Laws of Chemical Combination:

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

(i) Law of conservation of Mass: The law was first stated by Lavosier in 1774. It is also known as the law of indestructibility of matter. According to this law in "all chemical change the total mass of the system remains constant" or "in a chemical change mass is neither created nor destroyed". This law was tested by Landolt. All chemical reactions follow this law.

In chemical change,

Total masses of reactants = Total masses of products

This relationship hold good when reactants are completely conversed into products.

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

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

(ii) Law of Definite or Constant Proportions: This law was presented by Proust in 1799 and may be stated as follows:

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

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

(a) By heating calcium carbonate,

(b) By heating sodium bicarbonate,

(c) By burning carbon in oxygen,

(d) By reacting calcium carbonate with hydrochloric acid, Whatever sample of carbon dioxide is taken, it is observed that carbon and oxygen are always combined in the ratio of 12 : 32 or 3 : 8.

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

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

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

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

Nitrogen forms five stable oxides.

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

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

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

$$\underset{1vol.}{H_2} + \underset{1vol.}{Cl_2} \rightarrow \underset{2vol.}{2HCl} \quad \text{Ratio } 1:1:2$$

$$\underset{2vol.}{2H_2} + \underset{1vol.}{O_2} \rightarrow \underset{2vol.}{2H_2O} \quad \text{Ratio } 2:1:2$$

$$\underset{2vol.}{2CO} + \underset{1vol.}{O_2} \rightarrow \underset{2vol.}{2CO_2} \quad \text{Ratio } 2:1:2$$

$$\underset{1vol.}{N_2} + \underset{3vol.}{3H_2} \rightleftharpoons \underset{2vol.}{2NH_3} \quad \text{Ratio } 1:3:2$$

ATOMIC AND MOLECULAR MASSES

Atoms are too light and small to be weighed individually. The mass of an atom, therefore, is expressed with respect to a standard or reference, when it is called the relative atomic mass or, simply, the atomic mass.

The relative atomic mass of an element is a number which shows how many times an atom of the element is heavier than an atom of a reference element.

Thus, the atomic mass should never be confused with the absolute mass of an atom, which can only be calculated but not directly determined.

Depending upon the reference element, different scales of atomic mass of an atom time to time. All of them contributed significantly to the development of chemistry.

The Hydrogen Scale

Hydrogen, being the lightest element, was first chosen as the reference, and the atomic mass was defined as follows.

The atomic mass of an element is the number of times an atom of element is heavier than an atom of hydrogen.

$$\text{Atomic mass} = \frac{\text{mass of 1 atom of the element}}{\text{mass of 1 atom of hydrogen}}$$

Thus, on the hydrogen scale, a hydrogen atom is assigned a mass of exactly 1 and the masses of the atoms of other elements are determined accordingly.

The Oxygen Scale

From Gay –Lussac's law of combining volumes, we know that one atom of O combines with two atoms of H to form a molecule of water (steam). We also know that 16 parts by mass of oxygen combine with two parts by mass of hydrogen to form 18 parts by mass of water. So that atomic mass of O may be treated as 16.

Considering the greater reactivity of oxygen than hydrogen, chemists shifted the reference from H = 1.000 to O = 16.000. On this scale, an atom of oxygen is granted a mass of exactly 16 and the masses of the other atoms are determined accordingly.

However, when isotopes (atoms of the same element differing in mass number) were discovered, the oxygen scale became obsolete. It was discovered that oxygen has three isotopes = O^{16}, O^{17} and O^{18} with abundances of 99.759, 0.037 and 0.204% respectively. Hence an atom of natural sample of oxygen as a reference has no significance.

Aston, therefore, proposed a scale based on oxygen–16 (O^{16} = 15.9994) as the standard. This scale required the revision of all atomic –mass data compiled earlier on the natural–O scale.

From Aston's work, it was clear that, instead of a natural sample of an element, only a definite isotope could be chosen as a standard. And, for convenience, the choice of the isotope should be such that minimal correction is required in the previously determined atomic masses.

The Carbon –12 Scale

Carbon has three isotopes –C^{12}, C^{13} and C^{14} – of which C^{14} is present in negligible amounts in natural samples. C^{12} and C^{13} have natural abundances of 98.89 and 1.11% respectively.

It was realised that carbon–12 (C^{12} = 12.000) could also be chosen as a good standard. If this was done, the amount of correction required in the earlier data would be minimal. The atomic masses determined according to the natural –O scale would have to be reduced only by 0.004% to make them consistent with the carbon–12 scale. So the carbon–12 scale was finally adopted.

On this scale, one –twelfth the mass of an atom of the isotope ^{12}C is treated as the atomic mass unit (amu), and relative atomic masses are determined accordingly. Relative atomic mass is defined as follows.

The relative atomic mass of an element is the ratio of the mass of an atom of the element to one –twelfth the mass of an atom of carbon–12.

In other words, it is a number that shows how many times an atom of an element is heavier than one –twelfth the mass of an atom of this isotope carbon–12.

$$\text{Relative atomic mass} = \frac{\text{mass of 1 atom of the element}}{\frac{1}{12}\times \text{mass of 1 atom of } C^{12}}$$

$\frac{1}{12}^{\text{th}}$ The mass of 1 atom C^{12}, i.e., 1 amu = 1.66 x 10^{-24} g = 1.66 x 10^{-27} kg.

$\therefore$ The mass of 1 atom of an element = relative atomic mass x 1.66 x 10^{-24} g.

The relative atomic masses of some important elements are given in the table.

Element	Symbol	Atomic number	Atomic mass		Element	Symbol	Atomic number	Atomic mass
Hydrogen	H	1	1.0079		Nickel	Ni	28	58.693
Helium	He	2	4.0026		Copper	Cu	29	63.546
Lithium	Li	3	6.941		Zinc	Zn	30	65.409
Beryllium	Be	4	9.0122		Gallium	Ga	31	69.723
Boron	B	5	10.811		Germanium	Ge	32	72.64
Carbon	C	6	12.011		Arsenic	As	33	74.922
Nitrogen	N	7	14.007		Selenium	Se	34	78.96
Oxygen	O	8	15.999		Bromine	Br	35	79.904
Fluorine	F	9	18.998		Krypton	Kr	36	83.798
Neon	Ne	10	20.180		Rubidium	Rb	37	85.468
Sodium	Na	11	22.990		Strontium	Sr	38	87.62
Magnesium	Mg	12	24.305		Palladium	Pd	46	106.42
Aluminium	Al	13	26.982		Silver	Ag	47	107.87
Silicon	Si	14	28.086		Cadmium	Cd	48	112.41
Phosphorus	P	15	30.974		Tin	Sn	50	118.71

Sulphur	S	16	32.065	Antimony	Sb	51	121.76
Chlorine	Cl	17	35.453	Tellurium	Te	52	127.60
Argon	Ar	18	39.948	Iodine	I	53	126.90
Potassium	K	19	39.098	Xenon	Xe	54	131.29
Calcium	Ca	20	40.078	Cesium	Cs	55	132.91
Scandium	Sc	21	44.956	Barium	Ba	56	137.33
Titanium	Ti	22	47.867	Gold	Au	79	196.97
Vanadium	V	23	50.942	Mercury	Hg	80	200.59
Chromium	Cr	24	51.996	Lead	Pb	82	207.20
Manganese	Mn	25	54.938	Bismuth	Bi	83	208.98
Iron	Fe	26	55.845	Radium	Ra	88	226
Cobalt	Co	27	58.933	Thorium	Th	90	232.04

The Gram–atomic mass

Gram atomic mass is mass of one mole of atoms.

Example

How many gramatoms are there in 135 gm of Aluminium (At mass of Al = 27.0)?

Solution:

The relative atomic mass of Aluminium = 27.0

∴ The gram–atomic mass of Aluminium = 27.0 gm.

Given mass of Aluminium = 135

∴ The number of gram–atoms = $\frac{135.0\,gm}{27\,gm} = 5$

Relative Molecular Mass

The relative molecular mass of a substance is the ratio of the mass of a molecule of the substance to one twelfth the mass of an atom of carbon–12.

Relative molecular mass = $\dfrac{\text{mass of 1 molecule of the substance}}{\frac{1}{12}\times\text{mass of 1atom of}\,C^{12}}$

Thus, the relative molecular mass of a substance is the number that shows how many times a molecule of the substance is heavier than an atom of C^{12}.

The unit of molecular mass is the same as that of atomic mass (i.e., $\frac{1}{12}\times$mass of a C^{12} atom). So the relative molecular mass of a substance – element or compound can be easily calculated by adding the relative masses of all the individual atoms present in the molecule.

Gram–molecular mass

The gram molecular mass is mass of one mole of molecules.

Examples

Substance	Molecular Formula	Relative molecular mass		Gram–molecular mass
1. Hydrogen	H_2	2 x 1	= 2	2 g
2. Oxygen	O_2	2 x 16	= 32	32 g
3. Ozone	O_3	3 x 16	= 48	48 g
4. Chlorine	Cl_2	2 x 35.5	= 71	71 g
5. Neon	Ne	1 x 20	= 20	20 g
6. Water	H_2O	2 x 1 + 16	= 18	18 g
7. Carbon dioxide	CO_2	12 + 2 x 16	= 44	44 g
8. Methane	CH_4	12 + 4 x 1	= 16	16 g
9. Nitric acid	HNO_3	1 + 14 + 3 x 16	= 63	63 g
10. Ethanol	C_2H_5OH	2 x 12 + 5 x 1 + 16 +1	= 46	46 g

Formula Mass

Covalent substances like HCl, CO_2 and CH_4 exist as discrete molecules, but ionic solids do not. For example, a crystal of sodium chlorine does not contain discrete molecules of NaCl; rather it contains Na^+ : Cl^- ratio is 1 : 1 and the formula is NaCl. So, the relative mass of NaCl (58.5; Na = 23, Cl = 35.5) should be called the formula mass rather than the molecular mass.

The same is the case with all other ionic solids. Several covalent substances also do not exist in the form represented by their molecular formulae. For example, water molecules to form $(H_2O)_n$, both in the liquid and the solid state. However, the stoichiometry, i.e., the mass ratio of the elements in the compound, remains the same (1 : 8).

For all chemical calculations the formula mass is treated as the molecular mass as the stoichiometry is same in both cases. Also, the term 'molecular mass' is loosely used for formula mass.

Substance	Formula	Formula mass (amu)	
1. Sodium chloride	NaCl	23 + 35.5	= 58.5
2. Calcium chloride	$CaCl_2$	40 + 2 x 35.5	= 111
3. Calcium oxide	CaO	40 + 16	= 56
4. Sodium hydroxide	NaOH	23 + 16 + 1	= 40
5. Sodium carbonate	Na_2CO_3	2 x 23 + 12 + 3 x 16	= 106
6. Calcium carbonate	$CaCO_3$	40 + 12 + 3 x 16	= 100

The Mole Concept

A chemical change involves atoms, molecules, ions and electrons. We will realise that depicting a chemical change on the basis of the number of particles (atoms, molecules, ions, electrons or any other elementary particles) is easier and at the same time of greater significance than doing so on the basis of their masses.

$$H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$$

2 atoms of hydrogen and 2 atoms of chlorine will combine together chemically to form 2 molecules of hydrogen chloride.

$$Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow Ag^+Cl^-_{(s)} \text{ or } AgCl_{(s)}$$

1 ion of silver and 1 ion of chloride combine together to form a formula unit of silver chloride (solid).

Let us consider the first reaction in a little greater detail.

$$H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$$

1 molecule of hydrogen and 1 molecule of chlorine combine to form 2 molecules of hydrogen chloride.

or

2 atoms of hydrogen and 2 atoms of chlorine combine together to form 2 molecules of hydrogen chloride.

or

1 atom of hydrogen and 1 atom of chlorine can combine to form 1 molecule of hydrogen chloride.

or

1 mole of hydrogen atoms combines with 1 mole of chlorine atoms to form 1 mole of hydrogen chloride molecules.

Since one atom of hydrogen reacts with one atom of chlorine to give one molecule of hydrogen chloride, one may say that, in terms of mass, 1 gramatom of chlorine (35.5g) to give 1 grammolecule of hydrogen chloride (36.5 g).

Thus, 1 g (i.e., 1 gramatom) of hydrogen will contain the same number of atoms as 35.5 g (i.e., 1 gramatom) of hydrogen will contain the same number of atoms as 35.5 g (1 grammolecule) of hydrogen chloride.

Atomic and molecular masses are measured with respect to ^{12}C, which is considered to have a mass of exactly 12. So, the number of atoms in 1 gram–atomic mass of element and that of molecules in 1 gram–molecular mass of a substance must be the same as the atoms in exactly 12 g of ^{12}C. An assemblage of this number (N) of particles is known as a mole (mol when used as a unit), which may be defined as follows.

"A mole is the amount of a substance that contains the same number of particles–atoms, molecules, ions, electrons or any other elementary particles–as the number of atoms is exactly 12 g of carbon–12."

This number of particles (6.022×10^{23}), previously called the Avogadro number, is of great fundamental significance. Nowadays, the term 'Avogadro constant' is preferred, which is simply the Avogadro number suffixed by the unit of mol^{-1}.

Gram atomic mass/ gramatom

So, the gram atomic mass or the gramatom or a mole of an element is defined as the mass in grams of the same number of atoms as contained in exactly 12 g of C^{12}. It is equal to the relative atomic mass expressed in grams.

We can easily arrive at the following important relationships.

1. 1 mole of atom of an element

   = number of atoms in 1 gram atom of the element

   = 6.022×10^{23} atoms.

2. (a) The absolute mass of 1 atoms of an element

   $= \dfrac{\text{gram-atomic mass}}{6.022 \times 10^{23}} g$

   $= \dfrac{\text{gram-atomic mass}}{6.022 \times 10^{23}} \times 10^{-3} kg$

   (b) The absolute mass of 1 atom of an element

   = relative atomic mass x 1.66×10^{-24} g.

3. In a given mass of element,

   (a) The number of gramatoms (mole) = $\dfrac{\text{mass of the an element}}{\text{gram-atomic mass}}$, and

(b) The total number of atoms = $\frac{\text{mass of the element in grams}}{\text{gram-atomic mass}} \times 6.022 \times 10^{23}$

4. No. of moles = $\frac{\text{No. of particles}}{6.022 \times 10^{23}}$

Example

If you need 3.011 x 10^{23} atoms of Na for a chemical reactions, what is the mass of the metal required (relative atomic mass of Na = 23.0)?

Solution

The gram atomic mass of Na = 23.0 g.

For 6.022 x 10^{23} atoms of Na, the mass of the metal required = 23.0g

$\therefore$ For 3.011 x 10^{23} atoms of Na, the mass of the metal required

= $\frac{3.011 \times 10^{23}}{6.022 \times 10^{23}} \times 23.0\,g$ = 11.5 gm

Gram Molecular Mass (GMM)

"The gram –molecular mass or the mole of a substance is defined as the mass in grams of the same number of molecules or formula units of the substance as the number of atoms contained in exactly 12 g of C^{12}". It is equal to the relative molecular mass (M_r) expressed in grams.

As we have said earlier, the formula mass of a substance like an ionic solid or an associated liquid is considered as the molecular mass for purposes of calculation.

The following important relationships may be easily derived.

1. 1 mole of molecules of a substance

= number of molecules in 1 gram –mole of the substance

= 6.022 x 10^{23} molecules

2. The absolute mass of 1 molecule of a substance

= $\frac{\text{gram-molecular mass of the substance}}{6.022 \times 10^{23}}$ (in grams)

= $\frac{\text{gram-molecular mass of the substance}}{6.022 \times 10^{23}}$ x 10^{-3} (in kilograms)

3. In a given mass of a substance,

(a) The number of gram –moles = $\frac{\text{mass of the substance in grams}}{\text{gram-molecular mass}}$

(b) The total number of molecules = the number of gram –moles $\times 6.022 \times 10^{23}$

= $\frac{\text{mass of the substance in grams}}{\text{gram-molecular mass}} \times 6.022 \times 10^{23}$

(c) The number of atoms of a given element constituting the substance = the number of gram–moles of the substance x the number of atoms of the element in 1 molecule of the substance x 6.022 x 10^{23}

(d) The total number of atoms = the number of gram–moles of the substance x the total number of atoms in a molecule of the substance x 6.022 x 10^{23}

Example-1

How many moles of HCl molecules are there in 109.5 g of HCl? Calculate the number of molecules in the sample (Atomic weight of H = 1.0, Cl = 35.5).

Solution

The relative molecular mass of HCl = 1 + 35.5 = 36.5

The number of moles of HCl molecules in 109.5 g

$$= \frac{\text{mass of HCl (in grams)}}{\text{gram-molecular mass}} = \frac{109.5g}{36.5g} = 3.$$

The number of HCl molecules in 109.5 g

= number of moles x 6.022 x 10^{23}

= 3 $\times$ 6.022 $\times 10^{23}$

= 18.066 $\times 10^{23}$

= 1.8066 $\times 10^{24}$

Example-2

How many moles of water are there in 1 L of water? Assume a density of 1.0g mL^{-1}.

Solution

The mass of 1 L (i.e. 1000 mL) of water

= volume x density

= 1000 mL x $\frac{1.0g}{mL}$

= 1000 g.

The gram –molecular mass of water (H_2O) = 2 x 1 + 16 = 18 g.

$\therefore$ The number of moles of water 1 L = $\frac{1000\,g}{18\,g} = 55.55.$

A mole of ions

Ions are formed by the loss or gain of electron(s) from or by an atom. The mass of an electron is negligible in comparison to that of an atom. So, for all practical purposes, atoms have the same mass as do the ions derived from them.

A mole of ions is defined as the assemblage of the same number of ions as the number of atoms in exactly 12g of C^{12}.

In other words, mole ions are an assemblage of 6.022 x 10^{23} ions. Its mass, called the gram–ion, is the same as the gram atom of the element from which it is derived.

For example, a mole of Na^+ ions is a collection of 6.022 x 10^{23} Na^+ ion, which collectively weight 23 g. Similarly, a mole of Cl^- ions is an assemblage of 6.022 x 10^{23} Cl^- ions, which collectively weight 35.5 g. 1 mole of NaCl will furnish 1 mole of Na^+ and Cl^- ions.

$$NaCl \rightleftharpoons Na^+ + Cl^-$$

Mole of electrons:

Electrons have negligible mass and hence they are reckoned only by number and not by mass.

6.022 x 10^{23} electrons constitute a mole.

As you know, all electrons carry an electrical charge.

The charge carried by 1 electron $= 1.602 \times 10^{-19} C$

$\therefore$ The charge carried by 1 mole of electrons $= 1.602\times10^{-19}C\times6.022\times10^{23}$

$\approx$ 96,500 C.

The charge carried by 1 mol of electrons, i.e., 96,500 C, is called a Faraday.

Molar Volume

The volume occupied by one mole i.e., one gram–molecular mass of a gas, at a given temperature and pressure is called the molar volume of the gas.

As the volume of a gas is dependent on temperature and pressure, they must be mentioned along with the molar volume. If the molar volume is taken at STP, it is called the standard molar volume or the molar volume at STP.

It has been found that one gram–mole of every gas occupies 22.4 L at STP. In other words, a collection of 6.022×10^{23} molecules of any gas occupies 22.4 L at STP.

So, the standard molar volume of a gas = 22.4 L.

Thus, as the molecular mass of H_2 is 2, 2 g of H_2 occupies 22.4 L at STP. Similarly $2\times16=32g$ of O_2 will also occupy 22.4 L at STP. Some more examples are given below.

Gas	Molecular Mass	Gram–molecular mass	Standard Molar volume
H_2	2 x 1 = 2	2 g	22.4 L
N_2	2 x 14 = 28	28 g	22.4 L
O_2	2 x 16 = 32	32 g	22.4 L
Cl_2	2 x 35.5 = 71	71 g	22.4 L
CO_2	12 + 2 x 16 = 44	44 g	22.4 L
CH_4	12 + 4 x 1 = 16	16 g	22.4 L
SO_2	32 + 2 x 16 = 64	64 g	22.4 L

Example

What is the volume occupied by 17.75 g of Cl_2 at STP?

Solution

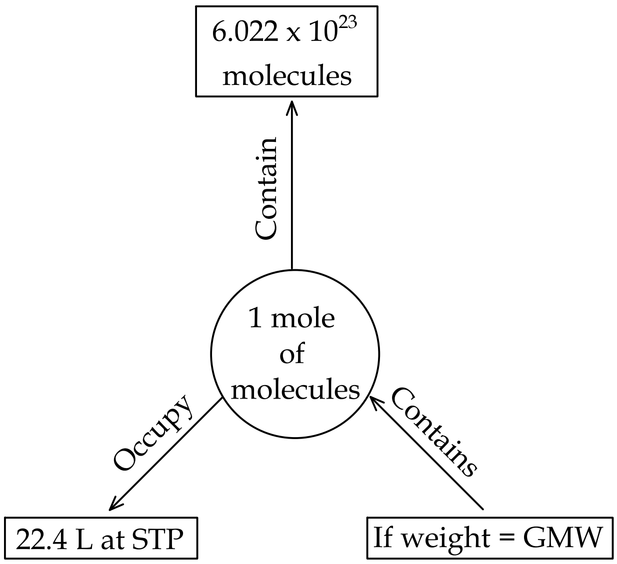
The relative molar mass of Cl_2 = 2 x 35.5 = 71.

The number of moles of Cl_2 in 17.75 g of the gas

$$= \frac{\text{the mass of } Cl_2 \text{ gas in atom}}{\text{gram-molecular mass}} = \frac{17.75}{71g} = \frac{1}{4}$$

The volume of 1 mol of Cl_2 at STP = 22.4 L.

$\therefore$ The volume of ¼ mol of Cl_2 at STP = ¼ x 22.4 = 5.6 L.



Mole concept chart

Molecular mass from vapour density

The density of a gas relative to that of hydrogen at the same temperature and pressure is known as the vapour density or the hydrogen density of the gas.

Thus, the vapour density of a gas = $\dfrac{\text{density of the gas at a given } P \text{ and } T}{\text{density of hydrogen at the same } P \text{ and } T}$

$$= \frac{\text{mass of given volume of the gas at a given } P \text{ and } T}{\text{mass of the same volume of hydrogen at the same } P \text{ and } T}$$

$$= \frac{\text{mass of } 22.4\,L \text{ of the gas at } STP}{\text{mass of } 22.4\,L \text{ of hydrogen at } STP}$$

$$= \frac{1 \text{ gram-molecule mass of the gas}}{2\,g}$$

$\therefore$ The relative molecular mass of a gas = 2 $\times$ vapour density.

The vapour densities of gases can be determined experimentally. So can those of many other substances like methanol, ethanol and acetone, which can be completely volatilized. Thus, it is easy to calculate the relative molecular masses of such substances.

Substance	Vapour density	Relative molecular mass (= 2 x vapour density)	Molecular formula
Nitrogen	14	28	N_2
Oxygen	16	32	O_2
Chlorine	35.5	71	Cl_2
Methane	8	16	CH_4
Ethane	15	30	C_2H_6
Carbon dioxide	22	44	CO_2
Methanol	16	32	CH_3OH
Ethanol	23	46	C_2H_5OH
Acetone	29	58	$(CH_3)_2CO$

Atomicity

The atomicity of a gaseous element can be easily calculated if we know its vapour density and atomic mass.

Atomicity = $\dfrac{\text{molecular mass}}{\text{atomic mass}} = \dfrac{2\times\text{vapour density}}{\text{atomic mass}}$.

For example, the vapour density of nitrogen is 14 and so its molecular mass = 2 $\times$ 14 = 28

As the atomic mass of nitrogen is 14, its atomicity = $\dfrac{28}{14}=2$.

Hence the molecular formula of nitrogen is N_2.

Percentage Composition

The percentage composition of a compound in terms of the different elements constituting it can be easily calculated if we know the molecular or the empirical formula of the compound and the atomic masses of its constituents.

Example

What is the percentage of Ca is $CaCO_3$? (Ca = 40, C = 12, O =16)

Solution

The relative molecular mass of $CaCO_3$ = 40 + 12 + 3 x 16 = 100.

100 g of $CaCO_3$ contain 40 g of Ca.

$\therefore$ The percentage of Ca in $CaCO_3$ = 40.

Empirical and Molecular Formula from Percentage Composition

Empirical Formula

The empirical formula of compound shows the simplest atomic ratio of the elements present in a molecule of the compound. For example, the empirical formula of hydrogen peroxide (H_2O_2) is HO, because the simplest atomic ratio of hydrogen and oxygen in a molecule of it is 1 : 1.

The empirical formula of a compound can be determined if the percentage composition of the elements in the compound is known.

The percentage of each element in the compound gives the mass (in grams) present in 100g of the compound. So, if we divide the percentage by the atomic mass of the element, we get the number of moles of atoms of the element present in 100 g of the compound. The ratio of the number of moles of the different elements thus calculated can be divided by a suitable common factor to obtain the simplest whole–number ratio of the atoms present in the compound. Thus the calculation is carried out in the following steps.

1. Tabulate the percentage composition of the different elements.
2. Divide the percentage of each element by the respective atomic mass to obtain the number of moles of atoms of the element.
3. Divide the numbers of moles by a suitable common factor to obtain them in the simplest whole–number ratio.
4. Deduce the empirical formula of the compound from the whole–number of the atoms.

Example

Determine the empirical formula of compound if it contains 5.9% H and 94.1% O.

Solution

Element	Percentage	Atomic mass	Moles of atoms	Atomic ratio
H	5.9	1	$\dfrac{5.9}{1}=5.9$	5.9/5.9 = 1
O	94.1	16	94.1/16 =5.9	5.9/5.9 = 1

The empirical formula is HO.

Molecular formula

The molecular formula can be derived from the empirical formula when the molecular mass of the compound is known.

As you know, the empirical formula gives the ratio of the atoms of the elements present in a molecule of a compound. The molecular formula gives the actual number of atoms of the different elements present in a molecule of the compound.

$(\text{Empirical formula})_n$ = molecular formula, where 'n' is an integer.

$\therefore$ Relative molecular mass = n x empirical formula mass.

So, if we know the empirical formula mass (which can be easily calculated) and the molecular mass (which can be experimentally determined), we can calculate 'n' and thus determine the molecular formula of the compound.

If the substance is a gas or a compound which can be completely volatilised, the molecular mass can be easily calculated from the relative density (or the vapour density).

Example-1

The empirical formula of a compound is HO and the relative molecular mass 34. Deduce he molecular formula of the compound.

Solution

Let the molecular formula be $(HO)_n$.

The empirical formula mass = 1 + 16 = 17.

We know that

n $\times$ empirical formula mass = relative molecular mass

$\Rightarrow n \times 17 = 34$

$\Rightarrow \frac{34}{17} = 2.$

Therefore the molecular formula of the compound is $(HO)_2$, i.e., H_2O_2.

Example-2

Deduce the molecular formula of a compound whose empirical formula is CH_3 and relative molecular mass 30.

Solution

Let the molecular formula be $(CH_3)_n$.

The empirical formula mass = 12 + 3 = 15.

We know that

n $\times$ empirical formula mass = relative molecular mass

$\Rightarrow n \times 15 = 30$

$\Rightarrow \frac{30}{15} = 2.$

Therefore the molecular formula of the compound is $(CH_3)_2$, i.e., C_2H_6.

Stoichiometry

We know that atoms combine in whole–number ratios to form compounds and also that reactants and products appear in chemical equations in simple ratios.

The relative proportions of elements in a compound or those of the reactants and products in a chemical reaction are known as stoichiometry.

We will see in the following sections how the percentage composition of a compound and also the masses and volumes taking part in a chemical reaction help us arrive at the stoichiometry.

Problems Based on Chemical Equations

A chemical equation is a precise way of expressing a lot of information about the way a reaction takes place. The following steps will help you use this information to solve various problems.

1. Write a chemical equation to represent the reaction and balance it.
2. Write the number of moles of each reactant and product. These numbers are the same as the numbers of molecules or ions on both sides of the equation.
3. Convert the number of moles of each reactant and product into mass or volume (if gaseous) as convenient or required.
4. Calculate the mass or volume of one or more of the reactants and products on the basis of the data given.

How well the method works can be easily seen in the following example.

Suppose you have to calculate the mass of water formed by the reaction between 11.2 L of H_2 and 5.6 L of O_2 taken at STP. Proceed as follows.

$2H_{2(g)}$ +	$O_{2(g)}$ $\longrightarrow$	$2H_2O_{(l)}$
2 mol	1 mol	2 mol
2 x 22.4 L (STP)	22.4 L (STP)	2 x 18 g
22.4 L (STP)	11.2 L (STP)	18g
11.2 L (STP)	5.6 L (STP)	9 g

So, the mass of the water formed in the reaction = 9 g.

Example

Hydrogen and chlorine react in the gram–atomic ratio of 1 : 1 to form HCl. If 4.0 g of hydrogen is available for the reaction, what is the mass of chlorine required for the complete reaction to take place (A_r of H = 1.0, Cl = 35.5)?

Solution

A_r of H = 1.0. $\therefore$ The gram atomic mass of H = 1.0 g.

A_r of Cl = 35.5. $\therefore$ The gram–atomic mass of Cl = 35.0 g.

The gram–atomic ratio for the reaction between hydrogen and chlorine = 1:1

$\Rightarrow$ 1 gram–atom of hydrogen requires 1 gram–atom of chlorine

$\Rightarrow$ 1.0g of hydrogen requires 35.5 g of chlorine

$\therefore$ 4.0 g of hydrogen requires 4.0 x 35.5 = 142.0g of chlorine.

Practice Problems

1. What masses of the following elements will contain 6.022×10^{23} atoms each?

   (a) Hydrogen (b) Sodium (c) Chlorine (d) Copper

2. Calculate the absolute masses of 1 atom of H, N, O and Na
3. How many gramatoms are there in 106.5g of chlorine (relative atomic mass of Cl =35.5)?
4. How many atoms of Na are present in 46.0 g of metal (A_r of Na = 23.0)?
5. For 196g of pure H_2SO_4, calculate

   (a) the number of moles of H_2SO_4 (b) the total number of H_2SO_4 molecules,

   (c) the total number of atoms, (d) the number of atoms of each kind, and

   (e) the absolute mass of an H_2SO_4 molecule. (A_r of H = 1, O = 16, S =32.)

6. The mass of an atom of an element is 3.986×10^{-26} kg. Find the relative atomic mass of the ement and identify it.

7. How many moles of carbon atoms are there in 540g of glucose($C_6H_{12}O_6$)?

   (A_r of C = 12, H = 1, O = 16.)

8. How many moles of Ca^{2+} ions and Cl^- ions will be furnished by 444 g of $CaCl_2$

   (A_r of Ca = 40, Cl = 35.5)?

9. What will be the mass of 1.4 L of CO_2 at STP?

10. Calculate the percentages of H and O in H_2O. (H = 1, O = 16)

11. Calculate the percentage of Cu and water of crystallization in blue vitriol $CuSO_4 . 5H_2O$).

    (Cu = 63.5, S = 32, O = 16, H = 1)

12. Calculate the percentage composition of ethanol (C_2H_5OH) in terms of all the elements present in the compound.

13. A hydrocarbon contains 20% H and 80% C. Deduce its empirical formula.

14. A compound contains 6.7% H, 40% C and 53.3% O. Determine its empirical formula.

15. Deduce the empirical formula of a compound containing 1% H, 33% S and 66% O.

16. Ammonium chloride contains 31.8% NH_3 and 66.3% Cl^-. Work out the empirical formula of the compound (H = 1, N = 14, O = 16, Cl = 35.5).

17. Acetylene and benzene have the same empirical formula CH, but vapour densities of 13 and 39 respectively. Deduce their molecular formulae.

18. An organic compound of vapour density 29 contains 62.07% C, 10.34% H and 27.59% O. Deduce the molecular formula of the compound.

19. An oxoacid of sulphur, relative molecular mass 194, contains 33% S and 66%. Deduce the molecular formula of the acid.

20. Anhydrous aluminum chloride contains 20.22% Al and 79.78% Cl. Though a solid, it volatilizes completely on heating. AT 350°C, the vapour density is 133.5 and at 750°C, 66.75. Deduce the molecular formulae of the salt at the two temperatures.

21. 2.5 gram–atoms of carbon are required for a particular reaction. How many grams of carbon would you take ($A_{t.}$ mass of C = 12.0)?

22. What is the volume of O_2 required to burn 18 kg of C and of the CO_2 product in the reaction? Assume that the volumes are measured at STP.

23. What is the volume of O_2 (measured at STP) produced when 61.25 g of $KClO_3$ is strongly heated (K = 39, Cl = 35.5, O = 16)? $2KClO_3 \longrightarrow 2KCl + 3O_2$

24. How much $KClO_3$ must be heated to produce as much O_2 as required to burn 24 g of carbon (K = 39, Cl = 35.5, O = 16)? $2KClO_3 \longrightarrow 2KCl + 3O_2$ $C + O_2 \longrightarrow CO_2$

25. MnO_2 oxidises HCl to Cl_2 and forms $MnCl_2$. What is the number of moles of HCl and the mass of MnO_2 required to produce 44.8 L of Cl_2 at STP (Mn = 55, Cl = 35.5, O =16)?

    $MnO_2 + 4HCl \longrightarrow MnCl_2 + Cl_2 + 2H_2O$

Assignment Problems

Single Correct Type

1. How many moles of electron weigh one kilogram?

(A) $6.023x10^{23}$ (B) $\frac{1}{9.108}\times10^{31}$ (C) $\frac{6.023}{9.108}\times10^{54}$ (D) $\frac{1}{9.108\times6.023}\times10^{8}$

2. Which has maximum number of atoms?

(A) 24g of C (12) (B) 56g of Fe (56) (C) 27g of Al (27) (D) 108g of Ag (108)

3. 3g of Mg is burnt in a closed vessel containing 3g of oxygen. The weight of excess reactant left is

(A) 0.5 g of oxygen (B) 1.0 g of oxygen (C) 1.0 g of Mg (D) 0.5 g of Mg

4. The weight of chlorine gas undergoing oxidation when excess of chlorine is sent into hot concentrated caustic soda solution containing 24 grams of NaOH in it is $3Cl_2+6NaOH\rightarrow5NaCl+NaClO_3+3H_2O$

(A) 21.3g (B) 17.75g (C) 3.55g (D) 10.65g

5. The density of a plant virus is $1.66g/cc$. If each virus particles are spheres of diameter 6^oA. The molecular weight of virus should be (assuming entire space is occupied by virus only)

(A) 36π (B) 72π (C) $5.976x10^{-23}\pi$ (D) 18π

6. The number of F^- ions in 4.2g of AlF_3

(A) 0.05 (B) $9x10^{22}$ (C) $3x10^{22}$ (D) 0.15

7. Find the number of molecules in 39.2g of nitrogen

(A) $16.86\ x\ 10^{23}$ (B) $8.43\ x\ 10^{23}$ (C) $2.11\ x\ 10^{23}$ (D) $4.22\ x\ 10^{23}$

8. What is the volume of 240g of oxygen at STP?

(A) 172L (B) 42L (C) 16.8L (D) 168L

9. The number of hydrogen atoms in 0.9g glucose $C_6H_{12}O_6$, is same as

(A) 0.48g N_2H_4 (B) 0.17g NH_3 (C) 0.30g C_6H_6 (D) 0.03g H_2

10. The weight of one millimole of water is

(A) $1.8\ x\ 10^{2}g$ (B) $1.8\ x\ 10^{-2}g$ (C) $5.4\ x\ 10^{-2}g$ (D) $2.7\ x\ 10^{-2}g$

11. Calculate the total number of moles of molecules present in a container if it contains 8g of O_2 and 7g of N_2

(A) 1mole (B) $\frac{1}{2}$ mole (C) $\frac{1}{4}$ mole (D) $\frac{3}{4}$ mole

12. If a container has $1.5x10^{24}$ molecules of NH_3 gas what is the weight of NH_3 in the container?

(A) 17g (B) 62g (C) 42.5g (D) 34g

13. A sample of $(NH_4)_2SO_4$ has 3.18 mol of hydrogen atoms. Calculate the number of moles of oxygen atoms in the sample?

(A) 0.265 (B) 0.795 (C) 1.06 (D) 1.59

More Than One Correct Type

14. Which of the following contains the same number of molecules?

(A) 1 g of O_2, 2g of SO_2

(B) 1 g of CO_2, 1 g of N_2O

(C) 112 mL of O_2 at STP, 224 mL of He at STP

(D) 1 g of O_2, 1 g of ozone

15. 0.22 g of a gas occupies a volume of 112 mL at a pressure of 1 atm and 273K. The gas can be

(A) NO_2 (B) N_2O (C) CO_2 (D) C_3H_8

16. A 100 mL mixture of CO and CO_2 is passed through a tube containing red hot charcoal. The volume now becomes 160 mL. The volumes are measured under the same conditions of temperature and pressure. Amongst the following which is /are correct statement(s).

(A) Ratio of moles of CO_2 and CO is 3 : 2

(B) Ratio of molecules of CO_2 and CO is 2 : 3

(C) Ratio of volumes of CO_2 and CO is 2 : 3

(D) Ratio of volumes of CO_2 and CO is 3 : 2

17. The volume occupied by 3.0115×10^{23} molecules of O_2 and 3.0115×10^{23} molecules of N_2 at STP is equal to

(A) gram molar volume of N_2 at STP

(B) double the volume of 14g of N_2 at STP

(C) 16g of CH_4 at STP

(D) 22.4 L at STP

18. The reaction $2C+O_2\longrightarrow 2CO$ is carried out by taking 24 grams of carbon and 96 g of O_2

(A) O_2 is left in excess (B) 56 g of CO is formed

(C) 2 moles of O_2 is left

(D) To use O_2 completely, 72 g carbon is needed

19. In a reaction vessel, 100gm H_2 and 100gm Cl_2 are mixed and suitable conditions are provided for the reaction. $H_2+Cl_2\rightarrow HCl$

(A) Cl_2 is the limiting reagent

(B) 102.8gm HCl is formed

(C) 2 moles HCl is formed

(D) H_2 is the limiting reagent

20. Number of atoms in 560 gm of Fe(at. mass =56) is

(A) twice that of 70 gm N

(B) half of that of 20 gm of H

(C) thrice that of 320 gm of O

(D)same as 120 gm of C

21. $\underline{P}$ and $\underline{Q}$ are two elements which forms P_2Q_3 and PQ_2. If 0.15 mole of P_2Q_3 weighs 15.9g and 0.15 mole of PQ_2 weighs $9.3g$

(A) atomic weight of P is 18

(B) atomic weight of P is 26

(C) atomic weight of Q is 18

(D) atomic weight of Q is 26

22. A bulb contains 1.6 g of O_2. It contains

(A) 0.05 mol of O_2

(B) 3.011×10^{22} molecules of O_2

(C) $1.12L$ of O_2 at STP

(D) 1.22L of O_2 at STP

23. From 3 moles of C_2H_6 gas is removed, it is followed by the removal of 3×10^{23} molecules further. The leftover gas is combusted in the presence of excess oxygen then $\left(N_A=6\times10^{23}\right)$ (density of H_2O is 1g/mL)

(A) 2 moles of C_2H_6 left for combustion
(B) volume of CO_2 at STP after combustion 44.8 litres
(C) volume of water produced is 54 mL
(D) none

24. Complete combustion of 1-mole of ethane $\left(C_2H_6\right)$ at STP

(A) produces $44.8L$ of CO_2
(B) produces $54g$ of H_2O
(C) requires 3.5 moles of oxygen
(D) requires 7 gram atoms of oxygen

25. The molar mass of Haemoglobin is about 65000 $g.mol^{-1}$. Every Haemoglobin contains 4 Iron atoms. Thus,

(A) Iron content in Haemoglobin is 0.35% by mass.
(B) 1 mole of Haemoglobin contains 56 grams of iron
(C) 1 mole of Haemoglobin contains 224 grams of iron
(D) If iron content is increased to 0.56% molar mass of Haemoglobin would be higher than $65000\,g.mol^{-1}$

26. 11.2 L of a gas at STP weighs 14 grams. The gas could be

(A) N_2O
(B) NO_2
(C) N_2
(D) CO

27. 100 mL of CO and CO_2 is passed through a tube contain red hot charcoal (carbon). The volume becomes 160 mL. The volumes are measured under the same conditions of temperature and pressure. (Reaction involved in this is: $CO_2+C\longrightarrow CO$)

Select the correct statements

(A) Mole percent of CO_2 in the mixture is 60
(B) Mole fraction of CO in the mixture is 0.4
(C) The mixture contains 40 mL of CO_2
(D) The mixture contains 40 mL of CO

28. Which of the following contain Avagadro number of atoms?

(A) one mole of Helium gas
(B) 22.4 litres of CO_2 at STP
(C) 11.2 litres of Hydrogen gas at STP
(D) 3.2 gms of methane

29. Equal masses of oxygen and ozone at a given temperature and pressure contain

(A) equal number of moles
(B) equal masses
(C) equal number of 'gramatoms'
(D) equal number of respective molecules

30. Which of the following statement is correct for decomposition of $NaHCO_3$?

$$NaHCO_3 \xrightarrow{\Delta} Na_2CO_3+CO_2+H_2O$$

(A) 84g of $NaHCO_3$ gives 106g of Na_2CO_3
(B) 16.8g of $NaHCO_3$ gives 0.1 mole of Na_2CO_3
(C) 0.1 mole of $NaHCO_3$ gives 1.12L of CO_2 at STP
(D) 84g of $NaHCO_3$ gives 22.4L of CO_2 at STP

31. How many of the following samples have same number of oxygen atoms as in 56 g of CO

(A) 40.83g of $KClO_3$ (B) 44g of CO_2 (C) 64g of SO_2 (D) 32g of O_2

32. From 440g of CO_2 sample, 24.092×10^{23} molecules were removed. Identify the correct statement?

(A) The number of moles of CO_2 left is 6 (B) The mass of CO_2 left is 264g

(C) 178g of CO_2 is removed from the sample

(D) In the original sample, $20N_0$ oxygen atoms were present

Integer Type

33. Consider the reaction $2A+B+3C\rightarrow P+2Q$ starting with 3 moles of A, 1.5 moles of B and 6 moles of C, the total number of moles of the final mixture is _________

34. Analysis of chlorophyll shows that it contains 2.68% by mass Mg. How many grams of chlorophyll contains 6.72×10^{20} atoms.

35. SO_3 is prepared by the following two reactions:

$$S_8(s)+8O_2(g)\longrightarrow 8SO_2(g)$$

$$2SO_2(g)+O_2(g)\longrightarrow 2SO_3(g)$$

If x grams of SO_3 are produced from 0.1 Mole S_8, then find $\sqrt{x}$ value.

36. If x grams of $CaCO_3$ on strong heating yields 28 grams of solid residue, then find $\frac{x}{10} value$

37. PQ_2 and P_2Q_3 are two compounds of the elements P and Q. 0.15 mole of each of these compounds weighs 9.3 and 15.9 gm. Respectively, then find the atomic weights difference of P and Q

38. Calculate the mass of $\frac{N_A}{4}$ molecules of Nitrogen in gms

39. If the number of electrons present in 18 gm of H_2O is equal to 'y' times of Avogadro's Number, then find $\frac{y}{5}$ value

40. The weight of iron which will be converted into its oxide by the action of 18g of steam is $7x$. The value of x is $\left[Fe+H_2O\longrightarrow Fe_3O_4+H_2\right]$ [Atomic weight of $Fe=56g$]

41. The volume of O_2 (in L) required at STP to completely oxidise 4L of CO is _____.

$(CO+O_2\rightarrow CO_2)$

42. The mass of 2.5 gm atoms of an element is 100 gm. If its gram atomic weight is x gm. Find $\frac{x}{10}$ value___

43. If 5 gram H_2 is mixed with 14 gram of nitrogen for the following reaction

$$N_2+3H_2\longrightarrow 2NH_3$$

At the end, mass of H_2 left unreacted.

44. $5L$ of ozone is converted to oxygen. At the end of reaction, the volume of mixture becomes $7L$. The volume of ozone, left at the end (in L) $O_{3(g)}\rightarrow O_{2(g)}$

Comprehension Type

Passage-I

10 moles of SO_2 and 4 moles of O_2 are mixed in a closed vessel of volume 2L. The mixture is heated in presence of Pt catalyst and the following reaction takes place:

$$\underset{(s)}{2SO_2} + \underset{(g)}{O_2} \rightarrow \underset{(g)}{2SO_3}.$$

45. Number of moles of SO_3 formed in the reaction is

(A) 10 (B) 4 (C) 8 (D) 14

46. Number of moles of excess reactant remaining is

(A) 4 (C)2 (C) 6 (D) 8

Paragraph-2

Calcium carbonate on heating gives calcium oxide and carbon dioxide the reaction is

$$CaCO_3 \uparrow \xrightarrow{\Delta} CaO + CO_2 \uparrow$$

47. In a container from 100 grams of calcium carbonate 0.25 moles of $CaCO_3$ removed. The number of molecules and volume of CO_2 gas formed when left over $CaCO_3$ is completely dissociated

(A) 0.5 N_0, 11.2 L at STP (B) 1 N_0, 22.4 L at STP

(C) 0.25 N_0, 6.6 L at STP (D) 0.75 N_0, 16.8 L at STP

48. The mass of left over calcium carbonate before dissociation is

(A) 25 g. (B) 50 g. (C) 75 g. (D) 99 g.

Paragraph - 3

In chemistry, 'mole' is an essential tool for the chemical calculations. It is a basic SI unit adopted by the 14th general conference on weights and measurements 1971. A mole contains as many elementary particles as the number of atoms present in 12g of ^{12}C. 1 mole of a gas at STP occupies 22.4 litre volume. Molar volume of solids and liquids is not definite. Molar mass of a substance also called gram-atomic mass or gram molar mass. The virtual meaning of mole is plenty, heap or the collection of large numbers. 1 mole of a substance contains 6.023×10^{23} elementary particles like atom or molecule. Atomic mass unit (amu) is the unit of atomic mass, e.g., atomic mass of single carbon is 12 amu.

49. The mass of one amu is approximately

(A) 1 g (B) 0.5 g (C) 1.66×10^{-24} g (D) $3.2\times10^{-24}g$

50. 5.6 litre of a gas at STP are found to have a mass of 22 g. The molecular mass of the gas is

(A) 22 (B) 44 (C) 88 (D) 33

Paragraph-4

The reaction between sulphuric acid and $NaOH$ can be represented as below $H_2SO_4 + NaOH \longrightarrow NaCl + H_2O$

51. If $9.8g$ of H_2SO_4 is mixed with 80 g. of $NaOH$ the amount of $NaCl$ formed

(A) $11.7g$ (B) $1.17g$ (C) $5.85g$ (D) $58.5g$

52. The mass of $NaOH$ left at the end of reaction

(A) $72g$ (B) $64g$ (C) $36g$ (D) $32g$

Matrix Match Type

53. Match the following

Column – I		Column – II (Left out reactant)	
(A)	$\underset{1g}{2H_2}+\underset{1g}{O_2} \rightarrow 2H_2O$	(p)	0.971 g
(B)	$\underset{1g}{3H_2}+\underset{1g}{N_2} \rightarrow 2NH_3$	(q)	0.666g
(C)	$\underset{1g}{H_2}+\underset{1g}{Cl_2} \rightarrow 2HCl$	(r)	0.875 g
(D)	$\underset{1g}{2H_2}+\underset{1g}{C} \rightarrow CH_4$	(s)	0.785 g
		(t)	0.075 g

54. Match the following

Column–I (Compound)		Column–II (Weight of oxygen in g. present)	
(A)	2 moles of oxygen atoms	(p)	16
(B)	1 mole of $KClO_3$	(q)	48
(C)	5.6 L of CO_2 at STP	(r)	32
(D)	1 mole of H_2O	(s)	8

55. Match the following

Column – I		Column – II	
(A)	1-gram molecule of oxygen	(p)	Contains 1 mole of oxygen atoms
(B)	1-mole of NO gas	(q)	Contains 2 moles of oxygen atoms
(C)	44 grams of CO_2	(r)	Occupy 22.4 L of STP
(D)	18 grams of water vapour	(s)	Contains $2N_A$ atoms
		(t)	Contains $3N_A$ atoms

56. Match the following:

	Column – I		Column – II
(A)	52g of He	(P)	13 atoms
(B)	52 moles of He	(Q)	44.8 L
(C)	34 g of NH_3	(R)	313.196×10^{23} atoms
(D)	52 u of He	(S)	78.299×10^{23} atoms

KEY

Assignment Problems

1.	D	2.	A	3.	B	4.	A	5.	A	6.	B
7.	B	8.	D	9.	A	10.	B	11.	B	12.	C
13.	D	14.	AB	15.	BCD	16.	AD	17.	ABCD	18.	ABCD
19.	ABC	20.	ADD	21.	BC	22.	ABC	23.	BC	24.	ABCD
25.	AC	26.	CD	27.	ABD	28.	ACD	29.	BC	30.	BC
31.	BCD	32.	ABD	33.	6	34.	1	35.	8	36.	5
37.	8	38.	7	39.	2	40.	6	41.	2	42.	4
43.	2	44.	1	45.	C	46.	B	47.	D	48.	C
49.	C	50.	C	51.	A	52.	A				

53. A-r, B-s, C-p, D-q

54. A-r, B-p, C-q, D-s

55. A-q,r,s; B-p,r,s; C-q,r,t; D-p,r,t

56. A-s ; B-r; C-q ; D-p