have profoundly improved the quality of our lives and helped to advance technology in countless ways. A few examples are polymers (including rubber and nylon), ceramics (such as cookware), liquid crystals (like those in electronic displays), adhesives, and coatings (for example, latex paint).

What is in store for the near future? One likely possibility is room-temperature *superconductors.* Electricity is carried by copper cables, which are not perfect conductors. Consequently, about 20 percent of electrical energy is lost in the form of heat between the power station and our homes. This is a tremendous waste. Superconductors are materials that have no electrical resistance and can therefore conduct electricity with no energy loss.

1.1.4 Food and Agriculture

How can the world's rapidly increasing population be fed? In poor countries, agricultural activities occupy about 80 percent of the workforce and half of an average family budget is spent on foodstuffs. This is a tremendous drain on a nation's resources. The factors that affect agricultural production are the richness of the soil, insects and diseases that damage crops, and weeds that compete for nutrients. Besides irrigation, farmers rely on fertilizers and pesticides to increase crop yield.

1.2 PARTICULATE NATURE OF MATTER

Chemistry deals with study of structure and composition of matter. Since ancient time people have been wondering about nature of matter. Suppose we take a piece of rock and start breaking it into smaller and smaller particles can this process go on far ever resulting in smaller and smaller particles or would it come to stop when such particles are formed which can no longer to broken into still smaller particles? Many people including Greek philosophers Plato and Aristotle believed that matter is continuous and the process of subdivision of matter can go on.

On the other hand, many people believed that the process of subdivision of mater can be repeated only a limited nuimber of times till such particles are obtained which cannot be further subdivided. They believed that mattr is composed of large number of very tiny particles and thus has particle naturew. The smallest indivisible particles of matter were given the name 'atom' from the Greek word "atoms" meaning 'indivisible'. It is generally agreed that the Greek philosopher Leucippus and his student Democritus were the first to propose this idea, about 440 B.C.. However, Maharshi Kanad had propounded the atomic concept of matter earlier (500 BC) and had named the smallest particle of matter as "PARMANU".

1.3 LAWS OF CHEMICAL COMBINATIONS

There was tremendous progress in Chemical Sciences after 18th century. It arose out of an interest in the nature of heat and the way things burn. Major progress was made through the careful use of *chemical balance* to determine the change in mass that occurs in chemical reactions. The great French Chemist Antoine Lavoisier used the balance to study chemical reactions. He heated mercury in a sealed flask that contained air. After several days, a red substance mercury (II) oxide was produced. The gas remaining in the flask was reduced in mass. The remaining gas was neither able to support combustion nor life. The remaining gas in the flask was identified as nitrogen. The gas which combined with mercury was oxygen. Further he carefully performed the experiment by taking a weighed quantity of mercury (II) oxide. After strong heating, he found that mercury (II) oxide, red in colour, was decomposed into mercury and oxygen. He weighed both mercury and oxygen and found that their combined mass was equal to that of the mercury (II) oxide taken. Lavoisier finally came to the conclusion that ***in every chemical reaction, total masses of all the reactants is equal to the masses of all the products***. This law is known as the **law of conservation of mass.**

There was rapid progress in science after chemists began accurate determination of masses of reactants and products. French chemist Claude Berthollet and Joseph Proust worked on the ratio (by mass) of two elements which combine to form a compound. Through a careful work, Proust demonstrated the fundamental law of definite or constant proportions in 1808. **In a given chemical compound, the proportions by mass of the elements that compose it are fixed, independent of the origin of the compound or its mode of preparation.**

In pure water, for instance, the ratio of mass of hydrogen to the mass of oxygen is always 1:8 irrespective of the source of water. In other words, pure water contains 11.11% of hydrogen and 88.89% of oxygen by mass whether water is obtained from well, river or from a pond. Thus, if 9.0 g of water are decomposed, 1.0 g of hydrogen and 8.0 g of oxygen are always obtained. Furthermore, if 3.0 g of hydrogen are mixed with 8.0 g of oxygen and the mixture is ignited, 9.0 g of water are formed and 2.0 g of hydrogen remains unreacted. Similarly sodium chloride contains 60.66% of chlorine and 39.34% of sodium by mass whether we obtained it from salt mines or by crytallising it from water of ocean or inland salt seas or synthesizing it from its elements sodium and chlorine. Of course, the key word in this sentence is 'pure'. Reproducible experimental results are highlights of scientific thoughts. In fact modern science is based on experimental findings. **Reproducible results indirectly hint for a truth which is hidden**. Scientists always worked for findings this truth and in this manner many theories and laws were discovered. This search for truth plays an important role in the development of science.

The Dalton's atomic theory not only explained the laws of conservations of mass and law of constant proportions but also predicted the new ones. He deduced **the law of multiple proportions** on the basis of his theory. The law states that **when two elements form more than one compound, the masses of one element in these compound for a fixed mass of the other element are in the ratio of small whole numbers**. For example, carbon and oxygen form two compounds: carbon monoxide and carbon dioxide. Carbon monoxide contains 1.3321 g of oxygen for each 1.0000 g of carbon, whereas carbon dioxide contains 2.6642 g of oxygen for 1.0000 g of carbon. In other words, carbon dioxide contains twice the mass of oxygen as is contained in carbon monoxide

(2.6642 g = 2 × 1.3321 g) for a given mass of carbon. Atomic theory explains this by saying that carbon dioxide contains twice as many oxygen atoms for a given number of carbon atoms as does carbon monoxide. The deduction of *law of multiple proportions* from atomic theory was important in convincing chemists of the validity of the theory.

1.4 DALTON'S ATOMIC THEORY

As we learnt earlier, Lavosier laid the experimental foundation of modern chemistry. But the British chemist John Dalton (1766–1844) provided the basic theory; all matter – whether element, compound, or mixture –is composed of small particles called atoms. The postulates, or basic assumptions of Dalton's theory are presented below in this section.

1.4.1 Postulates of Dalton's Atomic Theory

The English scientist John Dalton was by no means the first person to propose the existence of atoms, as we have seen in the previous section, such ideas date back to classical times. Dalton's major contribution was to arrange those ideas in proper order and give evidence for the existence of atoms. He showed that the mass relationship expressed by Lavoisier and Proust (in the form of law of conservation of mass and law of constant proportions) could be interpreted most suitably by postulating the existence of atoms of the various elements.

In 1803, Dalton published a new system of chemical philosophy in which the following statements comprise the atomic theory of matter:

1. Matter consists of indivisible atoms.
2. All the atoms of a given chemical element are identical in mass and in all other properties.
3. Different chemical elements have different kinds of atoms and in particular such atoms have different masses.
4. Atoms are indestructible and retain their identity in chemical reactions.
5. The formation of a compound from its elements occurs through the combination of atoms of unlike elements in small whole number ratio.

Dalton's fourth postulate is clearly related to the law of conservation of mass. Every atom of an element has a definite mass. Also in a chemical reaction there is rearrangement of atoms. Therefore after the reaction, mass of the product should remain the same. The fifth postulate is an attempt to explain the law of definite proportions. A compound is a type of matter containing the atoms of two or more elements in small whole number ratio. Because the atoms have definite mass, the compound must have the elements in definite proportions by mass.

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1.4.2 What is an Atom?

As you have just seen in the previous section that an atom is the smallest particle of an element that retains its (elements) chemical properties. An atom of one element is different in size and mass from the atoms of the other elements. These atoms were considered 'indivisible' by Indian and Greek 'Philosophers' in the beginning and the name 'atom' was given as mentioned earlier. Today, we know that atoms are not indivisible. They can be broken down into still smaller particles although they lose their chemical identity in this process. But inspite of all these developments atom still remains a **building block** of matter.

1.4.3 Molecules

*A **molecule** is* an *aggregate of at least two atoms in a definite arrangement held together by chemical forces* (also called *chemical bonds).* It is smallest particle of matter, an element or a compound, which can exist independently. A molecule may contain atoms of the same element or atoms of two or more elements joined in a fixed ratio, in accordance with the law of definite proportions stated. Thus, a molecule is not necessarily a compound, which, by definition, is made up of two or more elements. Hydrogen gas, for example, is a pure element, but it consists of molecules made up of two H atoms each. Water, on the other hand, is a molecular compound that contains hydrogen and oxygen in a ratio of two H atoms and one O atom. Like atoms, molecules are electrically neutral.

The hydrogen molecule, symbolized as H_2, is called *a* ***diatomic molecule*** because it *contains only two atoms.* Other elements that normally exist as diatomic molecules are nitrogen (N_2) and oxygen (O_2), as well as the Group 17 elements-fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2). Of course, a diatomic molecule can contain atoms of different elements. Examples are hydrogen chloride (HCl) and carbon monoxide (CO).

The vast majority of molecules contain more than two atoms. They can be atoms of the same element, as in ozone (O_3), which is made up of three atoms of oxygen, or they can be combinations of two or more different elements. *Molecules containing more than two atoms* are called ***polyatomic molecules.*** Like ozone, water (H_2O) and ammonia (NH_3) are polyatomic molecules.

1.4.4 Elements

Substances can be either elements or compounds. An ***element*** *is a substance that cannot be separated into simpler substances by chemical means.* To date, 118 elements have been positively identified. Eighty-three of them occur naturally on Earth. The others have been created by scientists via nuclear processes.

For convenience, chemists use symbols of one or two, letters to represent the elements. The first letter of a symbol is *always* capitalized, but the following letter is not. For example, Co is the symbol for the element cobalt, whereas CO is the formula for the carbon monoxide molecule. Table 1.1 shows the names and symbols of some of the more common elements; a complete list of the elements and their symbols appears inside the front cover of this book. The symbols of some elements are derived from their Latin names for example, Au from *auram* (gold), Fe from *ferrurn.* (iron), and Na from *natrium* (sodium) while most of them come from their English names.

Table 1.1: Some Common Elements and Their Symbols

Name	Symbol	Name	Symbol	Name	Symbol
Aluminium	Al	Fluorine	F	Oxygen	0
Arsenic	As	Gold	Au	Phosphorus	P
Barium	Ba	Hydrogen	H	Platinum	Pt
Bismuth	Bi	Iodine	I	Potassium	K
Bromine	Br	Iron	Fe	Silicon	Si

Calcium	Ca	Lead	Pb	Silver	Ag
Carbon	C	Magnesium	Mg	Sodium	Na
Chlorine	Cl	Manganese	Mn	Sulfur	S
Chromium	Cr	Mercury	Hg	Tin	Sn
Cobalt	Co	Nickel	Ni	Tungsten	W
Copper	Cu	Nitrogen	N	Zinc	Zn

Chemists use *chemical formulas* to *express the composition of molecules and ionic compounds in terms of chemical symbols.* By composition we mean not only the elements present but also the ratios in which the atoms are combined.

INTEXT QUESTIONS 1.1

1. Chemistry plays a vital role in many areas of science and technology. What are those areas?
2. Who proposed the particulate nature of matter?
3. What is law of conservation of mass?
4. What is an atom?
5. What is a molecule?
6. Why is the symbol of sodium Na?
7. How is an element different from a compound?

1.5 SI UNITS (REVISITED)

Measurement is needed in every walk of life. As you know that for every measurement a 'unit' or a 'reference standard' is required. In different countries, different systems of units gradually developed. This created difficulties whenever people of one country had to deal with those of another country. Since scientists had to often use each other's data, they faced a lot of difficulties. For a practical use, data had to be first converted into local units and then only it could be used.

In 1960, the 'General Conference of Weights and Measures', the international authority on units proposed a new system which was based upon the metric system. This system is called the 'International System of Units' which is abbreviated as SI units from its French name, Le Système Internationale d'Unitès. You have learned about SI units in your earlier classes also and know that they are based upon seven base units corresponding to seven base physical quantities. Units needed for various other physical quantities can be derived from these base SI units. The seven base SI units are listed in Table 1.2.

Table 1.2: SI Base Units

Physical Quantity	Name of SI Unit	Symbol for SI unit
Length	Metre	m
Mass	Kilogram	kg
Time	Second	s
Electrical current	Ampere	A
Temperature	Kelvin	K
Amount of substance	Mole	mol
Luminous intensity	Candela	cd

For measuring very large or very small quantities, multiples or sub-multiples of these units are used. Each one of them is denoted by a symbol which is **prefixed** to the symbol of the unit. For example, to measure long distances we use the unit **kilometre** which is a multiple of metre, the base unit of length. Here **kilo** is the prefix used for the multiple 10^3. Its symbol is k which is prefixed to the symbol of metre, m . Thus the symbol of kilometre is km and

$$1 \text{ km} = 1.0 \times 10^3 \text{ m} = 1000 \text{ m}$$

Similarly, for measuring small lengths we use centimetre (cm) and millimetre (mm) where

$$1 \text{ cm} = 1.0 \times 10^{-2} \text{ m} = 0.01 \text{ m}$$

$$1 \text{ mm} = 1.0 \times 10^{-3} \text{ m} = 0.001 \text{ m}$$

Some prefixes used with SI units are listed in Table 1.3.

Table 1.3: Some prefixes used with SI units

Prefix	Symbol	Meaning	Example
Tera	T	10^{12}	1 terametre (Tm) = 1.0×10^{12} m
Giga	G	10^9	1 gigametre (Gm) = 1.0×10^9 m
Mega	M	10^6	1 megametre (Mm) = 1.0×10^6 m
Kilo	k	10^3	1 kilometre (km) = 1.0×10^3 m
Hecta	h	10^2	1 hectametre (hm) = 1.0×10^2 m
Deca	da	10^1	1 decametre (dam) = 1.0×10^1 m
Deci	d	10^{-1}	1 decimetre (dm) = 1.0×10^{-1} m
Centi	c	10^{-2}	1 centimetre (cm) = 1.0×10^{-2} m
Milli	m	10^{-3}	1 millimetre (mm) = 1.0×10^{-3} m
Micro	μ	10^{-6}	1 micrometre (μm) = 1.0×10^{-6} m
Nano	n	10^{-9}	1 nanometre (nm) = 1×10^{-9} m
Pico	p	10^{-12}	1 picometre (pm) = 1×10^{-12}m

Before proceeding further try to answer the following questions:

INTEXT QUESTIONS 1.2

1. Name the SI Unit of mass
2. What symbol will represent 1.0×10^{-6} g ?
3. Name the prefixes used for (i) 10^2 and (ii) 10^{-9}

   (i) ....

   (ii) ....
4. What do the following symbols represent?

   (i) Ms (ii) ms

   (i) ....

   (ii) ....

1.6 RELATIONSHIP BETWEEN MASS AND NUMBER OF PARTICLES

Suppose you want to purchase 500 screws. How, do you think, the shopkeeper would give you the desired quantity? By counting the screws individually? No, he would give the screws by weight because it will take a lot of time to count them. If each screw weighs 0.8 g, he would weigh 400 g screws because it is the mass of 500 screws ($0.8 \times 500 = 400$ g). You will be surprised to note that the Reserve Bank of India gives the desired number of coins by weight and not by counting.This process of *counting by weighing* becomes more and more labour saving as the number of items to be counted becomes large. We can carry out the reverse process also. Suppose we take 5000 very tiny springs (used in watches) and weigh them. If the mass of these springs is found to be 1.5 g, we can conclude that mass of each spring is $1.5 \div 5000 = 3 \times 10^{-4}$ g.

Thus, we see that mass and number of identical objects or particles are inter-related. Since atoms and molecules are extremely tiny particles it is impossible to weigh or count them individually. Therefore we need a relationship between the mass and number of atoms and molecules (particles). Such a relationship is provided by 'mole concept'.

1.7 MOLE – A NUMBER UNIT

Mass of an atom or a molecule is an important property. However, while discussing the quantitative aspects of a chemical reaction, the *number* of reacting atoms or molecules is more significant than their masses.

It is observed experimently that iron and sulphur do not react with each other in a simple mass ratio. When taken in 1:1 ratio by mass (Fe:S), some sulphur is left unreacted and when taken in 2:1 ratio by mass (Fe:S) some iron is left unreacted.

Let us now write the chemical equation of this reaction

$$Fe + S \rightarrow FeS$$

From the above chemical equation, it is clear that 1 atom of iron reacts with 1 atom of sulphur to form 1 molecule of iron (II) sulphide (FeS). It means that if we had taken equal *number* of atoms of iron and sulphur, both of them would have reacted completely. Thus we may conclude that *substances react in a simple ratio by number of atoms or molecules.*

From the above discussion it is clear that the *number* of atoms or molecules of a substance is more relevant than their masses. In order to express their number we need a number unit. One commonly used number unit is 'dozen', which, as you know, means a collection of 12. Other number units that we use are 'score' (20) and 'gross'(144 or 12 dozens). These units are useful in dealing with small numbers only. The atoms and molecules are so small that even in the minute sample of any substance, their number is extremely large. For example, a tiny dust particle contains about 10^{16} molecules. In chemistry such large numbers are commonly represented by a unit known as **mole.** Its symbol is 'mol' and it is defined as.

A mole is the amount of a substance that contains as many elementary entities (atoms, molecules or other particles) as there are atoms in exactly 0.012 kg or 12 g of the carbon-12 isotope.

The term mole has been derived from the Latin word *'moles'* which means a 'heap' or a 'pile'. It was first used by the famous chemist Wilhelm Ostwald more than a hundred years ago.

Here you should remember that one mole always contains the **same number of entities**, no matter what the substance is. Thus *mole* is a number unit for dealing with elementary entities such as atoms, molecules, formula units, electrons etc., just as *dozen* is a number unit for dealing with bananas or oranges. In the next section you will learn more about this number.

1.8 AVOGADRO'S CONSTANT

In the previous section we have learned that a mole of a substance is that amount which contains as many elementary entities as there are atoms in exactly 0.012 kilogram or 12 gram of the carbon-12 isotope. This definition gives us a method by which we can find out the amount of a substance (in moles) if we know the number of elementary entities present in it or *vice versa.* Now the question arises how many atoms are there in exactly 12 g of carbon-12. This number is determined experimentally and its currently accepted value is 6.022045×10^{23}. Thus 1 mol = 6.022045×10^{23} entities or particles, or atoms or molecules.

For all practical purposes this number is · rounded off to 6.022×10^{23}.

The basic idea of such a number was first conceived by an Italian scientist Amedeo Avogadro. But, he never determined this number. It was determinned later and is known as *Avogadro's constant* in his honour.

This number was earlier known as *Avogadro's number*. This number alongwith the unit, that is, 6.022×10^{23} mol^{-1} is known as Avogadro constant. It is represented by the symbol N_A. Here you should be clear that mathematically a number does not have a unit. Avogadro's number 6.022×10^{23} will not have any unit but Avogradro's constant will have unit of mol^{-1}. Thus Avogradro's constant, $N_A = 6.022 \times 10^{23}$ mol^{-1}.

Significance of Avogadro's Constant

You know that 0.012 kg or 12 g of carbon –12 contains its *one mole* of carbon atoms. A mole may be defined as the amount of a substance that contains 6.022×10^{23} elementary entities like atoms, molecules or other particles. When we say one mole of carbon –12, we mean 6.022×10^{23} atoms of carbon –12 whose mass is exactly 12 g. This mass is called the *molar mass* of carbon-12. The ***molar mass*** *is defined as the mass (in grams) of 1 mole of a substance*. Similarly, a *mole of any substance* would contain 6.022×10^{23} particles or elementary entities. The nature of elementary entity, however,depends upon the nature of the substance as given below :

S.No.	Type of Substance	Elementary Entity
1.	Elements like Na, K, Cu which exist in atomic form	Atom
2.	Elements like O, N, H, which exist in molecular form (O_2, N_2, H_2)	Molecule
3.	Molecular compounds like NH_3, H_2O, CH_4	Molecule
4.	Ions like Na^+, Cu^{2+}, Ag^+, Cl^-, O^{2-}	Ion
5.	Ionic compounds like NaCl, $NaNO_3$, K_2SO_4	Formula unit

Formula unit of a compound contains as many atoms or ions of different types as is given by its chemical formula. The concept is applicable to all types of compounds. The following examples would clarify the concept.

Notes

Formula	Atoms/ions present in one formula unit
H_2O	Two atoms of H and one atom of O
NH_3	One atom of N and three atoms of H
$NaCl$	One Na^+ ion and one Cl^- ion
$NaNO_3$	One Na^+ ion and one NO_3^- ion
K_2SO_4	Two K^+ ions and one SO_4^{2-} ion
$Ba_3(PO_4)_2$	Three Ba^{2+} ions and two PO_4^{3-} ions

Now, let us take the examples of different types of substances and correlate their amounts and the number of elementary entities in them.

1 mole C $= 6.022 \times 10^{23}$ C atoms

1 mole O_2 $= 6.022 \times 10^{23}$ O_2 molecules

1 mole H_2O $= 6.022 \times 10^{23}$ H_2O molecules

1 mole NaCl $= 6.022 \times 10^{23}$ formula units of NaCl

1 mole Ba^{2+} ions $= 6.022 \times 10^{23}$ Ba^{2+} ions

We may choose to take amounts other than one mole and correlate them with number of particles present with the help of relation :

Number of elementary entities = number of moles × Avogadro's constant

1 mole O_2 $= 1 \times (6.022 \times 10^{23})$ $= 6.022 \times 10^{23}$ molecules of O_2

0.5 mole O_2 $= 0.5 \times (6.022 \times 10^{23})$ $= 3.011 \times 10^{23}$ molecules of O_2

0.1 mole O_2 $= 0.1 \times (6.022 \times 10^{23})$ $= 6.022 \times 10^{22}$ molecules of O_2

INTEXT QUESTIONS 1.3

1. A sample of nitrogen gas consists of 4.22×10^{23} molecules of nitrogen. How many moles of nitrogen gas are there?
2. In a metallic piece of magnesium, 8.46×10^{24} atoms are present. Calculate the amount of magnesium in moles.
3. Calculate the number of Cl_2 molecules and Cl atoms in 0.25 mol of Cl_2 gas.

1.9 MOLE, MASS AND NUMBER RELATIONSHIPS

You know that 1 mol $= 6.022 \times 10^{23}$ elementary entities

and Molar mass = Mass of 1 mole of substance

$= $ Mass of 6.022×10^{23} elementary entities.

As discussed earlier the elementary entity can be an atom, a molecule, an ion or a formula unit. As far as mole – number relationship is concerned it is clear that one mole of any substance would contain 6.022×10^{23} particles (elementary entities). For obtaining the molar mass, i.e., mole-mass relationship we have to use atomic mass scale.

1.9.1 Atomic Mass Unit

By inernational agreement, a unit of mass to specify the atomic and molecular masses has been defined. This unit is called *atomic mass unit* and its symbol is 'amu'. The mass of one C-12 atom, is taken as exactly 12 amu. Thus, C-12 atom serves as the *standard*. The **Atomic mass unit** is defined as a mass exactly equal to the $1/12^{th}$ of the mass of one carbon-12 atom.

$$1 \text{ amu} = \frac{\text{Mass of one C-12 atom}}{12}$$

Atomic mass unit is also called **unified atomic mass unit** whose symbol is 'u'. Another name of atomic mass unit is **dalton** (symbol Da). The latter is mainly used in biological sciences.

1.9.2 Relative Atomic and Molecular Masses

You are aware that atomic mass scale is a *relative scale* with C-12 atom (also written as ^{12}C) chosen as the standard. Its mass is taken as exactly 12. Relative masses of atoms and molecules are the number of times each atom or molecules is heavier than $\frac{1}{12}$th of the mass of one C-12 atom. Often, we deal with elements and compounds containing isotopes of different elements. Therefore, we prefer to use *average* masses of atoms and molecules. Thus

$$\text{Relative atomic mass} = \frac{\text{Average mass of 1 atom of the element}}{\frac{1}{12}\text{th of the mass of one C-12 atom}}$$

$$\text{and Relative molecular mass} = \frac{\text{Average mass of 1 molecule of the substance}}{\frac{1}{12}\text{th of the mass of one C-12 atom}}$$

Experiments show that one O-16 atom is 1.333 times as heavy as one C-12 atom. Thus

$$\text{Relative atomic mass of O-16} = 1.333 \times 12 = 15.996 \simeq 16.0$$

The relative atomic masses of all elements have been determined in a similar manner. Relative molecular masses can also be determined experimentally in a similar manner . In case we know the molecular formula of a molecule, we can calculate its relative molecular mass by adding the relative atomic masses of all its constituent atoms. Let us calculate the relative molecular mass of water, H_2O.

$$\text{Relative molecular mass of water, } H_2O = (2 \times \text{relative atomic mass of H}) + (\text{relative atomic mass of O})$$

$$= (2 \times 1) + (16) = 2 + 16 = 18$$

The relative atomic and molecular masses are just numbers and dimensionless, unit-less quantities.

1.9.3 Atomic, Molecular and Formula Masses

From the definition of atomic mass unit, we can calculate the atomic masses. Let us again take the example of oxygen-16 whose relative atomic mass is 16. By definition:

$$\text{Relative atomic mass of O-16} = 16 = \frac{\text{mass of one O-16 atom}}{\frac{1}{12}\text{th the mass of one C-12 atom}}$$

$$\text{Since 1 amu} = \frac{1}{12}\text{th the mass of one C-12 atom}$$

$$\therefore \quad 16 = \frac{\text{mass of one O-16 atom}}{1 \text{ amu}}$$

$$\text{Mass of one O-16 atom} = 16 \text{ amu}$$

Or $$\text{Atomic mass of O-16} = 16 \text{ amu.}$$

From this example we can see that numerical value of the relative atomic mass and atomic mass is the same. Only, the former has no unit while the latter has the unit *amu*.

Molecular and formula masses can be obtained by adding the atomic or ionic masses of all the constituent atoms or ions of the molecule or formula unit respectively. Let us understand these calculations with the help of following examples.

Example 1.1 : Calculate the molecular mass of ammonia, NH_3.

Solution : One molecule of NH_3 consists of one N atom and three H atoms.

$$\text{Molecular mass of } NH_3 = (\text{Atomic mass of N}) + 3\,(\text{Atomic mass of H})$$

$= [14 + (3 \times 1)]$ amu

$= 17$ amu

Example 1.2 : Calculate the formula mass of sodium chloride (NaCl).

Solution : One formula unit of sodium chloride consists of one Na^+ ion and one Cl^- ion.

Formula mass of NaCl $=$ (Ionic mass of Na^+) + (Ionic mass of Cl^-)

$= 23$ amu + 35.5 amu

$= 58.5$ amu.

You would have noticed in the above example that ionic mass of Na^+ ion has been taken as 23 amu which is the same as the atomic mass of Na atom. Since loss or gain of few electrons does not change the mass significantly, therefore atomic masses are used as ionic masses. Similarly we have taken ionic mass of Cl^- as 35.5 amu which is the same as the atomic mass of Cl^-.

1.9.4 Molar Masses

We know that molar mass is the mass of 1 mol of the substance. Also, 1 mol of any substance is the collection of its 6.022×10^{23} elementary entities. Thus

Molar mass $=$ Mass of 6.022×10^{23} elementary entities.

(i) Molar mass of an element

You know that the relative atomic mass of carbon–12 is 12. A 12g sample of it would contain 6.022×10^{23} atoms. Hence the molar mass of C-12 is 12 g mol^{-1}. For getting the molar masses of other elements we can use their relative atomic masses.

Since the relative atomic mass of oxygen -16 is 16, a 16 g sample of it would contain 6.022×10^{23} oxygen atoms and would constitute its one mole. Thus, the molar mass of O–16 is 16 g mol^{-1}. Relative atomic masses of some common elements have been listed in Table 1.4

Table 1.4 : Relative atomic masses of some elements (upto 1st place of decimal)

Element	Relative Atomic Mass	Element	Relative Atomic Mass
Hydrogen, H	1.0	Phosphorus, P	31.0
Carbon, C	12.0	Sulphur, S	32.1
Nitrogen, N	14.0	Chlorine, Cl	35.5
Oxygen, O	16.0	Potassium, K	39.1
Sodium, Na	23.0	Iron, Fe	55.9

(ii) Molar mass of a molecular substance

The elementary entity in case of a molecular substance is the molecule. Hence, *molar mass of such a substance would be the mass of its 6.022×10^{23} molecules*, which can be obtained from its relative molecular mass or by multiplying the molar mass of each element by the number of its moles present in one mole of the substance and then adding them.

Let us take the example of water, H_2O. Its relative molecular mass is 18. Therefore, 18 g of it would contain 6.022×10^{23} molecules. Hence, its molar mass is 18 g mol^{-1} . Alternately we can calculate it as :

$$\begin{aligned}\text{Molar mass of water, } H_2O &= (2 \times \text{molar mass of H}) + (\text{molar mass of O}) \\ &= (2 \times 1 \text{ g mol}^{-1}) + (16 \text{ g mol}^{-1}) \\ &= 18 \text{ g mol}^{-1}\end{aligned}$$

Table 1.5 lists molecular masses and molar masses of some substances.

Table 1.5 : Molecular masses and molar masses of some substances

Element or Compound	Molecular mass / amu	Molar mass / (g mol^{-1})
O_2	32.0	32.0
P_4	124.0	124.0
S_8	256.8	256.8
H_2O	18.0	18.0
NH_3	17.0	17.0
HCl	36.5	36.5
CH_2Cl_2	85.0	85.0

(iii) Molar masses of ionic compounds

Molar mass *of an ionic compound is the mass of its 6.022×10^{23} formula units.* It can be obtained by adding the molar masses of ions present in the formula unit of the substance. In case of NaCl it is calculated as

$$\begin{aligned}\text{Molar mass of NaCl} &= \text{molar mass of } Na^{+} + \text{molar mass of } Cl^{-} \\ &= (23 \text{ g mol}^{-1}) + (35.5 \text{ g mol}^{-1}) \\ &= 58.5 \text{ g mol}^{-1}\end{aligned}$$

Let us take some more examples of ionic compounds and calculate their molar masses.

Example 1.3 : Calculate the molar mass of

(i) K_2SO_4 (ii) $Ba_3(PO_4)_2$

Solution :

(i) Molar mass of K_2SO_4 = (2 × molar mass of K^+) + (molar mass of SO_4^{2-})

= (2 × molar mass of K^+) + (molar mass of S + 4 × molar mass of O)

= [(2 × 39.1) + (32.1 + 4 × 16)] g mol^{-1}]

= (78.2 + 32.1 + 64) g mol^{-1} = 174.3 g mol^{-1}

(ii) Molar mass of $Ba_3(PO_4)_2$ = (3 × molar mass of Ba^{2+}) + 2 (molar mass of PO_4^{3-})

= (3 × molar mass of Ba^{2+}) + 2 (molar mass of P + 4 × molar mass of O)

= [(3 × 137.3) + 2 (31.0 + 4 × 16.0)] g mol^{-1}

= (411.9 + 190.0) g mol^{-1} = 601.9 g mol^{-1}

Now you have learned about the mole, mass and number relationships for all types of substances. The following examples would illustrate the usefulness of these relationships.

Example 1.4 : Find out the mass of carbon -12 that would contain 1.0×10^{19} carbon-12 atoms.

Solution : Mass of 6.022×10^{23} carbon-12 atoms = 12 g

Mass of 1.0×10^{19} carbon-12 atoms $= \dfrac{12 \times 1 \times 10^{19}}{6.022 \times 10^{23}}$ g

$= 1.99 \times 10^{-4}$ g

Example 1.5 : How many molecules are present in 100 g sample of NH_3?

Solution : Molar mass of NH_3 = (14 + 3) g mol^{-1} = 17 g mol^{-1}

∴ 17 g sample of NH_3 contains 6.022×10^{23} molecules

Therefore, 100 g sample of NH_3 would contain $\dfrac{6.022 \times 10^{23}\ \text{molecule}}{17\text{g}} \times 100\text{g}$

$= 35.42 \times 10^{23}$ molecules

$= 3.542 \times 10^{24}$ molecules

Example 1.6 : Molar mass of O is 16 g mol^{-1}. What is the mass of one atom and one molecule of oxygen?

Solution : Mass of 1 mol or 6.022×10^{23} atoms of O $= 16$ g

$$\therefore \qquad \text{Mass of 1atom of O} = \frac{16\text{g}}{6.022 \times 10^{23}}$$

$$= 2.66 \times 10^{-23}\ \text{g}$$

Since a molecule of oxygen contains two atoms (O_2), its mass $= 2 \times 2.66 \times 10^{-23}$ g $= 5.32 \times 10^{-23}$ g.

INTEXT QUESTIONS 1.4

1. Calculate the molar mass of hydrogen chloride, HCl.
2. Calculate the molar mass of argon atoms, given that the mass of single atom is 6.634×10^{-26} kg.
3. Calculate the mass of 1.0 mol of potassium nitrate, KNO_3 (atomic masses : K = 39 amu; N = 14 amu, O = 16 amu).
4. The formula of sodium phosphate is Na_3PO_4. What is the mass of 0.146 mol of Na_3PO_4? (atomic masses : Na = 23.0 amu, P = 31.0 amu; O = 16.0 amu).

1.10 MASS, MOLAR MASS AND NUMBER OF MOLES

Mass, molar mass and number of moles of a substance are inter-related quantities. We know that :

Molar mass (M) = Mass of one mole of the substance.

Molar mass of water is 18 g mol^{-1}. If we have 18 g of water, we have 1mol of it. Suppose we have 36 g water (18×2), we have 2 mol of it. In general in a sample of water of mass ($n \times 18$) g, the number of moles of water would be n. We may generalize the relation as

$$\text{Number of moles (amount) of a substance} = \frac{\text{mass of the substance}}{\text{molar mass of the substance}}$$

$$n = \frac{m}{M}$$

or

$$m = n \times M$$

These relations are useful in calculations involving moles of substances.

Example 1.7 : In a reaction, 0.5 mol of aluminium is required. Calculate the amount of aluminium required in grams? (atomic mass of Al = 27 amu)

Solution :

$$\text{Molar mass of Al} = 27 \text{ g mol}^{-1}$$

$$\begin{aligned}\text{Required mass} &= \text{no. of moles} \times \text{molar mass}\\ &= (0.5 \text{ mol}) \times (27 \text{ g mol}^{-1})\\ &= 13.5 \text{ g}\end{aligned}$$

1.11 MOLAR VOLUME, V_m

Molar volume *is the volume of one mole of a substance*. It depends upon temperature and pressure. It is related to the density, by the relation.

$$\text{Molar volume} = \frac{\text{Molar mass}}{\text{Density}}$$

In case of gases, we use their volumes at **standard temperature and pressure (STP).** For this purpose **0 ^{0}C** or **273 K** temperature is taken as the **standard temperature** and **1bar** pressure is taken as the **standard pressure**. At STP, the molar volume of an ideal gas is 22.7 litre*. You will study that gases do not behave ideally and therefore their molar volume is not exactly 22.7 L. However, it is very close to 22.7 L and for all practical purposes we take the molar volume of all gases at STP as 22.7 L mol^{-1}.

INTEXT QUESTIONS 1.5

1. How many moles of Cu atoms are present in 3.05 g of copper (Relative atomic mass of Cu = 63.5).
2. A piece of gold has a mass of 12.6 g. How many moles of gold are present in it? (Relative atomic mass of Au = 197)
3. In a combustion reaction of an organic compound, 2.5 mol of CO_2 were produced. What volume would it occupy at STP (273K, 1bar) ?

1.12 MOLCULAR AND EMPIRICAL FORMULAE

In your previous classes, you have studied how to write chemical formula of a sustance. For example, water is represented by H_2O, carbon dioxide is represented

* Earlier 1 atmosphere pressure was taken as the standard pressure and at STP (273K, 1atm) the molar volume of an ideal gas was taken as 22.4 L mol^{-1}. The difference in the value is due to the change in the standard pressure (1bar) which is slightly less than 1atm.

by CO_2, methane is represented by CH_4, dinitrogen penta oxide is represented by N_2O_5, and so on. You are aware, formula for a molecule uses a symbol and subscript number to indicate the number of each kind of atoms present in the molcule (subscript 1 is always omitted). Such a formula is called **molecular formula** as it represents a molecule of a substance. A molecule of water consists of two hydrogen atoms and one oxygen atom. So its molecular formula is written as H_2O. Thus a **molecular formula shows the actual number of atoms of different elements in a molecule of a compound.**

There is another kind of formula, the empirical formul of a compound, which gives only relative number of atoms of different elements. These numbers are expressed as the simplest ratio. For example, empirical formula of glucose, which consists of carbon, hydrogen and oxygen in the ratio of 1:2:1 is CH_2O (empirical formulae are also called simplest formulae). Molecular formula of a substance is always an integral multiple of its empirical formula (i.e. molecular formula = X_n where X is empirical formula and n is an integer). For example molecular formula of glucose is $C_6H_{12}O_6$ which is 6 × its empirical formula. Thus, while empirical formula gives only a ratio of atoms, the molecular formula gives the actual number of atoms of each element in an individual molecule. In some cases the ratio of atoms shown in a molecular formula cannot be reduced to smaller integers. In such cases molecular and empirical formulae are the same, for example, sucrose $C_{12}H_{22}O_{11}$ which is popularly known as cane-sugar. In case of certain elements, a molecule consists of several atoms for example P_4, S_8, etc. In such cases, empirical formula will be symbol of the element only.

As you know, common salt, which is chemically called sodium chloride is represented as NaCl. This salt is ionic in nature and does not exist in molecular form. Therefore, NaCl is its empirical formula which shows that sodium and chlorine atoms are present in NaCl in the ratio of 1:1. Similar is the case with all ionic substanes. KCl, $NaNO_3$, MgO are examples of empirical formulae as these are all ionic compounds. Table 1.6 provides a few more examples.

Table 1.6: Molecular and Empirical Formulae

Substance	Molecular formula	Empirical formula
Ammonia	NH_3	NH_3
Carbon dioxide	CO_2	CO_2
Ethane	C_2H_6	CH_3
Fructose	$C_6H_{12}O_6$	CH_2O
Sulphur	S_8	S
Benzene	C_6H_6	CH
Sodium chloride	—	NaCl
Calcium oxide	—	CaO

1.13 CHEMICAL COMPOSITION AND FORMULAE

How much carbon is present in one kilogram of methane whose molecular formula is CH_4? How much nitrogen is present in one kilogram of ammonia, NH_3? If we have prepared a substance that is made of 58.8% carbon, 28.4% oxygen, 8.28% nitrogen and 6.56% hydrogen, what is its empirical formula? You have studied **atomic masses, formulae,** and the **mole concept.** Can we solve the problem using these basic concepts? The answer is 'yes'. Atomic masses, formulae and the mole concept are the basic tools needed to solve such problems. What is percentage composition? Let us take up this aspect in a little detail and try to understand.

1.13.1 Percentage Composition

If we know the formula of a compound, we can find out how much of each of the elements is present in a given quantity of the compound. Aluminium is obtained from its oxide. Al_2O_3 (which is found as the ore, bauxite). From the formula we can calculate how much aluminium can be obtained, at least in prinicple, from a given amount of aluminium oxide. Calculation is done by making use of the idea of **percentage composition**

Percentage mass of an element in a compound

$$= \frac{\begin{array}{c}\text{mass of element in one molecular formula or}\\ \text{in one empirical formula}\end{array}}{\begin{array}{c}\text{molecular mass or empirical formula}\\ \text{mass of compound}\end{array}} \times 100$$

$$= \frac{\text{Mass of element in 1mol of compound}}{\text{Molar mass of compound}} \times 100$$

Let us calculate percentage composition of aluminium oxide, Al_2O_3

$$\text{Pecentage of aluminium} = \frac{\text{Mass of aluminium in 1 mol } Al_2O_3}{\text{Molar mass of } Al_2O_3} \times 100$$

Molar mass of Al_2O_3= (2×27.0) g + (3×16.0) g = 102.0 g

Since 1 mol of Al_2O_3 contains 2 mol of Al atoms, the mass of Al is 2×27.0 g = 54.0 g Al

$$\text{Percentage of Aluminium} = \frac{54.0\text{ g}}{102.0\text{ g}} \times 100 = 52.9\ \%$$

We can calculate percentage of oxygen in the same way. One mole of Al_2O_3 contains 3 mole of O atoms, that is, 3×16.0 g oxygen therefore

$$\text{Percentage of oxygen} = \frac{3 \times 16.0\text{ g}}{102.0\text{ g}} \times 100 = 47.1\%$$

Example 1.8: Butanoic acid, has the formula $C_4H_8O_2$. What is the elemental analysis of butanoic acid?

Solution : Molecular formula of the butanoic acid is $C_4H_8O_2$.

In one mole of butanoic acid there are 4 mol of carbon atoms, 8 mol of hydrogen atoms and 2 mol of oxygen atoms. Thus, 1 molar mass of butanoic acid will be equal to the sum of 4 × molar mass of carbon atoms, 8 × molar mass of hydrogen atoms, and 2 × molar mass of oxygen atoms.

Molar mass of butanoic acid = $4 \times 12.0 \text{ g} + 8 \times 1.0 \text{ g} + 2 \times 16.0 \text{ g} = 88.0 \text{ g}$

$$\text{Percentage of C by mass} = \frac{48.0 \text{ g}}{88.0 \text{ g}} \times 100 = 54.5\%$$

$$\text{Percentage of H by mass} = \frac{8.0 \text{ g}}{88.0 \text{ g}} \times 100 = 9.1\%$$

$$\text{Percentage of O by mass} = \frac{32.0 \text{ g}}{88.0 \text{ g}} \times 100 = 36.4\%$$

The percentage of O in butanoic acid can also be calculated as follows :

Percentage of O by mass = 100 – (Percentage of C by mass + Percentage of H by mass)

$$= 100 - (54.5 + 9.1) = 36.4\%$$

1.14 DETERMINATION OF EMPIRICAL FORMULAE – FORMULA STOICHIOMETRY

We have just seen that if we know the formula of a compound we can calculate the percentage composition. Now the question arises, can we determine the formula of the compound if we know the percentage composition of a compound. The answer will be 'yes', but this formula will not be molecular formula; instead it would be **empirical formula** as it would give simplest ratio of different atoms present in a compound. Normally we determine the percentage composition of different elements present in an **unknown compound** and determine its formula. Let us take a simple example of water. Water consists of 11.11% hydrogen and 88.89% oxygen by mass. From the data, we can determine empirical formula of water. Now if we assume that we have a 100.00 g sample of water, then the percentage composition tells us that 100.0 g of water contains 11.11 g of hydrogen atoms and 88.89 g of oxygen atoms.

From the atomic mass table, we find that 1 mol of hydrogn atoms has a mass of 1.0g, and 1 mol of oxygen atoms has a mass of 16.0 g. Now we can write **unit conversion factors** so that the mass of hydrogen can be converted to moles of H atoms and the mass of oxygen can be converted to moles of O atoms.

$$\text{Moles of H} = \frac{\text{Mass of H}}{\text{Molar mass of H}} = \frac{11.11\ \text{g}}{1.0\ \text{g mol}^{-1}}$$

Similarly,

$$\text{Moles of O} = \frac{\text{Mass of O}}{\text{Molar mass of O}} = \frac{88.89\ \text{g}}{16.0\ \text{g mol}^{-1}} = 5.55\ \text{mol}$$

Thus in water, the ratio of moles of hydrogen atoms to moles of oxygen atoms is 11.11 : 5.55.

Since a mole of one element contains the same number of atoms as a mole of another element, the ratio of moles of atoms in a compound is also the ratio of the number of atoms. Therefore, the ratio of hydrogen atoms to oxygen atoms is 11.11:5.55. Now by dividing each by the smaller of the two numbers we can convert both numbers to integers

$$\frac{11.11}{5.55} = 2 \text{ and } \frac{5.55}{5.55} = 1$$

Thus ratio hydrogen and oxygen atoms in water is 2 : 1 and empirical formula of water is H_2O.

INTEXT QUESTIONS 1.6

1. For the compound Fe_3O_4, calculate percentage of Fe and O.
2. State percent composition for each of the following:
   (a) C in $SrCO_3$ (b) SO_3 in H_2SO_4
3. What are the empirical formulae of substances having the following molecular formulae?
   H_2O_2, C_6H_{12}, Li_2CO_3, $C_2H_4O_2$, S_8, H_2O, B_2H_6, O_3, S_3O_9, N_2O_3
4. A compound is composed of atoms of only two elements, carbon and oxygen. If the compound contain 53.1% carbon, what is its empirical formula.

1.15 CHEMICAL EQUATION AND REACTION STOICHIOMETRY

You have studied that a reaction can be represented in the form of a chemical equation. A balanced chemical equation carries a wealth of information qualitative as well as quantitative. Let us consider the following equation and learn what all information it carries.

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s) \qquad ...(1.1)$$

1. **Qualitative Information**

Qualitatively the equation (2.1) tells that iron reacts with oxygen to form iron oxide.

2. **Quantitative Information**

Quantitatively a balanced chemical equation specifies numerical relationship among the quantities of its reactants and products. These relationships can be expressed in terms of :

(i) **Microscopic quantities,** namely, atoms, molecules and formula units.

(ii) **Macroscopic quantities**, namely, moles, masses and volumes *(in case of gaseous substances)* of reactants and products.

Now let us again take the reaction (1.1) given earlier and get the quantitative information out of it.

1.15.1 Microscopic Quantitative Information

The reaction (1.1)

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s) \quad ...(1.2)$$

tells that *4 atoms* of iron react with 3 *molecules* of oxygen to form 2 *formula units* of iron oxide. Often this information is written below each reactant and product for ready reference as shown below:

$4Fe(s)$	$+$	$3O_2(g)$	$\rightarrow$	$2Fe_2O_3(s)$	...(1.2a)
4 atoms of Fe		3 molecules of O_2		2 formula units of Fe_2O_3	

1.15.2 Macroscopic Quantitative Information

The microscopic quantitative information discussed in the previous section can be converted into macroscopic information with the help of mole concept which you have learnt in unit 1.

(a) Mole Relationships

We know that Avogadro number of elementary entities like atoms, molecules, ions or formula units of a substance constitute one mole of it. Let us multiply the number of atoms, molecules and formula masses obtained in the previous section (Eq.2.1a) by Avogadro's constant, N_A

$4\ Fe(s)$	$+$	$3O_2(g)$	$\rightarrow$	$2Fe_2O_3(s)$	..(1.3)
4 atoms of Fe		**3 molecules of O_2**		**2 formula units of Fe_2O_3**	
$4 \times N_A$ atoms of Fe		$3 \times N_A$ molecules of O_2		$2 \times N_A$ formula units of Fe_2O_3	
4 mol of Fe		**3 mol of O_2**		**2 mol of Fe_2O_3**	

We may rewrite the above equation as

$4Fe(s)$	+	$3O_2(g)$	$\rightarrow$	$2Fe_2O_3(s)$	...(1.3a)
4 mol of Fe		**3 mol of O_2**		**2 mol of Fe_2O_3**	

The above equation (1.3a) gives us the mole relationship between reactants and products. Here *4 mol of Fe* react with *3 mol of O_2* and produce *2 mol of Fe_2O_3.*

(b) Mass Relationships

The mole relationships which you have learnt in the previous section, can be converted into mass relationship by using the fact that mass of one mole of any substance is equal to its *molar mass* which can be calculated from its formula with the help of relative atomic masses of its constituent elements.

In the reaction that we are discussing, the relative atomic masses of iron and oxygen are 55.8 and 16.0 respectively. Therefore

(i) molar mass of Fe $= 55.8 \text{ g mol}^{-1}$

(ii) molar mass of O_2 $= 2 \times 16.0 = 32 \text{ g mol}^{-1}$

(iii) molar mass of Fe_2O_3 $= (2 \times 55.8 + 3 \times 16.0) \text{ g mol}^{-1}$
$= 159.6 \text{ g mol}^{-1}$

Using these molar masses we can convert the mole relationship given by equation 2.1b into mass relationship as given below :

$4Fe(s)$	+	$3O_2(g)$	$\rightarrow$	$2Fe_2O_3(s)$
4 mol Fe		**3 mol O_2**		**2 mol Fe_2O_3**
(4×55.8) g Fe		(3×32) g O^2		(2×159.6) g Fe_2O_3
223.2 g Fe		96 g O_2		319.2 g Fe_2O_3

Thus 223.2 g iron would react with 96 g oxygen and produce 319.2 g iron oxide, We may rewrite the above equation as

$4Fe(s)$ +	$3O_2(g) \rightarrow$	$2Fe_2O_3(s)$		(1.3b)
223.2 g Fe	**96 g O_2**	**319.2 g Fe_2O_3**		

(c) Volume Relationships

We know that one mole of *any gas* occupies a volume of 22.7 L* at STP (standard temperature and pressure, 0^oC and 1 bar pressure). We can use this information to arrive at volume relationships between gaseous substances. The reaction that we are considering involves only one gaseous substance, O_2. We may rewrite the equation (2.1b) as

$4Fe(s)$	+	$3O_2(g)$	$\rightarrow$	$2Fe_2O_3(s)$	(2.1b)
4 mol		**3 mol**		**2 mol.**	
		(3×22.7) L at STP			
		68.1 L at STP			

* Earlier, the standard pressure was taken as 1 atmosphere and the volume of one mole of gas at STP was taken as 22.4 L.

Thus 4 mol of iron would react with 68.1 L of oxygen at STP to produce 2 mol of iron oxide. (The volume relationship becomes more useful for reactions involving 2 or more gaseous substances).

We can express microscopic as well macroscopic quantitative relationships involved in the above reaction as shown below:

$4Fe(s)$	+	$3O_2(s)$	$\rightarrow$	$2Fe_2O_3(s)$
4 atoms		**3 molecules**		**2 formula units**
4 mol		3 mol		2 mol
223.2 g		96 g		319.2 g
–		68.1 L at STP		–

We may use even mixed relations. For example, we may say 4 mol of iron would react with 68.1 L (at STP) of oxygen to produce 319.2 g of iron oxide.

Let us understand these relationships with two more examples.

(a) Let us work out the mole, mass and volume relationships for the reaction involved in manufacture of ammonia by Haber's process.

Microscopic relationship	$N_2(g)$ +	$3H_2(g)$ $\longrightarrow$	$2NH_3(g)$ ...(2.2)
Microscopic relationships	1 Molecule	3 Molecules	2 Molecules
(i) Moles	1 mol	3 mol	2 mol
(ii) Mass	28 g	$(3 \times 2.0) = 6.0$ g	$(2 \times 17.0) = 34$ g
(iii) Volume	1×22.7 L = 22.7 L	(3×22.7) = 68.1 L	(2×22.7) = 45.4 L
or	1 vol	3 vol	2 vol

(b) Let us take one more reaction, the combustion reaction of butane and work out the different types of relationships. The reaction is :

$2C_4H_{10}(g)$	+	$13O_2(g)$	$\rightarrow$	$8CO_2(g)$	+	$10H_2O(g)$
2 molecules		13 molecules		8 molecules		10 molecules
2 mol		13 mol		8 mol		10 mol
$2 \times (4 \times 12 + 10 \times 1)$ g		(13×32) g		$8 \times (12 + 2 \times 16)$ g		$10 \times (2 \times 1 + 16)$ g
116 g		416 g		352 g		180 g
$2 \times 22.7 = 45.4$ L		$13 \times 22.7 = 295.1$ L		$8 \times 22.7 = 181.6$ L		$10 \times 22.7 = 227$ L
2 vol		13 vol		8 vol		10 vol

Now let us use the mole, mass and volume relationships to make some calculations.

Example 1.9 : In the manufacture of ammonia by Haber process, nitrogen reacts with hydrogen at high temprature and high pressure in the presence of a catalyst and gives ammonia.

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

How much hydrogen would be needed to produce one metric ton of ammonia?

Solution : We should first find out the mass relationships for the reaction.

$$\underset{\substack{\textbf{1 mol}\\ 1\times 28\text{g} = 28\text{ g}}}{N_2(g)} + \underset{\substack{\textbf{3 mol}\\ 3\times 2\text{g} = 6.0\text{ g}}}{3H_2(g)} \longrightarrow \underset{\substack{\textbf{2 mol}\\ 2\times 17\text{g} = 34\text{ g}}}{2NH_3(g)}$$

We know that :

$$1 \text{ metric ton} = 1000 \text{ kg} = 10^3 \text{ kg} = 10^6 \text{ g}$$

From the mass relationship 34 g NH_3 requires 6.0 g H_2 for its manufacture.

$\therefore$ 10^6g NH_3 would require $\dfrac{6.0\times10^6}{34}$ g = 1.76×10^5g of H_2.

Thus 1 metric ton of ammonia will be obtained by using 1.176×10^5g of Hydrogen.

Example 1.10: In a rocket motor fuelled by butane, C_4H_{10}, how many kg of O_2 should be provided with each kg of butane to provide for complete combustion?

Solution : The combustion reaction of butane is

$$\underset{\substack{\textbf{2 mol}\\ \mathbf{2\times 58 = 116\ g}}}{2C_4H_{10}(g)} + \underset{\substack{\textbf{13 mol}\\ \mathbf{13\times 32 = 416\ g}}}{13O_2(g)} \longrightarrow 8CO_2(g) + 10H_2O(g)$$

Thus, to completely burn 116 g butane, oxygen required is 416g.

Therefore, to completely burn 1 kg (1000 g) butane, oxygen required will be

$$= \frac{416\times1000}{116} \text{ g } O_2$$

$$= 3586 \text{ g } O_2$$

$$= 3.586 \text{ kg } O_2 \approx 3.59 \text{ kg } O_2$$

Example 1.11: When lead sulphide; PbS and lead oxide, PbO, are heated together the products are lead metal and sulphur dioxide, SO_2 ,

$$PbS(s) + 2PbO(s) \xrightarrow{\text{heat}} 3Pb\ (1) + SO_2(g)$$

If 14.0 g of lead oxide reacts according to the above equation, how many (a) moles of lead (b) grams of lead, (c) atoms of lead and (d) grams of sulphur dioxide are formed?

(Atomic mass : Pb = 207.0, S = 32.1 ; O =16.0)

Solution : For each part of the question we will use the balanced equation

$$\underset{\textbf{1mol}}{PbS(s)} + \underset{\textbf{2mol}}{2PbO(s)} \xrightarrow{\text{heat}} \underset{\textbf{3 mol}}{3Pb\ (1)} + \underset{\textbf{1mol}}{SO_2(g)}$$

Now formula mass of PbO = (207.0 + 16.0) = 223.0 amu

Thus, one mole of lead oxide formula units have a mass of 223.0 g. Therefore,

14.0 g of PbO is $\dfrac{14.0 \text{ g PbO}}{223.0 \text{ g mol}^{-1} \text{ PbO}} = 6.28 \times 10^{-2}$ mol PbO

(a) The balanced equation shows that 2 mol of PbO form 3 mol of Pb. Therefore, 6.28×10^{-2} mol of PbO form

$$6.28 \times 10^{-2} \text{ mol PbO} \times \frac{3 \text{ mol Pb}}{2 \text{ mol PbO}} = 9.42 \times 10^{-2} \text{ mol Pb}$$

(b) The atomic mass of Pb is 207.0 ; this tells us that one mol of lead has a mass 207.0 g. Thus, 9.42×10^{-2} mol of Pb has a mass of

$$9.42 \times 10^{-2} \text{ mol Pb} \times \frac{207.0 \text{ g Pb}}{1 \text{mol Pb}} = 19.5 \text{ g Pb}$$

(c) 9.42×10^{-2} mol of Pb is

9.42×10^{-2} mol of Pb $\times$ 6.022×10^{23} atoms mol^{-1} = 5.67×10^{22} Pb atoms

(d) The balanced equation shows that 2 mol of PbO form 1 mol of SO_2.

Therefore, 6.28×10^{-2} mol of PbO formula unit forms

$$6.28 \times 10^{-2} \text{ mol PbO} \times \frac{1 \text{mol } SO_2}{2 \text{mol PbO}}$$

$$= 3.14 \times 10^{-2} \text{ mol } SO_2$$

Now the relative molecular mass of SO_2 = 32.1 + 2(16.0) = 64.1

Molar mass of SO_2 = 64.1 g mol^{-1}

Therefore, 3.14×10^{-2} mol of SO_2 molecules have a mass of 3.14×10^{-2} mol $\times$ 64.1 g mol^{-1} = 2.01 g

INTEXT QUESTIONS 1.7

1. How many grams of NH_3 can be made according to the reaction

   $N_2(g) + 3H_2(g) \longrightarrow 2NH_3\,(g)$

   from (a) 0.207 mol of N_2 (b) 22.6 g of H_2

2. In reaction

$$C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(\ell)$$

How many (a) moles of O_2 are consumed and (b) moles of H_2O are formed when 4.16×10^{-2} mol of C_2H_4 react?

1.16 LIMITING REAGENT

We generally find that substances which react with each other are not present in exactly the same proportionin a reaction mixture as stated by a balanced chemical equation. For example, if 2 mol each of hydrogen and oxygen are mixed and a spark is passed through the mixture, water is formed, according to the equation

$$\underset{\textbf{2 mol}}{2H_2} + \underset{\textbf{1mol}}{O_2} \longrightarrow \underset{\textbf{2 mol}}{2H_2O}$$

Here, 2 mol of hydrogen react with only 1 mol of oxygen, and 1 mol of oxygen therefore remains unreacted. In this example hydrogen is said to be the **limiting reagent or reactant** because its amount becomes zero and the reaction therefore stops before the other reactant; that is, the oxygen is used up completly. The amount of hydrogen present initially limits the amount of product that is formed.

Example 1.12: 3 mol of sulphur dioxide SO_2 is mixed with 2 mol of oxygen O_2, and after reaction is over sulphur trioxide, SO_3 is obtained.

(i) Which is the limiting reagent?

(ii) What is the maximum amount of SO_3 that can be formed?

Solution : (i) We must first write the balanced equation

$$2SO_2 + O_2 \rightarrow 2SO_3$$

According to the above equation

(a) 2 mol of SO_3 can be formal from 2 mol of SO_2.

$\therefore$ Amount of SO_3 that can be formed from 3 mol of SO_2.

$$= (3 \text{ mol } SO_2) \times \frac{2\text{mol } SO_3}{2\text{mol } SO_2} = 3 \text{ mol } SO_3$$

(b) 2 mol of SO_3 can be formed from 1 mol of O_2. Therefore, the amount of SO_3 that can be formed from 2 mol of O_2.

$$= (2 \text{ mol } O_2) \times \frac{2\text{mol } SO_3}{1\text{mol } O_2} = 4 \text{ mol } SO_3$$

According to the definition, the limiting reactant is that reactant which gives the smallest amount. In this case SO_2 is the limiting reactant.

(ii) The maximum amount of product that can be obtained is the amount formed by the limiting reagent. Thus a the maximum amount of SO_3 that can be obtained is 3 mol.

Example 1.13: 2.3 g of sodium metal is introduced into a 2L flask filled with chlorine gas at STP (273 K, 1bar). After the reaction is over, find :

(i) What is the limiting reagent in this reaction?

(ii) How many moles of sodium chloride are formed?

(iii) Which substance is left unconsumed at the end of the reaction? Find out its mass in grams.

(iv) What percentage of the substance present in excess is converted into sodium chloride?

(Given : Na = 23, Cl = 35.5)

Solution :

$$\underset{\textbf{2mol}}{2\,Na\,(s)} + \underset{\substack{\textbf{1 mol}\\ 22.7\text{ L at STP}}}{Cl_2(g)} \longrightarrow \underset{\textbf{2 mol}}{2NaCl\,(s)}$$

or

(i) Moles of sodium introduced $= \dfrac{2.3\,g}{23\,g\,mol^{-1}} = 0.1$ mol

From the above equation, it is clear that 2 mol NaCl is formed from 2 mol Na

Therefore 0.1 mol Na can produce $= \dfrac{2 \times 0.1}{2} = 0.1$ mol NaCl

Molar volume at STP = 22.7 L

Therefore moles of chlorine in 2 L volume at STP $= \dfrac{2L}{22.7\,Lmol^{-1}} = 0.088$ mol

From equation : 1 mol Cl_2 can produce 2 mol NaCl

Therefore 0.088 mol Cl_2 can produce $2 \times 0.088 = 0.176$ mol NaCl.

Since sodium produces less amount of NaCl, it is the limiting reagent.

(ii) Sodium being the limiting reagent, as calculated in (i), the moles of NaCl produced = 0.1 mol

(iii) From above equation, 2 mol NaCl is produced from 1 mol Cl_2

Therefore 0.1 mol NaCl is produced from $\dfrac{1 \times 0.1}{2} = 0.05$ mol Cl_2

Initial moles of $Cl_2 = 0.088$ mol

Moles of Cl_2 left unconsumed = (0.088 – 0.05) mol = 0.038 mol

Therefore, mass of Cl_2 left unconsumed = 0.038 g × 71.0 g mol^{-1} = 2.698 g

(because molar mass of Cl_2 = 2 × 35.5 = 71.0 g mol^{-1})

(iv) Moles of Cl_2 consumed = 0.05 mol out of 0.088 mol

$$\therefore \text{ Percent of } Cl_2 \text{ consumed and converted into NaCl} = \frac{0.05}{0.088} \times 100 = 56.8\ \%$$

Example 1.14: 2.0 g mixture of $MgCO_3$ and $CaCO_3$ are heated till no further loss of weight takes place. The residue weighs 1.04 g. Find the percentage composition of the mixture. (Mg = 24, Ca = 40, C = 12, O = 16)

Solution : Mixture of $MgCO_3$ and $CaCO_3$ taken = 2.0 g

Let the mass of $MgCO_3$ be = x g

Therefore the mass of $CaCO_3$ = (2.0 – x) g

The decomposition reactions are

$$\underset{\substack{\mathbf{(24+12+48)\ g} \\ \mathbf{84\ g}}}{MgCO_3\,(s)} \rightarrow \underset{\substack{\mathbf{(24+16)\ g} \\ \mathbf{40\ g\ (Residue)}}}{MgO(s)} + CO_2(g) \qquad \text{(i)}$$

$$\underset{\substack{\mathbf{(40+12+48)\ g} \\ \mathbf{100\ g}}}{CaCO_3\,(s)} \rightarrow \underset{\substack{\mathbf{(40+16)\ g} \\ \mathbf{56\ g\ (Residue)}}}{CaO(s)} + CO_2(g) \qquad \text{(ii)}$$

From the equation (i)

84 g $MgCO_3$ leaves a residue = 40 g

x g $MgCO_3$ will leave residue = $\dfrac{40x}{84}$ g

From the equation (ii)

100 g $CaCO_3$ leaves a residue = 56 g

(2.0 – x) g $CaCO_3$ will leave residue = $\dfrac{56 \times (2.0 - x)}{100}$ g

$$\text{Total mass of the residue} = \frac{40x}{84} + \frac{56 \times (2.0 - x)}{100} = 1.04 \text{ g (given)}$$

$$40 \times 100x + 84 \times 56 \times 2 - 84 \times 56x = 84 \times 100 \times 1.04$$

$$4000x + 9408 - 4704x = 8736$$

$$9408 - 8736 = (4704 - 4000)x$$

$$672 = 704x$$

Therefore, mass of $MgCO_3$ in the mixture = $x = \frac{672}{704} = 0.96$ g

Therefore, percentage of $MgCO_3 = \frac{0.96}{2.0} \times 100 = 48$ %

and percentage of $CaCO_3 = 100 - 48 = 52$ %

WHAT YOU HAVE LEARNT

- Chemistry plays an important role in many aspects of our life like health and medicine, energy and environment, materials and technology, food and agriculture.
- Matter has particulate matter.
- According to the law of conservation of mass, in any chemical reaction, the total mass of all the reactants is equal to the total mass of all the products.
- According to the law of definite proportion, in a chemical compound, the proportions by mass of the elements that compose it are fixed and independent of the origin of the compound or its mode of preparation.
- According to the law of multiple proportions when two elements form two or more compounds, the masses of one element that combine with a fixed mass of the other element are in the ratio of small whole numbers.
- John dalton gave the atomic theory in which he proposed that it is the smallest indivisible particle of matter. Atoms of the same element are all identicle while atoms of different elements differ. Atoms of different elements combine in a simple whole number ratio to form a molecule.
- An atom is the smallest particle of an element that retains its chemical properties.
- A molecule is the smallest particle of matter which can exist independently.
- An element is a substance that cannot be separated into simipler substnaces by chemical means.
- Mole is the amount of a substance which contains as may elementary entities as there are atoms present in 0.012 kg or 12 g of C-12. Thus mole denotes a number.
- The number of elementary entities present in one mole of a substance is 6.022×10^{23}.

- The mass of one mole of a substance is called its molar mass It is numerically equal to relative atomic mass or relative molecular mass expressed in grams per mole (g mol^{-1}) or kilogram per mole (kg mol^{-1}).
- Molar volume is the volume occupied by one mole of a substance. One mole of an ideal gas at standard pressure and temperature, STP (273 K and 1 bar) occupies 22.7 litres.
- In ionic substances, molar mass is numerically equal to the formula mass of the compound expressed in grams.
- If the molar mass of a substance is known, then the amount of a substance present in a sample having a definite mass can be calculated. If M is the molar mass, then, the amount of substance n, present in a sample of mass m is expressed as $n = \frac{m}{M}$.
- A chemical formula is used not only to represent the name of a compound but also to indicate its composition in terms of (i) relative number of atoms and (ii) relative number of moles of atoms.
- A molecular formula of a substance shows(i) the number of atoms of different elements in one molecule.(ii) the number of moles of atoms of different elements in one mole of molecule.
- An empirical formula shows only a ratio of (i) number of atoms, and (ii) moles of atoms in a compound.
- Molecular formula is always an integral multiple of the empirical formula.
- The empirical formula of a compound can be determined from its chemical analysis.
- In order to determine a compound's molecular formula, molecular mass also must be known.
- Stoichiometry is the quantitative study of the composition of chemical compounds (compound or formula stoichiometry) and of the substances consumed and formed in chemical reactions (reaction or equation stoichiometry).
- Chemical equations specify not only the identities of substances consumed and formed in a reaction, but also the relative quantities of these substances in terms of (a) atoms, molecules, and formula units and (b) moles of these entities.
- A balanced chemical equation demonstrates that all the atoms present in the reactants are accounted for in the product; atoms are neither created nor destroyed in a reaction.

- The stoichiometric ratios among the moles of reactants shown in a balanced equation are useful for determining which substance is entirely consumed and which substance(s) is (are) left over.

TERMINAL EXERCISE

1. How many atoms are present in a piece of iron that has a mass of 65.0 g/ (atomic mass; Fe = 55.9 amu).
2. A piece of phosphorus has a mass of 99.2 g. How many moles of phosphorus, P_4 are present in it? (atomic mass, P = 31.0 amu)
3. Mass of 8.46×10^{24} atoms of fluorine is 266.95 g. Calculate the atomic mass of fluorine.
4. A sample of magnesium consists of 1.92×10^{22} Mg atoms. What is the mass of the sample in grams? (atomic mass = 24.3 amu)
5. Calculate the molar mass in g mol^{-1} for each of the following:

   (i) Sodium hydroxide, NaOH

   (ii) Copper Sulphate $CuSO_4 .5H_2O$.

   (iii) Sodium Carbonate, $Na_2CO_3 .10H_2O$
6. For 150 gram sample of phosphorus trichloride (PCl_3), calculate each of the following:

   (i) Mass of one PCl_3 molecule.

   (ii) The number of moles of PCl_3 and Cl in the sample.

   (iii) The number of grams of Cl atoms in the sample.

   (iv) The number of molecules of PCl_3 in the sample.
7. Find out the mass of carbon-12, that would contain 1×10^{19} atoms.
8. How many atoms are present in 100 g sample of C-12 atom?
9. How many moles of $CaCO_3$ would weigh 5 g?
10. If you require 1.0×10^{23} molecules of nitrogen for the reaction $N_2 + 3H_2 \rightarrow 2NH_3$.

    (i) What is the mass (in grams) of N_2 required?

    (ii) How many moles of NH_3 would be formed in the above reaction from 1.0×10^{23} molecules of N_2?

    (iii) What volume would NH_3 gas formed in (ii) occupy at STP?

11. Write empirical formulae of the following compounds:

    CO, Na_2SO_3, C_4H_{10}, H_2O_2, KC1

12. The empirical formula of glucose is CH_2O which has a formula mass of 30 amu. If the molecular mass of glucose is 180 amu. Determine the molecular formula of glucose

13. What is ratio of masses of oxygen that are combined with 1.0 gram of nitrogen in the compound NO and N_2O_3?

14. A compound containing sulphur and oxygen on analysis reveals that it contains 50.1% sulphur and 49.9% oxygen by mass. What is the simplest formula of the compound?

15. Hydrocarbons are organic compound composed of hydrogen and carbon. A, 0.1647 g sample of a pure hydrocarbon on burning in a combustion tube produced 0.5694 g of CO_2 and 0.0845 g of H_2O. Determine the percentage of these elements in the hydrocarbon.

16. On combustion 2.4 g of a compound of carbon, hydrogen and oxygen gave 3.52 g of CO_2 and 1.44 g of H_2O. The molecular mass of the compound was found to be 60.0 amu.

    (a) What are the masses of carbon, hydrogen and oxygen in 2.4 g of the compound?

    (b) What are the empirical and molecular formulae of the compound?

17. (i) What mass of oxygen is required to react completely with 24 g of CH_4 in the following reaction?

    $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(1)$

    (ii) How much mass of CH_4 would react with 96 g of oxygen.

18. In the reaction $H_2 + C1_2 \rightarrow 2HC1$

    How many grams of chlorine, $C1_2$ are needed to react completely with 0.245 g of hydrogen, H_2, to give hydrogen chloride, HC1? How much HC1 is formed?

19. 3.65 g of H_2 and 26.7 g of O_2 are mixed and reacted. How many grams of H_2O are formed?

20. Caustic soda NaOH can be commercially prepared by the reaction of Na_2CO_3 with slaked line, $Ca(OH)_2$. How many grams of NaOH can be obtained by treating 2.0 kg of Na_2CO_3 with $Ca(OH)_2$?

21. A portable hydrogen generator utilizes the reaction

    $CaH_2 + H_2O \rightarrow Ca(OH)_2 + 2H_2$

    How many grams of H_2 can be produced by a 100 g cartridge of CaH_2?

22. The reaction $2Al + 3MnO \rightarrow Al_2O_3 + 3Mn$ proceeds till the limiting substance is consumed. A mixture of 220 g Al and 400 g MnO was heated to initiate the reaction. Which initial substance remained in excess and by how much ? (Al = 27, Mn = 55).

23. $KClO_4$ may be prepared by means of following series of reactions

    $Cl_2 + 2KOH \rightarrow KCl + KClO + H_2O$

    $3KClO \rightarrow 2KCl + KClO_3$

    $4KClO_3 \rightarrow 3KClO_4 + KCl$

    How much Cl_2 is needed to prepare 400 g $KClO_4$ by the above sequence?

    (K = 39, Cl = 35.5, O = 16, H = 1)

24. 2.0 g of a mixture of Na_2CO_3 and $NaHCO_3$ was heated when its weight reduced to 1.876 g. Determine the percentage composition of the mixture.

25. Calculate the weight of 60 % sulphuric acid required to decompose 150 g of chalk (calcium carbonate). Given Ca = 40, C = 12, O = 16, S = 32)

ANSWERS TO INTEXT QUESTIONS

1.1

1. Health, medicine, energy, food, agriculture etc.
2. Leucippus and his student Democritus
3. In every chemical reaction total masses of all the reactants is equal to the masses of all the products.
4. An atom is extremely small particles of matter that retains its identity during chemical reaction.
5. Molecule is an aggregate of at least two atoms in a definite arragement held togethrer its chemical forces.
6. It is derived from the Latin name of sodium i.e. Natrium
7. An elements comprises of atoms of one type only while a compound comprises atoms of two or more types combined in a simple but fixed ratio.

1.2

1. Kilogram
2. μg
3. (i) h (ii) n

4. (i) Megasecond, 10^6 s

(ii) millisecond, 10^{-3} s.

1.3

1. Moles of N_2 gas = $\dfrac{4.22\times10^{23}\text{ molecules}}{6.022\times10^{23}\text{ molecules mol}^{-1}}$ = 0.70 mol

2. Amount of magnesium (moles) = $\dfrac{8.46\times10^{24}\text{ atoms}}{6.022\times10^{23}\text{ atoms mol}^{-1}}$ = 14.05 mol

3. No. of Cl_2 molecules in 0.25 mol Cl_2 = $0.25 \times 6.022 \times 10^{23}$ molecules

= 1.5055×10^{23} molecules

Since each Cl_2 molecule has 2 Cl atoms, the number of Cl atoms = $2 \times 1.5055 \times 10^{23} = 3.011 \times 10^{23}$ atoms.

1.4

1. Molar mass of hydrogen chloride = molar mass of HCl

= 1 mol of H + 1 mol of Cl

= 1.0 g mol^{-1} + 35.5 g mol^{-1}

= 36.5 g mol^{-1}

2. Molar mass of argon atoms = mass of 1 mol of argon

= mass of 6.022×10^{23} atoms of argon.

= 6.634×10^{-26} kg $\times 6.022 \times 10^{23}$ mol^{-1}

= 39.95×10^{-3} kg mol^{-1}

= 39.95 g mol^{-1}

3. Molar mass of KNO_3 = mass of 1 mol of K + mass of 1 mol of N + mass of 3 mol of O.

Since molar mass of an element is numerically equal to its atomic mass but has the units of g mol^{-1} in place of amu = 39.1 g + 14.0 g + 3 × 16.0 g

∴ Molar mass of KNO_3 = 39.1 g + 14.0 g + 48.0 g = 101.1 g mol^{-1}

4. Mass of 1 mol of Na_3PO_4 = 3 × (mass of 1mol of Na) + mass of 1 mol of P + 4 × (mass of 1mol of oxygen)

= 3 (23.0 g) + 31.0 g + 4(16.0) g

= 69.0 g + 31.0 g + 64.0 g = 164.0 g

∴ Mass of 0.146 mol of Na_3PO_4 = 0.146 × 164.0 g = 23.94 g

Notes

1.5

1. Moles of Cu atoms in 3.05 g copper $= \dfrac{3.05 \text{ g}}{63.5 \text{ g mol}^{-1}} = 0.048$ mol

2. Moles of gold, Au $= \dfrac{12.6 \text{ g}}{197 \text{ g mol}^{-1}} = 0.064$ mol

3. Molar volume of any gas at STP (298 K, 1 bar) = 22.7 L

   $\therefore$ Volume occupied by 2.5 mol CO_2 at STP = 2.5 × 22.7 L = 56.75 L

1.6

1. Molar mass of Fe_3O_4 $= 3 \times 56.0 + 4 \times 16.0$

   $= (168.0 + 64.0) = 232.0 \text{ g mol}^{-1}$

   Percentage of Fe $= \dfrac{168.0}{232.0} \times 100 = 72.41\%$

   Percentage of O $= \dfrac{64.0}{232.0} \times 100 = 27.59\%$

2. (a) Molar mass of $SrCO_3$ = 87.6 + 12.0 + 48.0 = 147.6 g mol^{-1}

   Percentage of carbon C in $SrCO_3 = \dfrac{12.0}{147.6} \times 100 = 8.13\%$

   (b) Molar mass of H_2SO_4 = 2.0 + 32.1 + 64.0 = 98.1 g mol^{-1}

   Molar mass of SO_3 = 32.1 + 48.0 = 80.1 g mol^{-1}

   Percentage of SO_3 in $H_2SO_4 = \dfrac{80.1 \times 100}{98.1} = 81.65\%$

3.

Substance	Empirical formula
H_2O_2	HO
C_6H_{12}	CH_2
Li_2CO_3	Li_2CO_3
$C_2H_4O_2$	CH_2O
S_8	S
H_2O	H_2O
B_2H_6	BH_3
O_3	O_3
S_3O_9	SO_3
N_2O_3	N_2O_3

4. Percentage of carbon = 53.1%

Percentage of Oxygen = 46.9%

Suppose we take 100 g of the substance then moles of carbon

$$= \frac{53.1}{12.0} \text{ g} = 4.43 \text{ mol}$$

$$\text{mole of oxygen} = \frac{46.0}{16.0} = 2.93 \text{ mol}$$

$$\text{molar ratio of C and O} = \frac{4.43}{2.93} : \frac{2.93}{2.93}$$

$$= 1.50 : 1 \quad \text{or} \quad 3 : 2$$

Empirical formula of the compound is C_3O_2

1.7

1. In equation

$$\underset{\textbf{1 mol}}{N_2(g)} + \underset{\textbf{3 mol}}{3H_2(g)} \rightarrow \underset{\textbf{2 mol}}{2NH_3(g)}$$

0.207 mol of N_2 gives 0.414 mol of NH_3

0.414 mol of NH_3 = 0.414 mol × 17.0 g mol^{-1} = 7.038 g of NH_3

$$22.6 \text{ g of hydrogen} = \frac{22.6}{2.0} = 11.3 \text{ mol of hydrogen}$$

11.3 mol of hydrogen will give $\frac{2}{3} \times 11.3$ mol of NH_3 = 7.53 mol

Therefore, mass of NH_3 = 7.53 mol × 17.0 g mol^{-1} = 128.01 g

2. $$\underset{\textbf{1 mol}}{C_2H_4(g)} + \underset{\textbf{3 mol}}{3O_2(g)} \rightarrow \underset{\textbf{2 mol}}{2CO_2(g)} + \underset{\textbf{2 mol}}{2H_2O(g)}$$

(a) 4.16×10^{-2} mol of C_2H_4 will consume $3 \times 4.16 \times 10^{-2}$ mol of oxygen

$= 12.48 \times 10^{-2} = 1.248 \times 10^{-1}$ mol of O_2

(b) moles of H_2O formed $= 2 \times 4.16 \times 10^{-2}$ mol

$= 8.32 \times 10^{-2}$ mol of H_2O