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# SOME BASIC CONCEPTS IN CHEMISTRY

## CLASSIFICATION OF MATTER

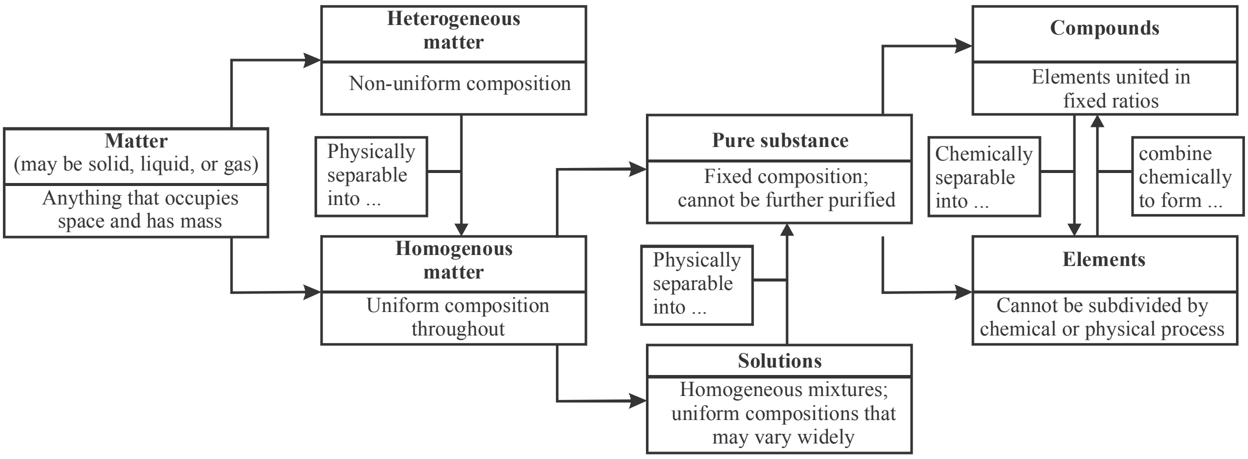



**Figure : Classifying matter**

**Pure substance**

* Every substance has a set of unique properties by which it can be recognized. Pure water, for example, is colourless, odourless & certainly does not contain suspended solids.
* A pure substance is that it cannot be separated into two or more different species by any physical technique such as heating in a bunsen flame. If it could be separated out, sample would be classified as a mixture.

**Mixture**

* A **mixture** in which the uneven texture of the material can be detected is called a **heterogeneous** mixture. Heterogeneous mixtures may appear completely uniform but on closer examination are not. Blood, for example, may not look heterogeneous until you examine it under a microscope and red and white blood cells are revealed.
* Milk appears smooth in texture to the unaided eye, but magnification would reveal fat and protein globules within the liquid.
* In a heterogeneous mixture the properties in one region are different from those in another region.
* A **homogeneous** mixture consists of two or more substances in the same phase. No amount of optical magnification will reveal a homogeneous mixture to have different properties in different regions.
* Homogeneous mixtures are often called **solutions**. Common examples include air (mostly a mixture of nitrogen and oxygen gases), gasoline (a mixture of carbon and hydrogen containing compounds called hydrocarbons), and an unopened soft drink.

**Separation of mixtures :**

The different methods used are as follows :

**(i)** **Filtration** to separate mixture in which one component is soluble in a particular solvent and the other is not.

**(ii)** **Simple distillation** to separate a non-volatile solute from a solution or to separate a mixture of liquids having large difference in their boiling points.

**(iii)** **Fractional distillation** to separate a mixture of liquids having small difference in their boiling points.

**(iv)** **Extraction** to dissolve out one of the components of the mixture with a suitable solvent.

**(v)** **Fractional crystallisation** to separate out two solids having different solubilities in the same solvent.

**(vi)** **Gravity separation** to separate a mixture ir which the components have different densities.

**(vii)** **Magnetic separation** to separate the components of a mixture one of which is magnetic and the other is non-magnetic.

**(viii)** **Sublimation** to separate the components of a mixture one of which sublimes and the other does not.

**(ix)** **Chromatography** to separate the components of a mixture on the basis of their difference in adsorption on a particular adsorbent.

## PROPERTIES OF MATTER

Every substance has unique or characteristic properties. These properties can be classified into two categories - physical properties and chemical properties.

**Physical properties**

* Properties which can be observed and measured without changing the composition of a substance, are called physical properties.
  For example, colour, melting point and boiling point.
* Physical properties allow us to classify and identify substances.

**Chemical properties :**

* They are characteristic reactions of different substances; these include acidity or basicity, combustibility etc.

**Intensive and Extensive properties :**

* **Extensive properties** depend on the amount of a substance present. The mass and volume are extensive properties, for example. In contrast, **intensive properties** do not depend on the amount of substance. A sample of ice will melt at 0°C, no matter whether you have an ice cube or an iceberg. Density is also an intensive property.

**UNCERTAINTY IN MEASUREMENT**

* Two kinds of numbers are encountered in scientific work: exact numbers (those whose values are known exactly) and inexact numbers (those whose values have some uncertainty).
* Numbers obtained by measurement are always inexact. The equipment used to measure quantities always has inherent limitations (equipment errors), and there are differences in how different people make the same measurement (human errors).
* **Precision and Accuracy :** The terms precision and accuracy are often used in discussing the uncertainties of measured values. Precision is a measure of how closely individual measurements agree with one another. Accuracy refers to how closely individual measurements agree with the correct, or "true," value.

**Significant figure**

* Suppose you determine the mass of a coin on a balance capable of measuring to the nearest 0.0001 g. You could report the mass as $2.2405 \pm 0.0001$g.
  The ± notation (read "plus or minus") expresses the magnitude of the uncertainty of your measurement. In much scientific work we drop the notation with the understanding that there is always some uncertainty in the last digit reported for any measured quantity.
* All digits of a measured quantity, including the uncertain one, are called significant figures. A measured mass reported as 2.2 g has two significant figures, whereas one reported as 2.2405 g has five significant figures.
* The greater the number of significant figures, the greater the certainty implied for the measurement.
* To determine the number of significant figures in a reported measurement, read the number from left to right, counting the digits starting with the first digit that is not zero.
* In any measurement that is properly reported, all nonzero digits are significant. Because zeros can be used either as part of the measured value or merely to locate the decimal point, they may or may not be significant:

1. Zeros between nonzero digits are always significant 1005 kg (four significant figures);
   7.03 cm (three significant figures).
2. Zeros at the beginning of a number are never significant; they merely indicate the position of the decimal point 0.02 g(one significant figure); 0.0026cm (two significant figures).
3. Zeros at the end of a number are significant if the number contains a decimal point 0.0200 g (three significant figures); 3.0 cm (two significant figures).

* A problem arises when a number ends with zeros but contains no decimal point. In such cases, it is normally assumed that the zeros are not significant. Exponential notation can be used to indicate whether end zeros are significant. For example, a mass of 10,300 g can be written to show three, four, or five significant figures depending on how the measurement is obtained:
  $1.03 \times 10^4$ g (three significant figures)
  $1.030 \times 10^4$ g (four significant figures)
  $1.0300 \times 10^4$ g (five significant figures)
  In these numbers all the zeros to the right of the decimal point are significant (rules 1 and 3). (The exponential term $10^4$ does not add to the number of significant figures.)

**Significant figures in calculations**

* When carrying measured quantities through calculations, the least certain measurement limits the certainty of the calculated quantity and thereby determines the number of significant figures in the final answer. The final answer should be reported with only one uncertain digit.
* To keep track of significant figures in calculations, we will make frequent use of two rules, one for addition and subtraction, and another for multiplication and division.

1. **For addition and subtraction**, the result has the same number of decimal places as the measurement with the fewest decimal places. When the result contains more than the correct number of significant figures, it must be rounded off.

| | |
|---|---|
| This number limits | 20.42 ← two decimal places |
| the number of significant | 1.322← three decimal places |
| figures in the result → | 83.1 ← one decimal place |
| | 104.842 ; round off to one decimal place (104.8) |

We report the result as 104.8 because 83.1 has only one decimal place.

2. **For multiplication and division,** the result contains the same number of significant figures as the measurement with the fewest significant figures. When the result contains more than the correct number of significant figures, it must be rounded off. For example, the area of a rectangle whose measured edge lengths are 6.221 cm and 5.2cm should be reported as 32 $cm^2$

Area = (6.221 cm) (5.2 cm) = 32.3492 $cm^2$

Round off to 32$cm^2$ because 5.2 has two significant figures.

* Notice that for addition and subtraction, decimal places are counted in determining how many digits to report in an answer, whereas for multiplication and division, significant figures are counted in determining how many digits to report.
* In determining the final answer for a calculated quantity, exact numbers are assumed to have an infinite number of significant figures. Thus, when we say, "There are 12 inches in 1 foot," the number 12 is exact, and we need not worry about the number of significant figures in it.
* In rounding off numbers, look at the left most digit to be removed:
  - If the left most digit removed is less than 5, the preceding number is left unchanged. Thus, rounding 7.248 to two significant figures gives 7.2.
  - If the left most digit removed is 5 or greater, the preceding number is increased by 1.

  Rounding 4.735 to three significant figures gives 4.74, and rounding 2.376 to two significant figures gives 2.4

**DIMENSIONAL ANALYSIS (The factor-label method)**

* It is a general problem-solving approach that uses the dimensions or units of each value to guide you through calculations.
* A conversion factor expresses the equivalence of a measurement in two different units

  (1 cm ≡ 10 mm; 1 g ≡ 1000 mg;

  12 eggs ≡ 1 dozen; 12 inches ≡ 1 foot).
* Because the numerator and the denominator describe the same quantity, the conversion factor is equivalent to the number 1. Therefore, multiplication by this factor does not change the measured quantity, only its units.
* A conversion factor is always written so that it has the form "new units divided by units of original number."

$$\text{Number in original unit}\left[\frac{\text{new unit}}{\text{original unit}}\right] = \text{new number in new unit}$$

↑ Quantity to express in new units — Conversion factor — ↑ Quantity now expressed in new units

**Useful conversion factors**

* 1 Å = $10^{-10}$ m
* 1 nm = $10^{-9}$ m
* 1 pm = $10^{-12}$ m
* 1 litre = $10^{-3}$ m = 1 $dm^3$
* 1 atm = 760 mm or torr=101, 325 Pa or $Nm^{-2}$
* 1 bar = $10^5$ $Nm^{-2}$ = $10^5$ Pa
* 1 calorie = 4.184 J
* 1 electron volt (eV) = 1.6022 × $10^{-19}$ J.

**Example 1 :**

If the value of Avogadro number is 6.023 × $10^{23}$ $mol^{-1}$ and the value of Boltzmann constant is 1.380 × $10^{-23}$ J $K^{-1}$, then find the number of significant digits in the calculated value of the universal gas constant. (R = N × K)

**Sol. 4.** $6.023 \times 10^{23} \times 1.380 \times 10^{-23} = 8.312$

It has four significant figure.

**LAWS OF CHEMICAL COMBINATION**

**(i) The law of conservation of mass :**

(a) This law was given by "Lavoiser" in 1744.

(b) This law states that "matter can neither be created nor destroyed, or in a chemical reaction, the mass of the reactants is equal to the mass of the products".

**Example :**

$$2Mg + O_2 \longrightarrow 2MgO$$
$$2\times 24 \quad 32 \qquad 2(24+16)$$
$$48 \quad 32 \qquad 2(40) = 80$$

Total mass of reactants = Total mass of product = 80

$$C + O_2 \longrightarrow CO_2$$
$$12 \quad 32 \qquad 12+32 = 44$$

Total mass of reactants = Total mass of product = 44

(c) The exception of this law is nuclear reactions where Einstein Equation ($E = mc^2$) is applicable.

**(ii) The law of constant composition or definite proportion :**

(a) This law was given by "Proust" in 1799.

(b) This law states that "All pure samples of the same chemical compound contain the same elements combined in the same proportion by mass".

**Example :** Different Samples of carbondioxide contain carbon and oxygen in the ratio of 12 : 32 or 3 : 8 by mass. Similarly in $H_2O$ ratio of weight of hydrogen to oxygen is 1 : 8.

**(iii) The law of multiple proportion :**

(a) This law was given by John Dalton in 1804.

(b) This law states that "When two elements A and B combine together to form more than one compound, then several masses of A which separately combine with a fixed mass of B, are in a simple ratio."

**Example :** CO and $CO_2$

12 : 16    12 : 32

Ratio = 16 : 32 = 1 : 2

Similarly $H_2O$ and $H_2O_2$

2 : 16    2 : 32

Ratio = 16 : 32 = 1 : 2

**(iv) The law of Reciprocal proportions :**

(a) This law was given by "Richer" in 1792–94.

(b) This law states that "When two elements combines separately with third element and form different types of molecules, their combining ratio is directly reciprocted if they combine directly".

**Example :** $C + O_2 \longrightarrow CO_2$ ,    $C + 2H_2 \rightarrow CH_4$

12   32     12   4

C combines with O to form $CO_2$ and with H to form $CH_4$. In $CO_2$ 12 gm of C reacts with 32 gm of O, where as in $CH_4$ 12 gm of C reacts with 4 gm of H. Therefore when O combines with H, they should combines in the ratio of 32 : 4 (i.e. 8 : 1) or in simple multiple of it. The same ratio is found to be true in $H_2O$ molecules, the ratio of weight of H and O in $H_2O$ is 1 : 8.

**(v) The law of Gaseous volume :**

(a) This law was given by "Gaylussac" in 1808.

(b) This law states that "Whenever gases react together, the volumes of the reacting gases as well as the products if they are gases, bear a simple whole number ratio, provided all the volumes are measured under similar conditions of Temperature & Pressure.

**Example :**

$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$

1 unit vol. 1 unit vol. 2 unit vol.

ratio = 1 : 1 : 2

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

1 unit vol. 3 unit vol. 2 unit vol.

ratio = 1 : 3 : 2

## AVOGADRO HYPOTHESIS

(a) Equal volumes of all gases under similar conditions of Temperature and pressure contain equal number of molecules.

**Example :** $2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$

2 unit vol. 1 unit vol. 2 unit vol.

Ratio of number of molecules = 2 : 1 : 2

(b) This law helped to remove anomaly between Dalton's atomic theory and Gaylussac's law of volume by making a clear distinction between atoms and molecules.

(c) It reveals that common elementary gases like Hydrogen, Nitrogen, Oxygen etc. are diatomic.

(d) It provides a method to determine the molecular weights of gaseous elements.

**Example 2 :**

4.2 g of $NaHCO_3$ on reaction with 10 g of acetic acid, liberated 2.2 g of $CO_2$ gas. Find the mass of residue left.

**Sol.** According to law of conservation of mass

mass of reactants = mass of products.

$4.2 + 10 = x + 2.2$ (where x is mass of residue)

$x = 4.2 + 10 - 2.2 = 12$ g

**Example 3 :**

In one experiment 4g of $H_2$ combine with 32g of $O_2$ to form 36g of $H_2O$. In another experiment when 50g of $H_2$ combine with 400g of $O_2$ then 450g of $H_2O$ is formed. Above two experiment follows

(1) The law of conservation of mass

(2) The law of constant composition

(3) The law of definite proportion

(4) All of these

**Sol.** **(4).** I experiment : $\dfrac{\text{mass of } H_2 \text{ combined}}{\text{mass of } O_2 \text{ combined}} = \dfrac{4}{32} = \dfrac{1}{8}$

II experiment :

$$\dfrac{\text{mass of } H_2 \text{ combined}}{\text{mass of } O_2 \text{ combined}} = \dfrac{500}{400} = \dfrac{1}{8}$$

Hence both law of conservation of mass and constant composition is obeyed.

## ATOM AND MOLECULES

**Atom :** It is the smallest particle of an element that takes part in a chemical reaction and is not capable of independent existence.

**Molecule :** It is the smallest particle of matter which is capable of independent existence.

A molecule is generally an assembly of two or more tightly bonded atoms. Molecules are of two type on the basis of elemental atoms.

**(i) Homo atomic molecules :** Molecules of an element containing one type of atoms only. **Ex. :** $O_2$, $H_2$, $Cl_2$ etc.

**(ii) Heteroatomic molecules :** Molecules of compounds containing more than one type of atoms.

Ex. : $NH_3$, $H_2O$, $CH_4$ etc.

**Atomic Mass :** Carbon as standard : The modern reference standard for atomic weight is carbon isotope of mass number 12. It is the number of times an atom of an element is heavier than 1/2 th of an atom of C–12.

$$\text{At. wt. of an element} = \frac{\text{Weight of 1 atom of element}}{1/12 \times \text{weight of 1 atom of C-12}}$$

* 1 amu (Atomic mass unit) = $1.66056 \times 10^{-24}$ g

* Mass of an atom of hydrogen = $1.6736 \times 10^{-24}$ g

In terms of amu, the mass of hydrogen atom

$$= \frac{1.6736 \times 10^{-24}\,g}{1.66056 \times 10^{-24}\,g} = 1.0078 \text{ amu} = 1.0080 \text{ amu}$$

Similarly, the mass of oxygen-16 ($^{16}O$) atom would be 15.995 amu.

Now a days, 'amu' has been replaced by '**u**' which is known as **unified mass.**

**Note :**

1. Atomic mass is not a weight but a unitless number.
2. Atomic mass is not absolute but relative to the weight of the standard reference elements C–12.
3. **Dulong's and Petit's law :** In case of heavy solid elements, it is observed that product of atomic mass and specific heat capacity is almost constant.
   Atomic mass × Specific heat capacity (Cal/gm°C) ≈ 6.4
   It should be remembered that this law is an empirical observation and this gives an approximate value of atomic mass.
4. Atoms of the same element which have different relative masses are called isotopes.
5. In case of isotopes, atomic mass of the element is average of relative masses of different isotope of the element

**Example :** There are two isotopes of chlorine

| | $_{17}C^{35}$ | and | $_{17}C^{37}$ |
|---|---|---|---|
| Relative mass | 35 | | 37 |
| Relative abundance (RA) | 3 | : | 1 |

Atomic mass of chlorine

$$= \frac{(\text{At. mass of I isotope} \times \text{RA}) + (\text{At. mass of II isotope} \times \text{RA})}{\text{Total RA}}$$

$$= \frac{(35 \times 3) + (37 \times 1)}{4} = 35.5$$

6. When atomic mass of elements are expressed in grams. They are called Gram Atomic Mass (GAM).
   1 GAM of Na = 23 gm of Na
   1 GAM of Ca = 40 gm of Ca

**Formula and Molecular Weights**

* The formula weight of a substance is the sum of the atomic weights of the atoms in the chemical formula of the substance. Using atomic weights, we find, for example, that the formula weight of sulphuric acid ($H_2SO_4$) is 98.1 amu.
  FW of $H_2SO_4$ = 2 (AW of H) + (AW of S) + 4 (AW of O)
  = 2 (1.0 amu) + 32.1 amu + 4 (16.0 amu) = 98.1 amu
* If the chemical formula is the chemical symbol of an element, such as Na, the formula weight equals the atomic weight of the element, in this case 23.0amu.
* If the chemical formula is that of a molecule, the formula weight is also called the molecular weight. The molecular weight of glucose ($C_6H_{12}O_6$). For example, is
  MW of $C_6H_{12}O_6$ = 6 (12.0 amu) + 12 (1.0 amu) + 6 (16.0 amu) = 180.0 amu
  Because ionic substances exist as three-dimensional arrays of ions, it is inappropriate to speak of molecules of these substances. Instead, we speak of formula units. The formula unit of NaCl, for instance, consists of one $Na^+$ ion and one $Cl^-$ ion. Thus, the formula weight of NaCl is defined as the mass of one formula unit:
  FW of NaCl = 23.0 amu + 35.5 amu=58.5 amu
* When molecular mass of compounds are expressed in grams. They are called Gram molecular mass (GMM) or (GMW).

**Determination of molecular weight :**

1. **Vapour density method :**

$$\text{Vapour density} = \frac{\text{Wt. of a certain vol. of a gas or vapour under certain temperatrue and pressure}}{\text{Wt. of the same volume of } H_2 \text{ under same temperature and pressure}}$$

Molecular weight = 2 × vapour density

2. **Diffusion method :**
   (a) It is based on Grahm's law of diffusion.
   (b) Grahm's law states that "The rate of diffusion of different gases, under similar conditions of temperature and pressure are inversely proportional to the square roots of their density (or molecular weights)".

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}}$$

When molecular mass of compounds are expressed in grams. They are called Gram molecular mass (GMM) or (GMW).

**PERCENTAGE COMPOSITION**

* Percentage composition of the compound is the relative mass of the each of the constituent composition in 100 parts of it. Mass % of an element

$$= \frac{\text{mass of that element in the compound} \times 100}{\text{molar mass of the compound}}$$

* Let us take an example of ethanol :
  Molecular formula of ethanol is : $C_2H_5OH$
  Molar mass of ethanol is :
  (212.01 + 61.008 + 16.00) g = 46.068 g

$$\text{Mass percent of carbon} = \frac{24.02g}{46.068g} \times 100 = 52.14\%$$

$$\text{Mass percent of hydrogen} = \frac{6.048g}{46.068g} \times 100 = 13.13\%$$

$$\text{Mass percent of oxygen} = \frac{16.00g}{46.068g} \times 100 = 34.73\%$$

**CHEMICAL FORMULA**

It is of two types :

(a) **Molecular formulae :** Chemical formulae that indicate the actual number and type of atoms in a molecule called molecular formulae.
**Example :** Molecular formula of Benzene is $C_6H_6$

(b) **Empirical formulae :** Chemical formulae that indicate only the relative number of atoms of each type in a molecule are called empirical formulae.
**Example :** Empirical formulae of Benzene in "CH".

**Determination of Chemical formulae :**

(a) **Determination of Empirical formulae :**
Step (I) : Determination of percentage of each element
Step (II) : Determination of mole ratio
Step (III) : Making it whole number ratio
Step (IV) : Simplest whole ratio

**Example 4 :**

Phosgene, a poisonous gas used during World war-I, contains 12.1% C, 16.2% O and 71.7% Cl by mass. What is the empirical formula of phosgene.

**Sol.**

| Element | % | Mole ratio | Simplest mole ratio |
|---|---|---|---|
| C | 12.1 | $\frac{12.1}{12} = 1.01$ | $\frac{1.01}{1.01} = 1$ |
| O | 16.2 | $\frac{16.2}{16} = 1.01$ | $\frac{1.01}{1.01} = 1$ |
| Cl | 71.7 | $\frac{71.7}{35.5} = 2.02$ | $\frac{2.02}{1.01} = 2$ |

Then empirical formulae = $COCl_2$

(b) **Determination of molecular formulae**
Step (I) : First of all find Empirical formulae
Step (II) : Calculate the Empirical weight
Step (III) : Molecular formulae = $(\text{Empirical formulae})_n$

$$n = \frac{\text{Molecular weight}}{\text{Empirical weight}}$$

**Example 5 :**

5.325 g sample of methyl bexzoate, a compound used in the manufacture of perfumes is found to contain 3.758 g of carbon 0.316 g hydrogen and 1.251g of oxygen. What is empirical formulae, of compound. If molecular weight of methyl benzoate is 136.0, calculate its molecular formula.

**Sol.**

| Element | % | Mole ratio | Simplest whole ratio |
|---|---|---|---|
| C | $\frac{3.758\times100}{5.325}=70.57$ | $\frac{70.57}{12}=5.88$ | $\frac{5.88}{147}=4$ |
| H | $\frac{0.316\times100}{5.325}=5.93$ | $\frac{5.93}{1}=5.93$ | $\frac{5.93}{147}=4$ |
| O | $\frac{1.251\times100}{5.325}=23.50$ | $\frac{23.50}{16}=1.47$ | $\frac{1.47}{1.47}=1$ |

Empirical formula = $C_2H_4O$

$$n=\frac{\text{Mol. wt}}{\text{Empirical formula wt}}=\frac{136}{68}=2$$

$\Rightarrow$ Molecular formula = $C_8H_8O_2$

## TRY IT YOURSELF - 1

**Q.1** How many significant figures are there in each of the following?
(i) $6.02\times10^{23}$ (ii) 21.00g (iii) 0.0362m

**Q.2** Express the following up to four significant figures.
(i) 55.0546 (ii) $1.7546\times10^{10}$ (iii) 0.002343

**Q.3** Express the results of the following to the appropriate number of significant figures:
(i) $\frac{32.4\times0.0867}{4.238}$ (ii) $0.42+452.32$

**Q.4** Boron has two isotopes Boron-10 and Boron-11 whose percentage abundance are 19.6% and 80.4% respectively? What is the average atomic mass of Boron?

**Q.5** Weight of CuO obtained by treating 2.16g Cu with nitric acid and subsequent ignition was 2.70g. In another experiment 1.15g of CuO on reduction yielded 0.92g Cu. Show that the data illustrates the law of definite proportions.

**Q.6** Calculate the volume of $H_2$ required to prepare 20 litre of $NH_3$, according to given reaction.
$N_2+3H_2\rightarrow2NH_3$

**Q.7** The relative abundance of naturally occurring element M is 78.99% $M^{24}$, 10% $M^{25}$ and 11.01% $M^{26}$ the atomic weight of M will be –

**Q.8** Aluminium oxide contains 52.9% aluminium and carbon dioxide contains 27.27% carbon. Assuming that the law of reciprocal proportions is true, calculate the percentage of aluminium in aluminium carbide.

**Q.9** Carbon occurs in nature as a mixture of C-12 and C-13. The average atomic mass of carbon is 12.011. What is the percentage abundance of Carbon-12 in nature?

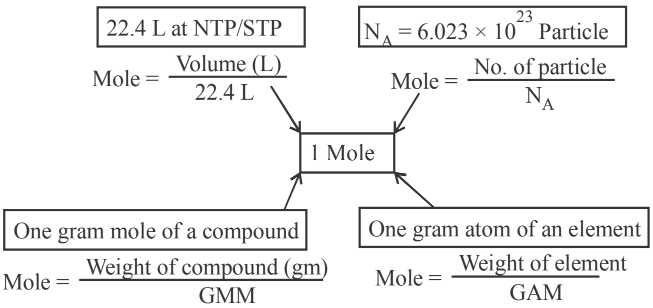
**Q.10** A compound on analysis gave the following results C = 54.54%, H = 9.09% and vapour density of the compound = 88. Determine the molecular formula of the compound.

**ANSWERS**

**(1)** (i) 3 (ii) 4 (iii) 3
**(2)** (i) 55.05 (ii) $1.755\times10^{10}$ (iii) 0.002343
**(3)** (i) 0.663, (ii) 452.74 **(4)** 10.804 u **(6)** 30 litre
**(7)** 24.32 **(8)** 74.97% **(9)** 98.9%
**(10)** $C_8H_{16}O_4$

## MOLE CONCEPT

**Mole :** Mole is a unit which represent $6.023\times10^{23}$ particles of same nature.




1 Mole = $6.023\times10^{23}$ Particles.
1 Mole atom = $6.023\times10^{23}$ atoms.
1 Mole molecule = $6.023\times10^{23}$ Molecules
1 Mole Electron = $6.023\times10^{23}$ Electrons.
The number $6.023\times10^{23}$ is called Avogadro number : ($N_A$)

**Relation of mole with mass :**

Mass of one mole atoms of an element = (Atomic mass of element) gm = Gram Atomic Mass (GAM)

**Example :** Mass of 1 Mole atoms of carbon = GAM of C=12gm

Mass of 1 Mole atoms of oxygen = GAM of O = 16 gm

Mass of 1 mole molecules of substance
= (Molecular weight of substance) gm
= Gram molecular mass (GMM)

**Ex.** Mass of 1 Mole Molecules of $O_2$ = GMM of $O_2$=32gm.
Mass of 1 Mole Molecules of $CO_2$ = GMM of $CO_2$ = 44gm

**Relation of mole with gas volume :**

Ideal gas equation : PV = nRT
Where P = Pressure of gas, V = Volume of gas,
n = Number of moles of gas, T = Temperature(Kelvin),
R = Gas constant = 0.082 Atm Ltr $K^{-1}$ $mole^{-1}$
Volume of one mole of a gas at NTP = 22.4 Litre.
Since 1 mole gas contain $6.023\times10^{23}$ molecules
$6.023\times10^{23}$ ($N_A$) molecule have volume at NTP = 22.4 lit.

$$\text{Moles at NTP}=\frac{\text{Volume (Litre)}}{22.4}$$

**Example 6 :**

Calculate the number of gram atoms for $2\times10^{23}$ atoms. If atomic weight of element is 24, then calculate mass of the atoms.

**Sol.** No. of gram atoms of element

$$= \frac{X}{N_A} = \frac{2\times10^{23}}{6.023\times10^{23}} = 0.33 \text{ moles.}$$

Mass of $2 \times 10^{23}$ atoms $= n \times GAM = 0.33 \times 24 = 7.92$ gm

**Example 7 :**

Calculate total number of moles of atoms present in 49 gm $H_2SO_4$.

**Sol.** Number of gram mole for 49 gm $H_2SO_4 = \frac{w}{GMM} = \frac{49}{98}$

$= 0.5$ mole

Since atomicity of $H_2SO_4 = 7$

Then total number of moles of atoms $= n \times$ atomicity

$= 0.5 \times 7 = 3.5$ moles

## CHEMICAL REACTION & MOLE CONCEPT

The ratio between reactant and product molecules is same to the ratio of their moles and volumes (gaseous substance) at NTP in a balanced chemical equation.

**Example :** $Fe_2(SO_4)_3 + BaCl_2 \longrightarrow BaSO_4 + FeCl_3$

Balanced equation : $Fe_2(SO_4)_3 + 3BaCl_2 \rightarrow 3BaSO_4 + 2FeCl_3$

| | | | | |
|---|---|---|---|---|
| Ratio of molecules | 1 | : 3 | : 3 | : 2 |
| Their moles's ratio | 1 | : 3 | : 3 | : 2 |

**Example :** $2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$

| | | | |
|---|---|---|---|
| Molecules ratio = | 2 | : 1 | : 2 |
| Mole's ratio = | 2 | : 1 | : 2 |
| Volume's ratio = | 2 | : 1 | : 2 |

## LIMITING REAGENT

It may be defined as the reactant which is completely consumed during the reaction is called limiting reagent. A reactant that is not completely consumed is often referred to as an excess reactants. Once one of the reactant is used up, the reaction stops. The moles of product are always determined by the starting moles of limiting reactants.

**Calculation of limiting reagent :**

**Method I :** By calculating the required amount by the equation and comparing it with given amount.
[Useful when only two reactant are there]

**Method II:** By calculating amount of any one product obtained taking each reactant one by one irrespective of other reactants. The one giving least product is limiting reagent.

**Method III :** Divide given moles of each reactant by their stoichiometric coefficient, the one with least ratio is limiting reagent. [Useful when number of reactants are more than two].

## PRODUCT YIELD

It is not very uncommon that the actual yield of a product is less than the theoretical maximum yield. The percentage yield of the product is defined as,

$$\% \text{ yield of product} = \frac{\text{Actual yield}}{\text{Theoretical maximum yield}} \times 100$$

## PERCENTAGE PURITY

Depending upon the mass of the product, the equivalent amount of reactant present can be determined with the help of given chemical equation. Knowing the actual amount of the reactant taken and the amount calculated with the help of a chemical equation, the purity can be determined, as,

$$\% \text{ purity} = \left[\frac{\text{Amount of reactant calculated from the chemical equation}}{\text{Actual amount of reactant taken}}\right] \times 100$$

**Example 8 :**

Calculate the mass of oxygen required to burnt 14g $C_2H_4$ completely :

**Sol.** $C_2H_4 + 3O_2 \longrightarrow 2CO_2 + 2H_2O$

Mole ratio 1 3 2 4

Moles of $C_2H_4$ to be burnt $= \frac{14}{28} = \frac{1}{2}$ mole.

$\because$ 1 mole $C_2H_4$ requires 3 mole $O_2$ for combustion

$\therefore \frac{1}{2}$ mole $C_2H_4$ requires $3 \times \frac{1}{2}$ mole $O_2 = \frac{3}{2}$ mol $O_2$

Mass of Oxygen $= 3/2 \times 32 = 48$ gm.

**Example 9 :**

If 20gm of $CaCO_3$ is treated with 20gm of HCl, how many grams of $CO_2$ can be generated according to following reactions?

$$CaCO_3(g) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(\ell) + CO_2(g)$$

**Sol.** Mole of $CaCO_3 = 20/100 = 0.2$

Mole of HCl $= \frac{20}{36.5} = 0.548$

$\left[\frac{\text{Mole}}{\text{Stoichiometric coefficient}}\right]$ for $CaCO_3 = \frac{0.2}{1} = 0.2$

$\left[\frac{\text{Mole}}{\text{Stoichiometric coefficient}}\right]$ for HCl $= \frac{0.548}{2} = 0.274$

According to reaction,

100 gm of $CaCO_3$ gives 44gm of $CO_2$

20gm $CaCO_3$ will give $\frac{44}{100} \times 20 = 8.8$gm $CO_2$

**Example 10 :**

Calculate the amount of (CaO) in kg that can be produced by heating 200 kg lime stone that is 90% pure $CaCO_3$.

**Sol.** Mass of pure $CaCO_3 = \frac{200\times90}{100} = 180$ kg

$CaCO_3 \rightarrow CaO + CO_2$

100 kg 56kg

180 x

$\frac{100}{180} = \frac{56}{x} \Rightarrow x = 100.8$ kg.

**Example 11 :**

If 6 moles of $KClO_3$ are decomposed according to following reactions calculate the moles of $KClO_4$ produced if mole of $O_2$ produced are 3 ?

$$KClO_3 \rightarrow KCl + \frac{3}{2}O_2 \quad ........ (1)$$

$$4KClO_3 \rightarrow 3KClO_4 + KCl \quad ........ (2)$$

**Sol.** Let x mole $KClO_3$ reacts in reaction (1) and y mole $KClO_3$ reacts in reaction (2)

$KClO_3 \rightarrow KCl + (3/2)\,O_2$ ........ (1)
x mole (3x/2) mole
$4KClO_3 \rightarrow 3KClO_4 + KCl$ ........ (2)
y mole

From question, $x + y = 6$ and $\frac{3x}{2} = 3$

$\therefore$ x = 2 mole and y = 4 mole

It means 4 mole $KClO_3$ reacts in reaction (2).

From eq. (1), 4 mole $KClO_3$ gives 3 mole $KClO_4$.

**PRINCIPLE OF ATOM CONSERVATION**

The principle of conservation of mass, expressed in the concepts of atomic theory means the conservation of atoms. And if atoms are conserved, moles of atoms shall also be conserved. This is known as the principle of atom conservation. This principle is in fact the basis of the mole concept.

**Example 12 :**

All carbon atoms present in $KH_3(C_2O_4)_2.2H_2O$ weighing 254gm is converted to $CO_2$. How many gram of $CO_2$ were obtained?

**Sol.** Applly POAC on carbon atom

$$4 \times \text{atom of } KH_3(C_2O_4)_2.2H_2O = 1 \times \text{mole of } CO_2$$

$$4 \times \frac{254}{254} = 1 \times \frac{W_{CO_2}}{44}$$

$\therefore$ Mass of $CO_2 = 4 \times 44 = 176$ gram

**STOICHIOMETRY AND PROBLEM SOLVING**

Stoichiometry refers to the quantitative relationship between the reactants and the products. It is quite useful in calculating the amount of the reactants required or those of the products formed for the chemical process. The calculations based on the knowledge of chemical equations are also called **Stoichiometry calculations steps :**

(i) Write the balanced chemical equation.

(ii) Write the molar relationship from the equation between the given and the required species.

(iii) Convert these moles into the desired parameters such as mass, volume, etc.

(iv) Apply unitary method to calculate the result.

**Example 13 :**

Calculate the volume of chlorine that can be obtained at STP, by reaction of 1.58 g of $KMnO_4$ and excess of hydrochloric acid.

$$2KMnO_4 + 16\,HCl \rightleftharpoons 2KCl + 2MnCl_2 + 8H_2O + 5Cl_2$$

2 mol 5 mol
(316 g) ($5 \times 22.4$ L at STP)

Thus, volume of $Cl_2$ produced at STP

$$= \frac{5 \times 22.4 \times 1.58}{316} = 0.560 \text{ or } 560 \text{ mL}$$

## TRY IT YOURSELF - 2

**Q.1** Calculate the actual mass of one molecule of carbondioxide ($CO_2$)

**Q.2** How many moles of $H_2SO_4$ are present in 4.9g $H_2SO_4$ ?

**Q.3** Calculate the weight of oxygen produced by the thermal decomposition of 10g of potassium chlorate.

**Q.4** Calculate the number of molecules in 1ml of $O_2$ at NTP.

**Q.5** Calculate the volume occupied at NTP by
(i) 2.5 mole of carbon dioxide
(ii) 14g of nitrogen gas

**Q.6** Calculate the number of moles of $Na_2SO_4$ produced from 1 mole of NaOH.

$$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$$

**Q.7** An enzyme contains 5.6% Fe, calculate number of Fe atoms present in 1g of enzyme.

**Q.8** Calcualte the amount of 50% $H_2SO_4$ required to decompose 25g of Marble (Calcium carbonate)

**Q.9** Calcualte the actual mass of a water molecule in gram.

**Q.10** Calculate volume of carbon dioxide produced on heating 10g of lime stone.

**Q.11** Calculate mass of $CO_2$ produced by heating 40g of 20% pure lime stone.

**Q.12** How many moles of lead nitrate is needed to produce 224 litre of oxygen at NTP?

**Q.13** Oxygen is prepared by catalytic decomposition of potassium chlorate, $KClO_3$, Decomposition of potassium chlorate gives potassium chloride (KCl) and oxygen ($O_2$). If 2.4 mol of oxygen is needed for an experiment, how many grams of potassium chlorate must be decomposed?

**Q.14** 10g of hydrogen is reacted with 50g of oxygen. What is the amount of water produced? Calculate the amount of unreacted reagent also.

**Q.15** Calculate the number of formula units, number of oxygen atom and total charge in 3gm $CO_3^{--}$.

**ANSWERS**

**(1)** $7.304 \times 10^{-23}$g **(2)** 0.05 **(3)** 3.92 g
**(4)** $2.69 \times 10^{19}$ molecules **(5)** (i)56L,(ii) 11.2L
**(6)** 1/2 mole **(7)** $6.02 \times 10^{20}$ atoms
**(8)** 49g **(9)** $2.99 \times 10^{-23}$g **(10)** 2.24 litre
**(11)** 3.52g **(12)** 20 mol **(13)** 196.0g
**(14)** 3.75g
**(15)**(i) $3.0 \times 10^{22}$ (ii) $9.0 \times 10^{22}$ (iii) $9.6 \times 10^3$ coulomb]

## EQUIVALENT WEIGHT

Equivalent weight of a substance (element or compound) is defined as "The number of parts by weight of it, that will combine with or displace directly or indirectly 1.008 parts by weight of hydrogen, 8 parts by weight of oxygen, 35.5 parts by weight of chlorine or the equivalent parts by weight of another element".

Equivalent weight of substance depends on the reaction in which that take parts.

Equivalent weight is a relative quantity so it is unit less. When equivalent weight of substances are expressed in grams. They are called Gram equivalent weight (GEW).

**n Factor :**

Equivalent weight is the ratio of atomic weight and a factor (say n-factor).

$$\text{Equivalent weight} = \frac{\text{atomic weight}}{\text{n-factor}}$$

In case of acid/base the n-factor is basicity/acidity (i.e. number of dissociable $H^+$ ions/number of dissociable $OH^-$ ion and in case of oxidizing agent/reducing agent, n-factor is number of moles of electrons gained/lost per mole of oxidizing agent/reducing agent. Therefore, in general, we can write.

$$\text{Equivalent weight (E)} = \frac{\text{atomic or molecular weight}}{\text{n-factor}}$$

$$\text{No. of equivalents of solute} = \frac{\text{wt}}{\text{eq. wt}} = \frac{W}{E} = \frac{W}{M/n}$$

No. of equivalents of solute

= No. of moles of solute × n-factor

**Evaluation of equivalent weight :**

1. $$\text{Equivalent weight of element} = \frac{\text{Atomic weight of element}}{\text{Valency of element}}$$

2. Equivalent weight of salt

= Equivalent weight of I part + Equivalent weight of II part

or $$\frac{\text{Molecular weight of salt}}{\text{Total charge on cation}}$$

**Example 14 :**

Equivalent weight of $AlCl_3$ ?

**Sol.** Equivalent weight $AlCl_3$ $= E_{Al} + E_{Cl}$

$= 27/3 \quad 35.5$

$= 9 + 35.5 = 44.5$

$$E_{AlCl_3} = \frac{\text{Molecular weight of } AlCl_3}{\text{Total positive charge}} = \frac{133.5}{3} = 44.5$$

3. **Equivalent weight of acid salt :**

$$\text{Equivalent weight of acid salt} = \frac{\text{Molecular weight of acid salt}}{\text{Replaceable 'H' atom}}$$

**Ex. :** E of $NaHCO_3$

$$E_{NaHCO_3} = \frac{M_{NaHCO_3}}{\text{Replaceable H atom}} = \frac{84}{1} = 84$$

**Ex. :** E of $NaHSO_4$

$$E_{NaHSO_4} = \frac{M_{NaHSO_4}}{\text{Repleaceable H atom}} = \frac{120}{1} = 120$$

4 **Equivalent weight of basic salt :**

$$\text{Eq. wt. of basic salt} = \frac{\text{Molecular weight of basic salt}}{\text{Replaceable 'OH' groups}}$$

| **Example : Basic Salt** | **Equivalent weight** |
|---|---|
| $Pb(OH)NO_3$ | 286/1 = 286 |
| $Mg(OH)Cl$ | 76.5/1 = 76.5 |

5. $$\text{Eq. wt. of Radicals/ions} = \frac{\text{Formula weight of ion}}{\text{Charge on ion}}$$

6. **Equivalent weight of oxidising and reducing agent :**

"Equivalent weight of an **oxidising** or **reducing agent** is equal to its molecular weight divided by the number of electrons gained or lost by per molecule".

$$\text{Eq. wt. of oxidant} = \frac{\text{Molecular weight of oxidant}}{\text{Number of } e^-s \text{ gained by 1 molecule}}$$

$$\textbf{Eq. wt of reductant} = \frac{\text{Molecular weight of Reductant}}{\text{Number of } e^-s \text{ lost by 1 molecule}}$$

**Example 15 :**

$Zn + FeSO_4 \longrightarrow ZnSO_4 + Fe$

Calculate the Eq. wt. of Zn and $FeSO_4$.

**Sol.** $Zn \longrightarrow ZnSO_4$

$Zn^o\ Zn^{+2}$ (oxidation)

$$E_{Zn} = \frac{\text{Atomic weight of Zn}}{\text{No. of } e^- \text{ lost by Zn}} = \frac{65}{2} = 32.5$$

$FeSO_4 \longrightarrow Fe\ ;\ Fe^{+2} \longrightarrow Fe^o$

$$E_{FeSO_4} = \frac{\text{Molecular weight of } FeSO_4}{\text{Number of } e^-\text{s gained by } FeSO_4} = \frac{152}{2} = 76$$

**Table Oxidising Agents (OA)/Reducting Agents (RA) with Eq. wt.**

| Species | Changes to | Reaction | Electrons exchanged or Change in O.N. | Eq.wt. |
|---|---|---|---|---|
| 1. $MnO_4^-$ (O.A.) | $Mn^{2+}$ in acidic medium | $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ | 5 | $E = \frac{M}{5}$ |
| 2. $MnO_4^-$ (O.A.) | $MnO_2$ in neutral medium | $MnO_4^- + 3e^- + 2H_2O \rightarrow MnO_2 + 4OH^-$ | 3 | $E = \frac{M}{3}$ |
| 3. $MnO_4^-$ (O.A.) | $MnO_4^{2-}$ in strongly basic medium | $MnO_4^- + e^- \rightarrow MnO_4^{2-}$ | 1 | $E = \frac{M}{1}$ |
| 4. $Cr_2O_7^{2-}$ (O.A.) | $Cr^{3+}$ in acidic medium | $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$ | 6 | $E = \frac{M}{6}$ |
| 5. $MnO_2$(O.A.) | $Mn^{2+}$ in acidic medium | $MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$ | 2 | $E = \frac{M}{2}$ |
| 6. $Cl_2$ (O.A.) | $Cl^-$ | $Cl_2 + 2e^- \rightarrow 2Cl^-$ | 2 | $E = \frac{M}{2}$ |
| 7. $CuSO_4$ (O.A.) (in iodometric titration) | $Cu^+$ | $Cu^{2+} + e \rightarrow Cu^+$ | 1 | $E = \frac{M}{1}$ |
| 8. $S_2O_3^{2-}$ (R.A.) | $S_4O_6^{2-}$ | $2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2e^-$ | 2 (for two molecules) | $E = \frac{2M}{2} = M$ |
| 9. $H_2O_2$ (O.A.) | $H_2O$ | $H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$ | 2 | $E = \frac{M}{2}$ |
| 10. $H_2O_2$ (O.A.) | $O_2$ | $H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$ (O.N. of oxygen in $H_2O_2$ is (–1) per atom) | 2 | E = M/2 |
| 11. $Fe^{2+}$ (R.A.) | $Fe^{3+}$ | $Fe^{2+} \rightarrow Fe^{3+} + e^-$ | 1 | E = M/1 |

**7. Equivalent weight of acid & base :**

**(i) Equivalent weight of acid :** Equivalent weight of an acid is weight which contains one gram equivalent weight of replacable hydrogen atoms.

[The number of maximum replaceable hydrogen atoms present in a molecule is called the basicity of the acid.]

$$\text{Eq. wt. of an acid} = \frac{\text{Molecular weight of the acid}}{\text{Basicity of the acid}}$$

| **Acid** | HCl | $H_2SO_4$ | $H_3PO_4$ |
|---|---|---|---|
| **Basicity** | 1 | 2 | 3 |
| **E. W.** | $\frac{M}{1} = 36.5$ | $\frac{M}{2} = 49$ | $\frac{M}{3} = 32.66$ |

| **Acid** | $H_3PO_3$ | $H_3PO_2$ | $HClO_4$ |
|---|---|---|---|
| **Basicity** | 2 | 1 | 1 |
| **E. W.** | $\frac{M}{2} = 41$ | $\frac{M}{1} = 66$ | $\frac{M}{1} = 100.5$ |

**(ii) Equivalent weight of base :** Equivalent weight of base is weight which contains one gram equivalent weight of replaceable hydroxyl radicals.

[The number of maximum replaceable hydroxyl (OH) groups present in the molecule of a base is called the ACIDITY of the base]

$$\text{Eq. wt. of Base} = \frac{\text{Molecular weight of the base}}{\text{Acidity of the base}}$$

| NaOH | $Ca(OH)_2$ | $Al(OH)_3$ |
|---|---|---|
| 1 | 2 | 3 |
| $\frac{M}{1} = 40$ | $\frac{M}{2} = 37$ | $\frac{M}{3} = 26$ |

**Method of determination of equivalent weight :**

**(1) By hydrogen displacement :** Equivalent weight of metals like Ca, Zn, Sn, Mg etc. which react with dilute acids to produce hydrogen can be determined by this method.

$$\text{Eq. wt. of metal} = \frac{\text{Weight of metal taken}}{\text{Weight of displaced } H_2\text{, gas at NTP}} \times 1.008$$

**(2) By oxide formation :**
Equivalent weight of metals like copper, magnessium, mercury, zinc etc. which form their oxides relatively easily, can be determined by this method.

$$\text{Eq. wt. of metal} = \frac{\text{Weight of metal taken}}{\text{Weight of oxygen } (w_2 - w_1)} \times 8$$

($w_2$ = weight of metal oxide, $w_1$ = weight of pure metal)

**(3) By metal chloride formation :**
Equivalent weight of metals like Na, K, Ag, Au etc. which form their chlorides easily can be determined by this method.

$$\text{Eq. wt. of metal} = \frac{\text{Weight of metal taken}}{\text{Weight of chlorine}} \times 35.5$$

**(4) By metal displacement :**
More active metal can displace less active metal from their salt solution.This displacement based on the law of equivalent.

$$\frac{\text{Weight of metal A}}{\text{Weight of metal B}} = \frac{\text{Equivalent weight of metal A}}{\text{Equivalent weight of metal B}}$$

**(5) By Electrolysis :** This method based on Faraday's second law of electrolysis.
The law states – "When same quantity of electricity is passed through the solutions at different electrolytes, the weight of different substances liberated as a result of electrolysis, are in the ratio of their equivalent weights".

$$\frac{\text{Weight of A deposited}}{\text{Weight of B deposited}} = \frac{\text{Equivalent weight of A}}{\text{Equivalent weight of B}}$$

**Example 16 :**
On reaction of 1 gm metal with dil. $H_2SO_4$, displaced hydrogen is 922 $cm^3$ at NTP. What will be equivalent weight of metal ?

**Sol.** Volume of displaced hydrogen at NTP = 922 $cm^3$
weight of Hydrogen gas = 922 × 0.00009 gm = 0.0829 gm

$$\text{Eq. wt. of metal} = \frac{\text{weight of metal taken}}{\text{weight of } H_2 \text{ gas displaced at NTP}} \times 1.008$$

$$= \frac{1}{0.0829} \times 1.008 = 12.147$$

**Example 17 :**
5 gm of a metal give 6.35 gm of its oxide. Calculate the equivalent weight of metal.

**Sol.** Weight of metal = 5 gm, Weight of oxide = 6.35 gm
∴ weight of oxygen = 6.35 – 5 gm = 1.35 gm

$$\text{Eq. wt. of metal} = \frac{\text{weight of metal}}{\text{weight of oxygen}} \times 8 = \frac{5}{1.35} \times 8 = 29.63$$

So equivalent weight of metal is 29.63.

**EXPRESSION OF CONCENTRATION OF SOLUTION**

"The amount of solute which dissolved in unit volume of solution is called concentration of solution".

$$\text{Concentration} = \frac{\text{Amount of solute}}{\text{Volume of solution}}$$

**(i) Weight-weight age percent (w/W) :**
Weight of solute present in 100 gm of the solution.

$$\text{Weight percent} = \frac{\text{weight of solute (gm)}}{\text{weight of solution (gm)}} \times 100$$

$$\%\text{ by weight} = \frac{w}{W} \times 100$$

**Example 18 :**
What is the weight percentage of urea solution in which 10gm urea dissolved in 90gm of water.

**Sol.** $\text{Wt. \% of urea} = \frac{\text{Weight of urea}}{\text{weight of solution}} \times 100 = \frac{10}{90+10} \times 100$

$= 10\%$ urea $\text{sol}^n$ (w/W)

**(ii) Volume-volume percent (v/V) : (In liquid-liquid solution)**
Volume of solute in ml. present in 100 ml of the solution is called volume – volume percentage.

$$\text{Volume – volume \%} = \frac{\text{Volume of solute (ml.)}}{\text{Volume of solution (ml.)}} \times 100$$

$$\%\text{ by volume} = \frac{v}{V} \times 100$$

**Example 19 :**
A solution is prepared by mixing of 10 ml ethanol with 190 ml of water. What is volume percentage of ethanol

**Sol.** Volume percentage of ethanol

$$= \frac{\text{Volume of ethanol}}{\text{Volume of solution}} \times 100 = \frac{10}{10+190} \times 100 = 5\%$$

Thus 5% ethanol aqueous solution.

**(iii) Weight – volume percentage (w/V) :**
Weight of solute in gm. present in 100 ml of the solution is called weight — volume percentage.

$$\text{weight – volume \%} = \frac{\text{weight of solute (gm)}}{\text{volume of solution (ml)}} \times 100$$

$$\%\text{ of strength} = \frac{w}{V} \times 100$$

**(iv) Normality :** The number of gram equivalents of the solute dissolved per litre of the solution.
It is denoted by 'N'.

$$\text{Normality} = \frac{\text{Number of gram equivalents of solute}}{\text{volume of solution (lit.)}}$$

$\because$ $\text{Gram eq. of solute} = \dfrac{\text{weight of solute (gm)}}{\text{Equivalent weight of solute}}$

$\therefore$ $\text{Normality} = \dfrac{\text{weight of solute (gm)}}{\text{Equivalent weight of solute}} \times \dfrac{1}{\text{volume of solution (lit.)}}$

**Formula :** $N = \dfrac{w}{E} \times \dfrac{1}{V(\text{lit.})}$ ......... (i)

$N = \dfrac{w}{E} \times \dfrac{1000}{V(\text{ml})}$ ......... (ii) $N = n_E \times \dfrac{1}{V(\text{lit})}$ ......... (iii)

$n_E = N \times V$ ......... (iv)

w = weight of solute (gm), E = Equivalent weight of solute, V = volume of solution,
$n_E$ = number of gram equivalent of solute.
Number of Gram equivalents
= Normality of solution × volume of solution (lit.)
Milli gram equivalents
= Normality of solution × volume of solution (ml)

**Example 20 :**
4 gm NaOH is present in 100 ml of the solution what is the normality –

**Sol.** $\text{Normality} = \dfrac{w}{E} \times \dfrac{1000}{V\,(\text{ml})} \times \dfrac{4}{40} \times \dfrac{1000}{100} = 1\,N$

**Example 21 :**
12.6 gm oxalic acid present in 550 gm of the solution. Density of the solution is 1.10 gm/ml. What is the normality.

**Sol.** $N = \dfrac{w}{E} \times \dfrac{d}{W} \times 1000 = \dfrac{12.6}{63} \times \dfrac{1.10}{550} \times 1000 = 0.4\,N$

**(v) Molarity :** The number of gram moles of the solute dissolved per litre of the solution. It is denoted by 'M'.

$$\text{Molarity} = \frac{\text{Number of gram moles of solute}}{\text{volume of solution (lit.)}}$$

$\because$ $\text{Gram moles} = \dfrac{\text{weight of solute (gm)}}{\text{Molecular weight of solute}}$

$\therefore$ $\text{Molarity} = \dfrac{\text{weight of solute (gm)}}{\text{Molecular weight of solute}} \times \dfrac{1}{\text{volume of solution (lit.)}}$

**Formula :**

$M = \dfrac{w}{M'} \times \dfrac{1}{V(\text{lit.})}$ . ...... (i) $M = \dfrac{w}{M'} \times \dfrac{1000}{V\,(\text{ml.})}$ ......... (ii)

$M = n_M \times \dfrac{1}{V(\text{lit.})}$ ......... (iii) $n_M = M \times V(\text{lit.})$ ......... (iv)

Where, w = weight of solute,
M' = Molecular weight of solute, V = volume of solution,
$n_M$ = number of gram moles.
Gram moles=Molarity of solution × volume of solution (litre)
Milli moles = Molarity of solution × volume of solution (ml)

**Example 22 :**
3.65 gm HCl gas present in 100 ml of its aqueous sol$^n$. What is the molarity ?

**Sol.** $\text{Molarity} = \dfrac{w}{M'} \times \dfrac{1000}{\text{volume(ml.)}} = \dfrac{3.65}{36.5} \times \dfrac{1000}{100} = 1\,M$

"1M sol$^n$ of HCl".

**(vi) Molality :** The number of gram moles of solute dissolved in 1000 gm or 1 kg of the solvent. It is denoted by 'm'

$$\text{molality} = \frac{\text{Gram moles of solute}}{\text{weight of solvent (kg)}}$$

$\therefore$ $\text{Gram moles of solute} = \dfrac{\text{weight of solute (gm)}}{\text{Molecular weight of solute}}$

Molality

$$= \frac{\text{weight of solute (gm)}}{\text{Molecular weight of solute}} \times \frac{1}{\text{weight of solvent (kg)}}$$

**Formula :**

$m = \dfrac{w}{M'} \times \dfrac{1}{W(\text{kg})}$ ...... (i) $m = \dfrac{w}{M'} \times \dfrac{1000}{W(\text{gm})}$ ...... (ii)

$m = n_M \times \dfrac{1}{W(\text{kg})}$ ..... (iii)

Where w = weight of solute, M = molecular weight of solute, W = weight of solvent, $n_M$ = no. of moles of solute

**Example 23 :**
8 gm NaOH dissolved in 500 ml of its aqueous sol$^n$. If density of the solution is 1.2 gm/ml. then find the molality of the solution

**Sol.** Weight of solute = 8 gm, Volume of solution = 500 ml
Density of sol$^n$ = 1.2 gm/ml
$\therefore$ Weight of solution = 500 × 1.2 = 600 gm.
$\therefore$ Weight of solvent = weight of solution – weight of solute = 600 – 8 = 592 gm

$\therefore$ $m = \dfrac{w}{M'} \times \dfrac{1000}{W} = \dfrac{8}{40} \times \dfrac{1000}{592} = 0.34$

**(vii) Formality :** The no. of gram formula weight of a solute dissolved per litre of the solution is called formality of the solution. It is denoted by 'F'.

$$\text{Formality} = \frac{\text{Wt. of solute (gm)}}{\text{Formula wt. of solute}} \times \frac{1}{\text{Volume of solution (lit.)}}$$

$$F = \frac{w}{f} \times \frac{1}{V(\text{litre})} \quad \text{.........(i)} \;;\; F = \frac{w}{f} \times \frac{1000}{V(\text{ml})} \quad \text{........(ii)}$$

$$F = n_f \times \frac{1}{v(\text{litre})} \quad \text{..........(iii)}$$

Where, w = weight of solute, f = formula weight of solute, V = volume of solution, $n_f$ = no. of gram formula weight.

**(viii) Mole fraction :** The mole fraction of a component in a solution is the ratio of the number of moles of that component to the total number of moles present in the solution. Suppose : $\left.\begin{matrix}A - \text{Solute} \\ B - \text{Solvent}\end{matrix}\right\}\text{Solution}$

$n_A$ = No. of moles of solute, $n_B$ = No. of moles of solvent

Then mole fraction of solute $= X_A = \frac{n_A}{n_A + n_B}$

Mole fraction of solvent $= X_B = \frac{n_B}{n_A + n_B}$ ; $X_A + X_B = 1$

**For gaseous mixture :**

A binary system of two gases A & B

$P_A$ = Partial pressure of A, $P_B$ = Partial pressure of B

$P = P_A + P_B$ = Total pressure of gaseous mixture

Mole fraction of gas A, $X_A = \frac{P_A}{P_A + P_B} = \frac{P_A}{P}$

Mole fraction of gas B, $X_B = \frac{P_B}{P_A + P_B} = \frac{P_B}{P}$

**(ix) Mole percentage :**

Mole percentage = Mole fraction × 100

Mole percent of A = $X_A \times 100$

Mole percent of B = $X_B \times 100$

**(x) ppm. (Part per million) :**

The parts of the component per million parts ($10^6$) of the solution. $\text{ppm} = \frac{w}{w + W} \times 10^6$

Where, w = weight of solute, W = weight of solvent

**TITRATION**

Titration is a procedure of determination of concentration of unknown solution with the help of known concentrated sol$^n$.

In this procedure for determining the concentration of solution A by adding a carefully measured volumes of a solution with known concentration of B until the reaction of A & B is just complete.

**Law of equivalence :** The fundamental basis of titration is the 'Law of equivalence" which states that at end point of a titration volumes of the two titrants reacted have the same number of equivalents or milli equivalents.

**Acid base Titration :** One gm equivalent of acid neutralised by one gm equivalent of base. It means :

One equivalent of acid = one equivalent of base

Acid $[N_1V_1]$ = Base $[N_2V_2]$

[∵ gm equivalent = Normality × volume]

**Example 24 :**

Find the number of milli equivalents of $H_2SO_4$ present in 10ml of N/2 $H_2SO_4$ solution.

**Sol.** milli equivalents = Normality × volume (ml)

$= \frac{1}{2} \times 10 = 5$ milli equivalent of $H_2SO_4$

**Example 25 :**

10 milli equivalent KOH are present in its 100 ml sol$^n$. What is the normality :–

**Sol.** $\text{Normality} = \frac{\text{milli equalents of solute}}{\text{volume of solution (ml)}} = \frac{10}{100} = \frac{1}{10}$

= 0.1 N sol$^n$ of KOH

**STRENGTH EXPRESSION**

**Strength of $H_2O_2$. (Hydrogen peroxide)**

The strength of $H_2O_2$ is expressed in term of volume or weight percentage.

The strength of $H_2O_2$ is commonly expressed as 'volume'. This refers to the volume of oxygen which as solution of $H_2O_2$ will give at NTP.

$H_2O_2 \longrightarrow H_2O + \frac{1}{2} O_2$ ; $2H_2O_2 \longrightarrow 2H_2O + O_2$

20 volume of $H_2O_2$ means that 1 litre of this solution will give 20 litre of oxygen at NTP.

* Volume strength of $H_2O_2 = N \times 5.6$

* Volume strength of $H_2O_2 = 11.2 \times M$

$= \frac{11.2 \times \text{Percentage strength} \times 10}{\text{Mol. wt. of } H_2O_2 (34)}$, where M is molarity.

**Example 26 :**

Find the volume strength of 1.6 M $H_2O_2$ solution.

**Sol.** Strength of the solution = Molarity × mol. mass

$= 1.6 \times 34 = 54.4 \text{ gL}^{-1}$

$2H_2O_2 \rightarrow 2H_2O + O_2$

68g     22.4L

54.4     $\frac{22.4}{68} \times 54 = 19.92$ L

∴ Volume strength = 19.92 V

**Percentage labelling of oleum :**

Oleum is fuming sulphuric acid which contains extra $SO_3$ dissolved in $H_2SO_4$. To convert this extra $SO_2$ into $H_2SO_4$, water has to be added ($SO_3 + H_2O \rightarrow H_2SO_4$). The amount of sulphuric acid obtained when just sufficient water is added into 100g of oleum so that all $SO_3$ present in it is converted into $H_2SO_4$ is called percentage labelling of oleum.

* In oleum labelled as (100 + x)%

$\% \text{ of free } SO_3 = \left(\frac{80 \times x}{18}\right)(w/w)$.

**Example 27 :**

Given sample of oleum is labelled as 109%, calculate the % of free $SO_3$ in it ?

**Sol.** Let the weight of oleum sample be 100g, which on required dilution becomes 109g, so this means 9g of $H_2O$ was added.

$$SO_3 + H_2O \rightarrow H_2SO_4$$

Moles of $H_2O$ added = Moles of $SO_3$ present in oleum = $\frac{9}{18}$

Mass of of $SO_3$ in oleum = $\frac{9}{18} \times 80 = 40g$

$\therefore$ % of $SO_3$ in oleum = 40%

## TRY IT YOURSELF - 3

**Q.1** A solution is prepared by dissolving 1.0g of NaOH in water to get 250ml of solution. Calculate its molarity.

**Q.2** How many gram equivalents of $H_2SO_4$ are present in 200ml of (N/10) $H_2SO_4$ solution ?

**Q.3** x g of metal reacts with chlorine to form y g of metal chloride. Calculate equivalent mass of metal.

**Q.4** How many moles and how many grams of HCl are present in 250cm$^3$ of 0.5 M HCl solution ?

**Q.5** 100ml decinormal HCl is mixed to 100ml seminormal $H_2SO_4$ solution. Calculate normality of resulting mixture.

**Q.6** 200ml $\frac{N}{10}$ $H_2SO_4$ is mixed into 300ml $\frac{N}{100}$ NaOH. Calculate normality of resulting mixture.

**Q.7** Calculate the molality of a solution containing 5.3g of anhydrous $Na_2CO_3$ in 400g of water.

**Q.8** Find the equivalent weight of (a) $CaCO_3$ (b) $K_2SO_4 \times Al_2(SO_4)_3 \times 24\,H_2O$ (Mol. mass = M)

**Q.9** Calculate equivalent weight of $H_3PO_4$ and $Ca(OH)_2$ on the basis of given reaction.

$H_3PO_4 + NaOH \rightarrow NaH_2PO_4 + H_2O$

$Ca(OH)_2 + HCl \rightarrow Ca(OH)Cl + H_2O$

**Q.10** 1.26g of crystalline oxalic acid was dissolved in water to prepare 250ml of solution. Calculate molarity of solution.

## ANSWERS

**(1)** 0.1 M **(2)** 0.02 gram equivalent

**(3)** $\frac{x}{y-x} \times 35.5$ **(4)** 0.125 moles, 4.5625g

**(5)** 0.3 N **(6)** 0.03 N **(7)** 0.125 m

**(8)** (a) 50, (b) M/8 **(9)** 98, 74 **(10)** 0.04 M

## IMPORTANT POINTS

* The number of molecules in 1 cm$^3$ of gas at STP is known as Loschmidt number.
* Moles (gases) at NTP = $\frac{\text{volume (L)}}{22.4}$
* Dilution formula : $M_1V_1 = M_2V_2$
  For mixing two solutions of the same substance $M_1V_1 + M_2V_2 = M_3(V_1 + V_2)$
* 1 mol of $H_2O \neq 22400$ cc of $H_2O$ (because it is a liquid). Instead, 1 mol of $H_2O$ = 18cc of $H_2O$ (because density of $H_2O$ = 1g/cc)
* 1M $H_2SO_4$ = 2N $H_2SO_4$

## ADDITIONAL EXAMPLES

**Example 1 :**

An inorganic substance on analysis gave the following results Na = 29.1%, S = 40.5% and O = 30.4%. Calculate its empirical formula.

**Sol.**

| Element | % | At. wt. | Relative no. of atoms | Ratio | Simple whole no. ratio |
|---|---|---|---|---|---|
| Na | 29.1 | 23 | $\frac{29.1}{23} = 1.265$ | $\frac{1.265}{1.265} = 1$ | 2 |
| S | 40.5 | 32 | $\frac{40.5}{32} = 1.266$ | $\frac{1.265}{1.265} = 1$ | 2 |
| O | 30.4 | 16 | $\frac{30.4}{16} = 1.90$ | $\frac{1.90}{1.265} = 1.5$ | 3 |

Simplest whole number ratio of Na : S : O will be 2 : 2 : 3 and thus empirical formula will be $Na_2S_2O_3$.

**Example 2 :**

Find the vapour density of a gas is 11.2, then 11.2 g of this gas at N.T.P. will occupy a volume-

**Sol.** Vapour density of any gas occupies a volume of 11.2 litres at N.T.P.

**Example 3 :**

Find the weight percentage of 10 volume of $H_2O_2$.

**Sol.** $\% \text{ weight} = \frac{\text{volume}}{5.6} \times \frac{17}{10} = \frac{10}{5.6} \times \frac{17}{10} = 3.04\,\%$

**Example 4 :**

What is the mass of 1 molecule of CO.

**Sol.** Gram molecular weight of CO = 12 + 16 = 28 g

$6.023 \times 10^{23}$ molecules of CO weight 28 g

1 molecule of CO weighs $= \frac{28}{6.02 \times 10^{23}} = 4.65 \times 10^{-23}$ g

**Example 5 :**

Calculate the volume at STP occupied by 240 gm of $SO_2$.

**Sol.** Molecular weight of $SO_2 = 32 + 2 \times 16 = 64$

64 g of $SO_2$ occupies 22.4 litre at STP

240 g of $SO_2$ occupies $= \frac{22.4}{64} \times 240 = 84$ litre at STP

**Example 6 :**

$6 \times 10^{20}$ molecules of $CO_2$ are removed from 220 milligram of $CO_2$. What are the remaining moles of $CO_2$.

**Sol.** Mole of 220 mg. of $CO_2 = \frac{220 \times 10^{-3}}{44} = 5 \times 10^{-3}$ moles;

Moles of $CO_2$ removed are $= \frac{6 \times 10^{20}}{6 \times 10^{23}} = 10^{-3}$ moles

Remaining moles of $CO_2 = [5 \times 10^{-3} - 10^{-3}] = 4 \times 10^{-3}$ moles

**Example 7 :**

12 L of $H_2$ and 11.2 L of $Cl_2$ are mixed and exploded. Find the composition by volume of mixture.

**Sol.**

| $H_2$ + | $Cl_2 \rightarrow$ | 2HCl |
|---|---|---|
| 1L | 1L | 2L |
| 11.2 L | 11.2 L | 22.4L |

Volume of $H_2$ = [12 – 11.2] = 0.8 L,
Volume of $Cl_2$ = Zero, Volume of HCl = 22.4 L

**Example 8 :**

10 moles $SO_2$ and 15 moles $O_2$ were allowed to react over a suitable catalyst. 8 moles of $SO_3$ were formed. The remaining moles of $SO_2$ and $O_2$ respectively are -

(1) 2 moles, 11 moles (2) 2 moles, 8 moles
(3) 4 moles, 5 moles (4) 8 moles, 2 moles

**Sol. (1).**

| $2SO_2$ + | $O_2$ | $\longrightarrow$ $2SO_3$ |
|---|---|---|
| 10 | 15 | 0 |
| 10 – 2x | 15 – x | 2x |
| $\because$ 2x = 8 | x = 4 | |

Hence, remaining, $SO_2$ = 10 – 8 = 2 moles,
$O_2$ = 15 – 4 = 11 moles

**Example 9 :**

If 0.5 mol of $BaCl_2$ is mixed with 0.2 mole of $Na_3PO_4$, find the maximum amount of $Ba_3(PO_4)_2$ that can be formed.

| | $3\,BaCl_2$ + | $2\,Na_3PO_4 \rightarrow$ | $6\,NaCl$ | + $Ba_3(PO_4)_2$ |
|---|---|---|---|---|
| molar ratio | 3 | 2 | 6 | 1 |
| initial moles | 0.5 | 0.2 | 0 | 0 |

Limiting reagent is $Na_3PO_4$ hence it would be consumed, and the yield would be decided by it inital moles.
2 moles of $Na_3PO_4$ give 1 mole of $Ba_3(PO_4)_2$,
0.2 moles of $Na_3PO_4$ would give 0.1 mole of of $Ba_3(PO_4)_2$

**Example 10 :**

On reduction 1.644 gm of hot iron oxide give 1.15 gm of iron. Evaluate the equivalent weight of iron.

**Sol.** Weight of iron oxide = 1.6444 gm
Weight of iron after reduction = 1.15 gm
weight of displaced oxygen = 1.6444 – 1.15 = 0.4944 gm

$\therefore$ Equivalent weight of iron = $\frac{1.15}{0.4944} \times 8 = 18.61$

Thus equivalent weight of metal is = 18.61.

**Example 11 :**

A metallic chloride contain 47.22% metal calculate the equivalent weight of metal.

**Sol.** Suppose weight of metallic chloride = 100 gm
Then weight of metal = 47.22 gm
Weight of chlorine = 100 – 47.22 = 52.78 gm

$\therefore$ Equivalent weight of metal = $\frac{47.22}{52.78} \times 35.5 = 31.76$

**Example 12 :**

What is the volume of 10 N acetic acid required to prepare 400 ml of N–solution.

**Sol.** Equivalents of acetic acid in N $sol^n$ = Equivalents of acetic acid in 10 N $sol^n$.

$N_1V_1 = N_2V_2$ ; $1 \times 400 = 10 \times V_2$ ; $V_2 = 40$ ml.

**Example 13 :**

6.8 gm $H_2O_2$ present in 100 ml of its $sol^n$ . What is the molarity of solution.

**Sol.** $\because$ Weight of $H_2O_2$ in 100 ml of $H_2O_2$ $sol^n$ = 6.8 gm
$\because$ Weight of $H_2O_2$ in 1000 ml of its $sol^n$ = 6.8 × 10 = 68 gm
Molecular weight of $H_2O_2$ = 34

Then, Molarity = $\frac{68}{34}$ = 2M or % wt. = Molarity × 2 × $\frac{17}{10}$

$\therefore$ Molarity = $\frac{\%\,weight \times 10}{2 \times 17} = \frac{6.8 \times 10}{2 \times 17} = 2M$

# QUESTION BANK

## CHAPTER 1 : SOME BASIC CONCEPTS OF CHEMISTRY

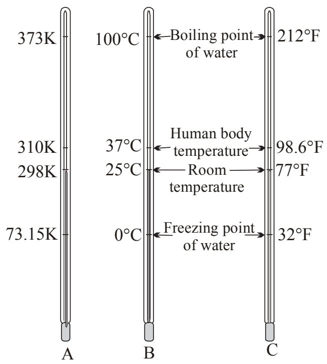
## EXERCISE - 1 [LEVEL-1]

**Choose one correct response for each question.**

### ART - 1 : PROPERTIES OF MATTER AND THEIR MEASUREMENT

**Q.1** Which of the following statements about a compound is incorrect?
(A) A molecule of a compound has atoms of different elements.
(B) A compound cannot be separated into its constituent elements by physical methods of separation.
(C) A compound retains the physical properties of its constituent elements.
(D) The ratio of atoms of different elements in a compound is fixed.

**Q.2** Thermometers using different temperature scales are given in the picture. Here A, B, C refer to –




Choose the correct option.
(A) A – Celsius, B – Fahrenheit, C – Kelvin
(B) A – Kelvin, B – Celsius, C – Fahrenheit
(C) A – Fahrenheit, B – Celsius, C – Kelvin
(D) A – Kelvin, B – Fahrenheit, C – Celsius

**Q.3** Which of the following is not true of mixtures –
(A) Mixtures can be homogeneous or heterogeneous.
(B) Components in a mixture are present in a fixed ratio.
(C) Properties of a mixture are the average of its components.
(D) Components of a mixture can be separated easily by simple physical methods.

### ART - 2 : UNCERTAINTY IN MEASUREMENT

**Q.4** ___ is are meaningful digit(s) which is/are known with certainty.
(A) Scientific notation (B) Precision
(C) Accuracy (D) Significant figures

**Q.5** Two students performed the same experiment separately and each one of them recorded two readings of mass which are given below. Correct reading of mass is 3.0 g. On the basis of given data, mark the correct option out of the following statements.

| Student | Readings (i) | Readings (ii) |
|---|---|---|
| A | 3.01 | 2.99 |
| B | 3.05 | 2.95 |

(A) Results of both the students are neither accurate nor precise.
(B) Results of student A are both precise and accurate.
(C) Results of student B are neither precise nor accurate.
(D) Results of student B are both precise and accurate.

**Q.6** $1.00 \times 10^2$ has ___ significant figures.
(A) two (B) one
(C) three (D) zero

**Q.7** Which of the following statement(s) is/are true?
I. Every experimental measurement has zero amount of uncertainty associated with it.
II. One would always like the result to be precise and accurate.
III. Precision and accuracy are often referred to while we talk about the measurement.
(A) I and II
(B) II and III
(C) I and III
(D) All the above statements are true.

**Q.8** How many significant figures are present in $0.010100 \times 10^3$ ?
(A) 7 (B) 5
(C) 3 (D) 10

**Q.9** How many significant figures are in each of the following numbers :
(a) 4.003 (b) $6.023 \times 10^{23}$ (c) 5000
(A) 3, 4, 1 (B) 4, 3, 2
(C) 4, 4, 4 (D) 3, 4, 3

**Q.10** The result of the operation $2.5 \times 1.25$ should be on the basis of significant figures?
(A) 3.125 (B) 3.13
(C) 3.1 (D) 31.25

### PART - 3 : LAWS OF CHEMICAL COMBINATIONS

**Q.11** Which of the following statements best explains the law of conservation of mass?
(A) 100 g of water is heated to give steam.
(B) A sample of $N_2$ gas is heated at constant pressure without any change in mass.
(C) 36 g of carbon combines with 32 g of oxygen to form 68 g of $CO_2$.
(D) 10 g of carbon is heated in vacuum without any change in mass.

**Q.12** Consider the following statements,
a. Matter consists of indivisible atoms.
b. All the atoms of a given elements have identical properties including identical mass.
c. Atoms of different elements differ in mass.
d. Compounds are formed when atoms of different elements combine in a fixed ratio.
e. Atoms can neither be created nor destroyed in a chemical reaction.
The correct statements are
(A) a, b, c are correct statements
(B) c, d, e are correct statements
(C) a, c, e are correct statements
(D) all are correct statements

**Q.13** Which of the following statements is correct about the given reaction: $4Fe\,(s) + 3O_2(g) \rightarrow 2Fe_2O_3(g)$
(A) Total mass of iron and oxygen in reactants = total mass of iron and oxygen in product; therefore, it follows law of conservation of mass.
(B) Total mass of reactants = total mass of product; therefore, law of multiple proportions is followed.
(C) Amount of $Fe_2O_3$ can be increased by taking anyone of the reactants (iron or oxygen) in excess.
(D) Amount of $Fe_2O_3$ produced will decrease if the amount of anyone of the reactants (iron or oxygen) is taken in excess.

**Q.14** 4.88g of $KClO_3$ when heated produced 1.92 g of $O_2$ and 2.96 g of KCl. Which of the following statements regarding the experiment is correct?
(A) The result illustrates the law of conservation of mass.
(B) The result illustrates the law of multiple proportions.
(C) The result illustrates the law of constant proportion.
(D) None of these

**Q.15** Which law states that matter can neither be created nor destroyed?
(A) Law of definite proportions
(B) Law of conservation of mass
(C) Law of multiple proportions
(D) Avogadro law

**Q.16** Hydrogen combines with oxygen to form two compounds namely, water & hydrogen peroxide.

| Hydrogen + | oxygen → | Water |
|---|---|---|
| 2 g | 16 g | 18 g |

| Hydrogen + | oxygen → | Hydrogen peroxide |
|---|---|---|
| 2 g | 32 g | 34 g |

Here, the masses of oxygen which combine with a fixed mass of hydrogen (2 g) bear simple ratio _____.
(A) 2 : 1 (B) 1 : 2
(C) 3 : 4 (D) 4 : 3

**Q.17** Proust worked with the two samples of cupric carbonate one of which was of natural origin and the other was synthetic one. He found that the composition of elements present in it was same for both the samples as shown below.

| Sample | % of Cu | % of $O_2$ | % of C |
|---|---|---|---|
| Natural sample | 51.35 | 9.74 | 38.91 |
| Synthetic sample | 51.35 | 9.74 | 38.91 |

Which law is in favour of the above data?
(A) Law of multiple proportions
(B) Gay Lussac's law of gaseous volumes
(C) Avogadro law
(D) Law of definite proportions

**Q.18** Which of the following reactions is not correct according to the law of conservation of mass?
(A) $2Mg\,(s) + O_2(g) \rightarrow 2MgO(s)$
(B) $C_3H_8\,(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
(C) $P_4(s) + 5O_2(g) \rightarrow P_4O_{10}\,(s)$
(D) $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O\,(g)$

## PART - 4 : ATOMIC AND MOLECULAR MASSES

**Q.19** Naturally occurring chlorine is 75.53% $Cl^{35}$ which has an atomic mass of 34.969 amu and 24.47% $Cl^{37}$ which has a mass of 36.966 amu. Calculate the average atomic mass of chlorine-
(A) 35.5 amu (B) 36.5 amu
(C) 71 amu (D) 72 amu

**Q.20** 'amu' has been replaced by 'u' which is known as ____.
(A) unified mass (B) uni mass
(C) unitech mass (D) unit mass

**Q.21** Given, that the abundances of isotopes $^{54}Fe$, $^{56}Fe$ and $^{57}Fe$ are 5%, 90% and 5% respectively, the atomic mass of Fe is –
(A) 55.85 (B) 55.95
(C) 55.75 (D) 56.05

**Q.22** Use the data given in the table to calculate the molar mass of naturally occurring argon :

| Isotope | Isotopic molar mass | Abundance |
|---|---|---|
| $^{36}Ar$ | 35.96755 g $mol^{-1}$ | 0.337% |
| $^{38}Ar$ | 37.96272 g $mol^{-1}$ | 0.063% |
| $^{40}Ar$ | 39.9624 g $mol^{-1}$ | 9.600% |

(A) 39.948 g/mol (B) 39.498 g/mol
(C) 38.948 g/mol (D) 39.849 g/mol

## PART - 5 : MOLE CONCEPT

**Q.23** The mass of one mole of a substance in grams is called its ______.
(A) Avogadro mass (B) molar mass
(C) atomic mass (D) formula mass

**Q.24** How many carbon atoms are present in 0.35 mol of $C_6H_{12}O_6$ -
(A) $6.023 \times 10^{23}$ carbon atoms
(B) $1.26 \times 10^{23}$ carbon atoms
(C) $1.26 \times 10^{24}$ carbon atoms
(D) $6.023\times 10^{24}$ carbon atoms

**Q.25** How many molecules are in 5.23 gm of glucose ($C_6H_{12}O_6$)-
(A) $1.65 \times 10^{22}$ (B) $1.75 \times 10^{22}$
(C) $1.75 \times 10^{21}$ (D) None of these

**Q.26** What is the weight of $3.01 \times 10^{23}$ molecules of ammonia
(A) 17 gm (B) 8.5 gm
(C) 34 gm (D) None of these

**Q.27** How many number of molecules and atoms respectively are present in 2.8 litres of a diatomic gas at STP?
(A) $6.023 \times 10^{23}$, $7.5 \times 10^{23}$
(B) $6.023 \times 10^{23}$, $15 \times 10^{22}$
(C) $7.5 \times 10^{22}$, $15 \times 10^{22}$
(D) $15 \times 10^{22}$, $7.5 \times 10^{23}$

**Q.28** The number of atoms present in one mole of an element is equal to Avogadro number. Which of the following element contains the greatest number of atoms?
(A) 4 g He (B) 46 g Na
(C) 0.4 g Ca (D) 12 g He

**Q.29** Which one of the following will have largest number of atoms?
(A) 1 g Au (s) (B) 1 g Na (s)
(C) 1 g Li (s) (D) 1 g of $Cl_2$(g)

**Q.30** How many number of aluminium ions are present in 0.051 g of aluminium oxide?
(A) $6.023 \times 10^{20}$ ions (B) 3 ions
(C) $6.023 \times 10^{23}$ ions (D) 9 ions

**Q.31** For three moles of ethane ($C_2H_6$), choose the correct statement –
(A) Number of moles of carbon atoms is 6.
(B) Number of moles of hydrogen atoms is 18.
(C) Number of molecules of ethane is $18.069 \times 10^{23}$
(D) All of these

**Q.32** How many moles of oxygen gas can be produced during electrolytic decomposition of 180 g of water?
(A) 2.5 moles (B) 5 moles
(C) 10 moles (D) 7 moles

**Q.33** What is the mass ratio of fluorine to boron in a boron trifluoride molecule?
(A) 1.8 to 1 (B) 3.0 to 1
(C) 3.5 to 1 (D) 5.3 to 1

**Q.34** What is the mass of oxygen in 148 grams of calcium hydroxide ($Ca(OH)_2$)?
(A) 16 grams (B) 24 grams
(C) 32 grams (D) 64 grams

## PART - 6 : PERCENTAGE COMPOSITION

**Q.35** A compound containing only sulphur and oxygen is 50% sulphur by weight. What is the empirical formula for the compound?
(A) SO (B) $SO_2$
(C) $SO_3$ (D) $S_2O$

**Q.36** A hydrocarbon was found to be 20% hydrogen by weight. If 1 mole of the hydrocarbon has a mass of 30 grams, what is its molecular formula?
(A) CH (B) $CH_2$
(C) $CH_3$ (D) $C_2H_6$

**Q.37** A hydrocarbon contains 75% carbon by mass. What is the empirical formula for the compound?
(A) $CH_2$ (B) $CH_3$
(C) $CH_4$ (D) $C_2H_5$

**Q.38** A compound contains 69.5% oxygen and 30.5% nitrogen and its molecular weight is 92. The formula of compound is –
(A) $N_2O$ (B) $NO_2$
(C) $N_2O_4$ (D) $N_2O_5$

**Q.39** An organic compound on analysis was found to contain 10.06% carbon, 0.84% hydrogen and 89.10% chlorine. What will be the empirical formula of the substance?
(A) $CH_2Cl_2$ (B) $CHCl_3$
(C) $CCl_4$ (D) $CH_3Cl$

**Q.40** What is the mass percentage of carbon, in ethanol?
(A) 50.00% (B) 52.14%
(C) 55.00% (D) 51.04%

**Q.41** Two elements 'P' and 'Q' combine to form a compound. Atomic mass of 'P' is 12 and 'Q' is 16. Percentage of 'P' in the compound is 27.3. What will be the empirical formula of the compound?
(A) $P_2Q_2$ (B) PQ
(C) $P_2Q$ (D) $PQ_2$

**Q.42** An oxide of iodine (I = 127) contains 25.4g of iodine for 8g of oxygen. Its formula could be
(A) $I_2O_3$ (B) $I_2O$
(C) $I_2O_5$ (D) $I_2O_7$

**Q.43** The empirical formula of a compound is $CH_2$.One mole of this compound has a mass of 42 g. Its molecular formula is –
(A) $C_3H_6$ (B) $C_3H_8$
(C) $CH_2$ (D) $C_2H_2$

**Q.44** The empirical formula and molecular mass of a compound are $CH_2O$ and 180 g respectively. What will be the molecular formula of the compound?
(A) $C_9H_{18}O_9$ (B) $CH_2O$
(C) $C_6H_{12}O_6$ (D) $C_2H_4O_2$

## PART - 7 : STOICHIOMETRY AND STOICHIOMETRIC CALCULATIONS

**Q.45** The concentration of sodium chloride in sea water is about 0.5 molar. How many grams of NaCl are present in 1 kg of sea water?
(A) 30 grams (B) 60 grams
(C) 100 grams (D) 300 grams

**Q.46** A sample of a hydrate of $CuSO_4$ with a mass of 250 grams was heated until all the water was removed. The sample was then weighed and found to have a mass of 160 grams. What is the formula for the hydrate?
(A) $CuSO_4 \cdot 10\,H_2O$ (B) $CuSO_4 \cdot 7\,H_2O$
(C) $CuSO_4 \cdot 5\,H_2O$ (D) $CuSO_4 \cdot 2\,H_2O$

**Q.47** $ZnSO_3(s) \rightarrow ZnO(s) + SO_2(g)$
What is the STP volume of $SO_2$ gas produced by the above reaction when 145 grams of $ZnSO_3$ are consumed?
(A) 23 liters (B) 36 liters
(C) 45 liters (D) 56 liters

**Q.48** $CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(\ell) + CO_2(g)$
If the reaction above took place at standard temperature and pressure and 150 grams of $CaCO_3(s)$ were consumed, what was the volume of $CO_2$ (g) produced at STP?
(A) 11 L (B) 22 L
(C) 34 L (D) 45 L

**Q.49** 8 litre of $H_2$ and 6 litre of $Cl_2$ are allowed to react to maximum possible extent. Find out the final volume of reaction mixture. Suppose P & T remains constant throughout the course of reaction
(A) 7 litre (B) 14 litre
(C) 2 litre (D) None of these.

**Q.50** The molarity of a solution of ethanol in water in which the mole fraction of $C_2H_5OH$ is 0.040 is–
(Assume density of solution to be 1 g/ml)
(A) 2.09 M (B) 2.31 M
(C) 20.9 M (D) 23.1 M

**Q.51** 0.250 g of an element M reacts with excess of fluorine to produce 0.547 g of the hexafluoride $MF_6$. What is the element ? (Atomic weights of F = 19, Cr = 52, Mo = 96, S = 32, Te = 127.6)
(A) Cr (B) Mo
(C) S (D) Te

**Q.52** Which of the following has the highest normality? (consider each of the acid is 100% ionised.)
(A) 1 (M) $H_2SO_4$ (B) 1 (M) $H_3PO_3$
(C) 1 (M) $H_3PO_4$ (D) 1 (M) $HNO_3$

**Q.53** The weight of AgCl precipitated when a solution containing 5.85 g of NaCl is added to a solution containing 3.4 g of $AgNO_3$ is
(A) 28 g (B) 9.25 g
(C) 2.870 g (D) 58 g

**Q.54** A sample of nitric acid is 69% by mass and it has a concentration of 15.44 moles per litre. Its density is –
(A) 1.86 g/cc (B) 1.41 g/cc
(C) 2.60 g/cc (D) 1.02 g/cc

**Q.55** If the concentration of glucose ($C_6H_{12}O_6$) in blood is 0.9 g $L^{-1}$, what will be the molarity of glucose in blood?
(A) 5M (B) 50 M
(C) 0.005 M (D) 0.5 M

**Q.56** What volume of water is to be added to 100$cm^3$ of 0.5 M NaOH solution to make it 0.1 M solution?
(A) 200 $cm^3$ (B) 400 $cm^3$
(C) 500 $cm^3$ (D) 100 $cm^3$

**Q.57** What will be the molality of the solution made by dissolving 10 g of NaOH in 100 g of water?
(A) 2.5 m (B) 5 m
(C) 10 m (D) 1.25 m

**Q.58** The weight of lime obtained by heating 200 kg of 95% pure lime stone is
(A) 98.4 kg (B) 106.4 kg
(C) 112.8 kg (D) 122.6 kg

**Q.59** What will be the molarity of a solution, which contains 5.85 g of NaCl (s) per 500 mL?
(A) 4 mol $L^{-1}$ (B) 20 mol $L^{-1}$
(C) 0.2 mol $L^{-1}$ (D) 2 mol $L^{-1}$

**Q.60** What will be the molality of the solution containing 18.25 g of HCl gas in 500 g of water?
(A) 0.1 m (B) 1 M
(C) 0.5 m (D) 1 m

**Q.61** A solution is prepared by adding 2 g of a substance A to 18 g of water. Calculate the mass per cent of the solute.
(A) 8% (B) 9%
(C) 10% (D) 11%

**Q.62** A sample of $H_2O_2$ solution labelled as 28 volume has density of 26.5 g/L. Mark the INCORRECT option representing concentration of same solution in other units
(A) $M_{H_2O_2} = 2.5$ (B) $M_{H_2O_2} = 13.88$
(C) Mole fraction of $H_2O_2$ = 0.2 (D) $\%\frac{w}{v} = 17$

**Q.63** What is the concentration of copper sulphate (in mol $L^{-1}$) if 80 g of it is dissolved in enough water to make a final volume of 3 L?
(A) 0.0167 (B) 0.167
(C) 1.067 (D) 10.67

**Q.64** $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
How many moles of methane are required to produce 22 g $CO_2$ (g) after combustion?
(A) 1 mol (B) 0.5 mol
(C) 0.25 mol (D) 1.25 mol

**Q.65** A solution is made by dissolving 49 g of $H_2SO_4$ in 250 mL of water. The molarity of the solution prepared is –
(A) 2 M (B) 1 M
(C) 4 M (D) 5 M

**Q.66** The molarity of pure water is
(A) 18 M (B) 50.0 M
(C) 55.6 M (D) 100 M

**Q.67** 10 mL of gaseous hydrocarbon on combustion give 40 mL of $CO_2(g)$ and 50 mL of $H_2O$ (vapour). The hydrocarbon is –
(A) $C_4H_5$ (B) $C_8H_{10}$
(C) $C_4H_8$ (D) $C_4H_{10}$

**Q.68** If 1.6 g of $SO_2$ and 1.5 × $10^{22}$ molecules of $H_2S$ are mixed and allowed to remain in contact in a closed vessel until the reaction: $2H_2S + SO_2 \rightarrow 3S + 2H_2O$, proceeds to completion. Which of the following statement is true?
(A) Only S and $H_2O$ remain in the reaction vessel.
(B) $H_2S$ will remain in excess
(C) $SO_2$ will remain in excess
(D) None

## PART - 8 : EQUIVALENT WEIGHTS

**Q.69** One g equivalent of a substance is present in -
(A) 0.25 mole of $O_2$ (B) 0.5 mole of $O_2$
(C) 1.00 mole of $O_2$ (D) 8.00 mole of $O_2$

**Q.70** Sulphur forms two chlorides $S_2Cl_2$ and $SCl_2$. The equivalent mass of sulphur in $SCl_2$ is 16. The equivalent weight of sulphur in $S_2Cl_2$ is -
(A) 8 (B) 16
(C) 32 (D) 64

**Q.71** Equivalent weight of a divalent metal is 24. The volume of hydrogen liberated at STP by 12 g of the same metal when added to excess of an acid solution is -
(A) 2.8 litres (B) 5.6 litres
(C) 11.2 litres (D) 22.4 litres

**Q.72** 0.84 g of a metal carbonate reacts exactly with 40mL of N/2 $H_2SO_4$. The equivalent weight of the metal carbonate is -
(A) 84 (B) 64
(C) 42 (D) 32

**Q.73** 1.0 g of a metal combines with 8.89 g of Bromine. Equivalent weight of metal is nearly:
(Atomic weight of Br = 80)
(A) 8 (B) 9
(C) 10 (D) 7

**Q.74** $H_3PO_4$ is a tribasic acid and one of its salt is $NaH_2PO_4$. What volume of 1 M NaOH solution should be added to 12 g $NaH_2PO_4$ to convert it into $Na_3PO_4$ ? (Atomic weight of P = 31)
(A) 100 mL (B) 200 mL
(C) 80 mL (D) 300 mL

**Q.75** 0.84 g. of metal hydride contains 0.04 g of hydrogen. The equivalent weight of metal is –
(A) 80 (B) 40
(C) 20 (D) 60

**Q.76** When an element forms an oxide in which oxygen is 20% of the oxide by mass, the equivalent mass of the element will be –
(A) 32 (B) 40
(C) 60 (D) 128

**Q.77** One g of hydrogen is found to combine with 80g of bromine. One g of calcium (valency = 2) combines with 4 g of bromine. The equivalent weight of calcium is –
(A) 10 (B) 20
(C) 40 (D) 80

**Q.78** A metal oxide is reduced by heating it in a stream of hydrogen. It is found that after complete reduction 3.15 g of the oxide have yielded 1.05g of the metal. We may conclude that.
(A) Atomic weight of the metal is 4.
(B) Equivalent weight of the metal is 8.
(C) Equivalent weight of the metal is 4.
(D) Atomic weight of the metal is 8.

## EXERCISE - 2 [LEVEL-2]

**Choose one correct response for each question.**

**Q.1** In 5 g atom of Ag (Atomic weight of Ag = 108), calculate the weight of one atom of Ag -
(A) $17.93 \times 10^{-23}$ gm (B) $16.93 \times 10^{-23}$ gm
(C) $17.93 \times 10^{23}$ gm (D) $36 \times 10^{-23}$ gm

**Q.2** How many molecules are present in one ml of water vapours at STP -
(A) $1.69 \times 10^{19}$ (B) $2.69 \times 10^{-19}$
(C) $1.69 \times 10^{-19}$ (D) $2.69 \times 10^{19}$

**Q.3** How many years it would take to spend Avogadro's number of rupees at the rate of 1 million rupees in one second -
(A) $19.098 \times 10^{19}$ years (B) 19.098 years
(C) $19.098 \times 10^{9}$ years (D) None of these

**Q.4** An atom of an element weighs $6.644 \times 10^{-23}$ g. Calculate g atoms of element in 40 kg-
(A) 10 gm atom (B) 100 gm atom
(C) 1000 gm atom (D) $10^4$ gm atom

**Q.5** Calculate the number of $Cl^-$ and $Ca^{+2}$ ions in 222 g anhydrous $CaCl_2$ –
(A) 2N ions of $Ca^{+2}$ 4 N ions of $Cl^-$
(B) 2N ions of $Cl^-$ & 4N ions of $Ca^{+2}$
(C) 1N ions of $Ca^{+2}$ & 1N ions of $Cl^-$
(D) None of these.

**Q.6** Calculate the weight of lime (CaO) obtained by heating 200 kg of 95% pure lime stone ($CaCO_3$).
(A) 104.4 kg (B) 105.4 kg
(C) 212.8 kg (D) 106.4 kg

**Q.7** The density of $O_2$ at NTP is 1.429g / litre. Calculate the standard molar volume of gas-
(A) 22.4 lit. (B) 11.2 lit
(C) 33.6 lit (D) 5.6 lit.

**Q.8** Choose the correct statements –
(A)The number of atoms in 52 mole of He is $31.3 \times 10^{24}$
(B) The number of atoms in 52 amu of He is 13.
(C) The number of atoms in 52g of He is $78.26 \times 10^{23}$
(D) All of these

**Q.9** From 160g sample of $SO_2$, $1.2046 \times 10^{24}$ atoms are removed, find the volume of remaining $SO_2$ at STP.
(A) 11.2 litre (B) 12.2 litre
(C) 5.2 litre (D) 15.4 litre

**Q.10** 1.25g of a solid diabasic acid is completely neutralized by 25mL of 0.25 molar $Ba(OH)_2$ solution. Molecular mass of the acid is–
(A) 100 (B) 150
(C) 120 (D) 200

**Q.11** 0.30 g of an organic compound containing C, H and oxygen on combustion yields 0.44 g $CO_2$ and 0.18 g $H_2O$. If one mol of compound weighs 60, then molecular formula of the compound is –
(A) $CH_2O$ (B) $C_3H_8O$
(C) $C_4H_6O$ (D) $C_2H_4O_2$

**Q.12** The density of oxygen gas at NTP is –
(A) 1.429 g/L (B) 1.429 g/mL
(C) 14.29 g/L (D) 0.1429 g/L

**Q.13** A 5.82g silver coin is dissolved in nitric acid when sodium chloride is added to the solution all the silver is precipitated as AgCl. The AgCl precipitate weighs 7.20 g. What is the percentage of silver in coin ?
(A) 98% (B) 93.1 %
(C) 86% (D) 82 %

**Q.14** In the reaction $Br_2 + Na_2CO_3 \rightarrow NaBr + NaBrO_3 + CO_2$ The equiv. wt. of $NaBrO_3$ is
(A) $\frac{\text{Mol. wt}}{1}$ (B) $\frac{\text{Mol. wt}}{10}$
(C) $\frac{\text{Mol. wt}}{5}$ (D) $\frac{\text{Mol. wt}}{4}$

**Q.15** 19g of a mixture containing $NaHCO_3$ and $Na_2CO_3$ on complete heating liberated 1.12 L of $CO_2$ at STP. The weight of the remaining solid was 15.9 g. What is the weight (in g) of $Na_2CO_3$ in the mixture before heating?
(A) 8.4 (B) 15.9
(C) 4.0 (D) 10.6

**Q.16** The number of molecules of $CO_2$ liberated the complete combustion of 0.1 g atom graphite in air is
(A) $3.01 \times 10^{22}$ (B) $6.02 \times 10^{23}$
(C) $6.02 \times 10^{22}$ (D) $3.01 \times 10^{23}$

**Q.17** In which one of the following, does the given amount of chlorine exert the least pressure in a vessel of capacity 1 $dm^3$ at 273 K?
(A) 0.071 g (B) 0.0355 g
(C) 0.02 mole (D) $6.023 \times 10^{21}$ molecules

**Q.18** What is the molarity of $H_2SO_4$ solution that has a density of 1.84 g/cc at 35°C and contains 98% by weight?
(A) 4.18 M (B) 8.14 M
(C) 18.4 M (D) 18 M

**Q.19** 50 $cm^3$ of 0.2 N HCl is titrated against 0.1 N NaOH solution. The titration was discontinued after adding 50$cm^3$ of NaOH. The remaining titration is completed by adding 0.5N KOH. The volume of KOH required for completing the titration is –
(A) 12 $cm^3$ (B) 10 $cm^3$
(C) 21.0 $cm^3$ (D) 16.2 $cm^3$

**Q.20** 20 ml of methane is completely burnt using 50ml of oxygen. The volume of the gas left after cooling to room temperature is –
(A) 80 ml (B) 40 ml
(C) 60 ml (D) 30 ml

**Q.21** 100 ml of 0.1 M acetic acid is completely neutralized using a standard solution of NaOH. The volume of ethane obtained at STP after the complete electrolysis of the resulting solution is
(A) 112 ml (B) 56 ml
(C) 224 ml (D) 560 ml

**Q.22** The total number of electrons in 18 ml of water (density = 1 g $ml^{-1}$) is–
(A) $6.02 \times 10^{23}$ (B) $6.02 \times 10^{25}$
(C) $6.02 \times 10^{24}$ (D) $6.02 \times 18 \times 10^{23}$

**Q.23** The volume of 0.1 M oxalic acid that can be completely oxidized by 20 ml of 0.025 M $KMnO_4$ solution is
(A) 125 ml (B) 25 ml
(C) 12.5 ml (D) 37.5 ml

**Q.24** The equivalent mass of a certain bivalent metal is 20. The molecular mass of its anhydrous chloride is –
(A) 91 (B) 111
(C) 55.5 (D) 75.5

**Q.25** The number of water molecules present in a drop of water weighing 0.018 gm is
(A) $6.022 \times 10^{26}$ (B) $6.022 \times 10^{23}$
(C) $6.022 \times 10^{19}$ (D) $6.022 \times 10^{20}$

**Q.26** Empirical formula of a compound is $CH_2O$ and its molecular mass is 90, the molecular formula of the compound is
(A) $C_3H_6O_3$ (B) $C_2H_4O_2$
(C) $C_6H_{12}O_6$ (D) $CH_2O$

**Q.27** The mass of 112 $cm^3$ of $NH_3$ gas at STP is
(A) 0.085 g (B) 0.850 g
(C) 8.500 g (D) 80.500 g

**Q.28** 10 g of a mixture of BaO and CaO requires 100$cm^3$ of 2.5M HCl to react completely. The percentage of calcium oxide in the mixture is approximately
(Given : molar mass of BaO = 153)
(A) 52.6 (B) 55.1
(C) 44.9 (D) 47.4

**Q.29** 25 $cm^3$ of oxalic acid completely neutralised 0.064 g of sodium hydroxide. Molarity of the oxalic acid solution is
(A) 0.064 (B) 0.045
(C) 0.015 (D) 0.032

**Q.30** 5.5 mg of nitrogen gas dissolves in 180 g of water at 273 K and one atm pressure due to nitrogen gas. The mole fraction of nitrogen in 180 g of water at 5 atm nitrogen pressure is approximately
(A) $1 \times 10^{-6}$ (B) $1 \times 10^{-5}$
(C) $1 \times 10^{-3}$ (D) $1 \times 10^{-4}$

**Q.31** 50 $cm^3$ of 0.04 M $K_2Cr_2O_7$ in acidic medium oxidizes a sample of $H_2S$ gas to sulphur. Volume of 0.03 M $KMnO_4$ required to oxidize the same amount of $H_2S$ gas to sulphur, in acidic medium
(A) 60 $cm^3$ (B) 80 $cm^3$
(C) 90 $cm^3$ (D) 120 $cm^3$

**Q.32** 0.06% (w/v) aqueous solution of urea is isotonic with
(A) 0.06% glucose solution
(B) 0.6% glucose solution
(C) 0.01 M glucose solution
(D) 0.1 M glucose solution

## EXERCISE - 3 (NUMERICAL VALUE BASED QUESTIONS)

**NOTE : The answer to each question is a NUMERICAL VALUE.**

**Q.1** A sample consisting of choclate-brown powder of $PbO_2$ is allowed to react with excess of KI and iodine liberated is reacted with $N_2H_4$ in another container. The volume of gas liberated from this second container at STP was measured out to be 1.12 litre. Find out volume of decimolar NaOH (in litre) required to dissolve $PbO_2$ completely. (Assume all reactions are 100% complete)

**Q.2** A sample of $Fe_2(SO_4)_3$ and $FeC_2O_4$ was dissolved in dil. $H_2SO_4$. The complete oxidation of reaction mixture required 40ml of N/16 $KMnO_4$. After the oxidation, the reaction mixture was reduced by Zn and dil. $H_2SO_4$. On again oxidation by same $KMnO_4$, 60ml were required. The ratio of millimoles of $Fe_2(SO_4)_3$ and $FeC_2O_4$ is 7 : A. Find the value of A. (mol wt. $Fe_2(SO_4)_3$ = 400).

**Q.3** 20% surface sites have adsorbed $N_2$. On heating $N_2$ gas evolved from sites and were collected at 0.001 atm and 298K in a container of volume is 2.46 $cm^3$. Density of surface sites is $6.023 \times 10^{14}/cm^2$ and surface area is 1000 $cm^2$, find out the no. of surface sites occupied per molecule of $N_2$.

**Q.4** The value of n in the molecular formula $Be_nAl_2Si_6O_{18}$ is

**Q.5** A student performs a titration with different burettes and finds titre values of 25.2 mL, 25.25 mL, and 25.0 mL. The number of significant figures in the average titre value is :

**Q.6** Reaction of $Br_2$ with $Na_2CO_3$ in aqueous solution gives sodium bromide and sodium bromate with evolution of $CO_2$ gas. The number of sodium bromide molecules involved in the balanced chemical equation is

**Q.7** The volume (in mL) of 0.1 M $AgNO_3$ required for complete precipitation of chloride ions present in 30 mL of 0.01 M solution of $[Cr(H_2O)_5Cl]Cl_2$, as silver chloride is close to

**Q.8** 29.2 % (w/w) HCl stock solution has a density of 1.25 g $mL^{-1}$. The molecular weight of HCl is 36.5g $mol^{-1}$. The volume (mL) of stock solution required to prepare a 200 mL solution of 0.4 M HCl is.

**Q.9** If the value of Avogadro number is $6.023 \times 10^{23}$ $mol^{-1}$ and the value of Boltzmann constant is $1.380 \times 10^{-23}$ J $K^{-1}$, then the number of significant digits in the calculated value of the universal gas constant is

**Q.10** A compound $H_2X$ with molar weight of 80 g is dissolved in a solvent having density of 0.4 g $ml^{-1}$. Assuming no change in volume upon dissolution, the molality of a 3.2 molar solution is –

# EXERCISE - 4 [PREVIOUS YEARS AIEEE / JEE MAIN QUESTIONS]

**Q.1** The weight of $2.01 \times 10^{23}$ molecules of CO is - **[AIEEE-2002]**
(A) 9.3 gm (B) 7.2 gm
(C) 1.2 gm (D) 3 gm

**Q.2** In an organic compound of molar mass 108 gm $mol^{-1}$ C, H and N atoms are present in 9 : 1 : 3.5 by weight. Molecular formula can be – **[AIEEE-2002]**
(A) $C_6H_8N_2$ (B) $C_7H_{10}N$
(C) $C_5H_6N_3$ (D) $C_4H_{18}N_3$

**Q.3** Number of atoms in 560 gm of Fe (atomic mass $56 gmol^{-1}$) is – **[AIEEE-2003]**
(A) is twice that of 70 gm N (B) is half that of 20 gm H
(C) both are correct (D) None is correct

**Q.4** $6.02 \times 10^{20}$ molecules of urea are present in 100 ml of its solution. The concentration of urea solution is - **[AIEEE-2004]**
(A) 0.001 M (B) 0.01 M
(C) 0.02 M (D) 0.1 M
(Avogadro constant, $N_A = 6.02 \times 10^{23}\ mol^{-1}$)

**Q.5** How many moles of magnesium phosphate, $Mg_3(PO_4)_2$ will contain 0.25 mole of oxygen atoms ? **[AIEEE 2006]**
(A) $3.125 \times 10^{-2}$ (B) $1.25 \times 10^{-2}$
(C) $2.5 \times 10^{-2}$ (D) 0.02

**Q.6** In the reaction,
$2Al_{(s)} + 6HCl_{(aq)} \rightarrow 2Al^{3+}{}_{(aq)} + 6Cl^{-}{}_{(aq)} + 3H_2(g)$, **[AIEEE 2007]**
(A) 6L $HCl_{(aq)}$ is consumed for every 3L $H_2(g)$ produced.
(B) 33.6 L $H_{2(g)}$ is produced regardless of temperature and pressure for every mole Al that reacts.
(C) 67.2 L $H_{2(g)}$ at STP is produced for every mole Al that reacts.
(D) 11.2 L $H_{2(g)}$ at STP is produced for every mole $HCl_{(aq)}$ consumed.

**Q.7** Amount of oxalic acid present in a solution can be determined by its titration with $KMnO_4$ solution in the presence of $H_2SO_4$. The titration gives unsatisfactory result when carried out in the presence of HCl, because HCl – **[AIEEE 2008]**
(A) gets oxidised by oxalic acid to chlorine
(B) furnishes $H^+$ ions in addition to those from oxalic acid
(C) reduces permanganate to $Mn^{2+}$
(D) oxidises oxalic acid to carbon dioxide and water

**Q.8** A 5.2 molal aqueous solution of methyl alcohol, $CH_3OH$, is supplied. What is the mole fraction of methyl alcohol in the solution? **[AIEEE 2011]**
(A) 0.100 (B) 0.190
(C) 0.086 (D) 0.050

**Q.9** A gaseous hydrocarbon gives upon combustion 0.72 g of water and 3.08 g. of $CO_2$. The empirical formula of the hydrocarbon is – **[JEE MAIN 2013]**
(A) $C_2H_4$ (B) $C_3H_4$
(C) $C_6H_5$ (D) $C_7H_8$

**Q.10** The ratio of masses of oxygen and nitrogen in a particular gaseous mixture is 1 : 4. The ratio of number of their molecule is: **[JEE MAIN 2014]**
(A) 1 : 8 (B) 3 : 16
(C) 1 : 4 (D) 7 : 32

**Q.11** At 300 K and 1 atm, 15 mL of a gaseous hydrocarbon requires 375 mL air containing 20% $O_2$ by volume for complete combustion. After combustion the gases occupy 330 mL. Assuming that the water formed is in liquid form and the volumes were measured at the same temperature and pressure, the formula of the hydrocarbon is : **[JEE MAIN 2016]**
(A) $C_3H_8$ (B) $C_4H_8$
(C) $C_4H_{10}$ (D) $C_2H_{12}$

**Q.12** 1 gram of a carbonate ($M_2CO_3$) on treatment with excess HCl produces 0.01186 mole of $CO_2$. The molar mass of $M_2CO_3$ in g $mol^{-1}$ is: **[JEE MAIN 2017]**
(A) 11.86 (B) 1186
(C) 84.3 (D) 118.6

**Q.13** The most abundant elements by mass in the body of a healthy human adult are: Oxygen (61.4%); Carbon (22.9%), Hydrogen (10.0%); and Nitrogen (2.6%). The weight which a 75 kg person would gain if all $^1H$ atoms are replaced by $^2H$ atoms is: **[JEE MAIN 2017]**
(A) 10 kg (B) 15 kg
(C) 37.5 kg (D) 7.5 kg

**Q.14** The ratio of mass percent of C and H of an organic compound ($C_XH_YO_Z$) is 6 : 1. If one molecule of the above compound ($C_XH_YO_Z$) contains half as much oxygen as required to burn one molecule of compound $C_XH_Y$ completely to $CO_2$ and $H_2O$. The empirical formula of compound $C_XH_YO_Z$ is: **[JEE MAIN 2018]**
(A) $C_3H_4O_2$ (B) $C_2H_4O_3$
(C) $C_3H_6O_3$ (D) $C_2H_4O$

**Q.15** A solution of sodium sulfate contains 92 g of $Na^+$ ions per kilogram of water. The molality of $Na^+$ ions in that solution in mol $kg^{-1}$ is: **[JEE MAIN 2019 (JAN)]**
(A) 16 (B) 8
(C) 12 (D) 4

**Q.16** In order to oxidise a mixture one mole of each of $FeC_2O_4$, $Fe_2(C_2O_4)_3$, $FeSO_4$ and $Fe_2(SO_4)_3$ in acidic medium , the number of moles of $KMnO_4$ required is - **[JEE MAIN 2019 (APRIL)]**
(A) 3 (B) 2
(C) 1 (D) 1.5

**Q.17** The percentage composition of carbon by mole in methane is : **[JEE MAIN 2019 (APRIL)]**
(A) 80% (B) 25%
(C) 75% (D) 20%

**Q.18** The strength of 11.2 volume solution of $H_2O_2$ is :
[Given that molar mass of H = 1 g $mol^{-1}$ and O = 16 g $mol^{-1}$] **[JEE MAIN 2019 (APRIL)]**
(A) 13.6% (B) 3.4%
(C) 34% (D) 1.7%

**Q.19** For a reaction, $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ ;

identify dihydrogen ($H_2$) as a limiting reagent in the following reaction mixtures.

**[JEE MAIN 2019 (APRIL)]**

(A) 14g of $N_2$ + 4g of $H_2$ (B) 28g of $N_2$ + 6g of $H_2$

(C) 56g of $N_2$ + 10g of $H_2$ (D) 35g of $N_2$ + 8g of $H_2$

**Q.20** At 300 K and 1 atmospheric pressure, 10 mL of a hydrocarbon required 55 mL of $O_2$ for complete combustion and 40 mL of $CO_2$ is formed. The formula of the hydrocarbon is : **[JEE MAIN 2019 (APRIL)]**

(A) $C_4H_8$ (B) $C_4H_7Cl$

(C) $C_4H_{10}$ (D) $C_4H_6$

**Q.21** Amongst the following statements, that which was not proposed by Dalton was : **[JEE MAIN 2020 (JAN)]**

(A) All the atoms of a given element have identical properties including identical mass. Atoms of different elements differ in mass.

(B) Chemical reactions involve reorganisation of atoms. These are neither created nor destroyed in a chemical reaction.

(C) When gases combine or reproduced in a chemical reaction they do so in a simple ratio by volume provided all gases are at the same T & P.

(D) Matter consists of indivisible atoms.

**Q.22** The ammonia ($NH_3$) released on quantitative reaction of 0.6 g urea ($NH_2CONH_2$) with sodium hydroxide (NaOH) can be neutralized by : **[JEE MAIN 2020 (JAN)]**

(A) 100 mL of 0.2 N HCl (B) 400 mL of 0.2 N HCl

(C) 100 mL of 0.1 N HCl (D) 200 mL of 0.2 N HCl

**Q.23** 0.3 g $[ML_6]Cl_3$ of molar mass 267.46 g/mol is reacted with 0.125 M $AgNO_3$(aq) solution, calculate volume of $AgNO_3$ required in ml. **[JEE MAIN 2020 (JAN)]**

**Q.24** Ferrous sulphate heptahydrate is used to fortify foods with iron. The amount (in grams) of the salt required to achieve 10 ppm of iron in 100 kg of wheat is _____.

**[JEE MAIN 2020 (JAN)]**

Atomic weight : Fe = 55.85 ; S = 32.0 ; O = 16.00

**Q.25** The molarity of $HNO_3$ in a sample which has density 1.4 g/mL and mass percentage of 63% is _____. (Molecular Weight of $HNO_3$ = 63) **[JEE MAIN 2020 (JAN)]**

**Q.26** 5 g of zinc is treated separately with an excess of

(a) dilute hydrochloric acid and

(b) aqueous sodium hydroxide.

The ratio of the volumes of $H_2$ evolved in these two reactions is : **[JEE MAIN 2020 (JAN)]**

(A) 1 : 4 (B) 1 : 2

(C) 2 : 1 (D) 1 : 1

# EXERCISE - 5 (PREVIOUS YEARS AIPMT/NEET EXAM QUESTIONS)

**Choose one correct response for each question.**

**Q.1** The number of moles of $KMnO_4$ reduced by one mole of KI in alkaline medium is – **[AIPMT 2005]**
(A) one (B) two
(C) five (D) one fifth

**Q.2** The mass of carbon anode consumed (giving only carbondioxide) in the production of 270 kg of aluminium metal from bauxite by the Hall process is
(Atomic mass : Al = 27) **[AIPMT 2005]**
(A) 270 kg (B) 540 kg
(C) 90 kg (D) 180 kg

**Q.3** An element, X has the following isotopic composition, $^{200}X : 90\%$, $^{199}X : 8.0\%$, $^{202}X : 2.0\%$
The weighted average atomic mass of the naturally occuring element X is closest to – **[AIPMT 2007]**
(A) 201 amu (B) 202 amu
(C) 199 amu (D) 200 amu

**Q.4** What volume of oxygen gas ($O_2$) measured at 0°C and 1 atm, is needed to burn completely 1L of propane gas ($C_3H_8$) measured under the same conditions
(A) 10 L (B) 7 L **[AIPMT 2008]**
(C) 6 L (D) 5 L

**Q.5** How many moles of lead (II) chloride will be formed from a reaction between 6.5g of PbO and 3.2g of HCl ?
(A) 0.029 (B) 0.044 **[AIPMT 2008]**
(C) 0.333 (D) 0.011

**Q.6** An organic compound contains carbon, hydrogen and oxygen. Its elemental analysis gave C, 38.71% and H, 9.67%. The empirical formula of the compound would be: **[AIPMT 2008]**
(A) $CH_4O$ (B) $CH_3O$
(C) $CH_2O$ (D) CHO

**Q.7** 10 g of hdyrogen and 64g of oxygen were filled in a steel vessel and exploded. Amount of water produced in this reaction will be: **[AIPMT 2009]**
(A) 3 mol (B) 4 mol
(C) 1 mol (D) 2 mol

**Q.8** The number of atoms in 0.1 mol of a triatomic gas is: ($N_A = 6.02 \times 10^{23}\ mol^{-1}$) **[AIPMT (PRE) 2010]**
(A) $6.026 \times 10^{22}$ (B) $1.806 \times 10^{23}$
(C) $3.600 \times 10^{23}$ (D) $1.800 \times 10^{22}$

**Q.9** 25.3 g of sodium carbonate, $Na_2CO_3$ is dissolved in enough water to make 250 mL of solution. If sodium carbonate dissociates completely, molar concentration of sodium ion, $Na^+$ and carbonate ions, $CO_3^{2-}$ are respectively (Molar mass of $Na_2CO_3$ = 106 g $mol^{-1}$) **[AIPMT (PRE) 2010]**
(A) 0.955 M and 1.910 M (B) 1.910 M and 0.955 M
(C) 1.90 M and 1.910 M (D) 0.477 M and 0.477 M

**Q.10** Which has the maximum number of molecules among the following **[AIPMT (MAINS) 2011]**
(A) 44g $CO_2$ (B) 48g $O_3$
(C) 8g $H_2$ (D) 64g $SO_2$

**Q.11** $6.02 \times 10^{20}$ molecules of urea are present in 100mL of its solution. The concentration of solution is –
(A) 0.1 M (B) 0.02 M **[NEET 2013]**
(C) 0.01 M (D) 0.001M

**Q.12** Equal masses of $H_2$, $O_2$ and methane have been taken in a container of volume V at temperature 27°C in identical conditions. The ratio of the volumes of gases $H_2 : O_2$ : methane would be – **[AIPMT 2014]**
(A) 8 : 16 : 1 (B) 16:8:1
(C) 16 : 1 : 2 (D) 8 : 1 : 2

**Q.13** When 22.4 litres of $H_2(g)$ is mixed with 11.2 litres of $Cl_2(g)$, each at STP, the moles of HCl (g) formed is equal to – **[AIPMT 2014]**
(A) 1 mol of HCl (g) (B) 2 mol of HCl (g)
(C) 0.5 mol of HCl (g) (D) 1.5 mol of HCl (g)

**Q.14** 1.0 g of magnesium is burnt with 0.56 g $O_2$ in a closed vessel. Which reactant is left in excess and how much?
(At. wt. Mg = 24; O = 16) **[AIPMT 2014]**
(A) Mg, 0.16 g (B) $O_2$, 0.16 g
(C) Mg, 0.44 g (D) $O_2$, 0.28 g

**Q.15** A mixture of gases contains $H_2$ and $O_2$ gases in the ratio of 1 : 4 (w/w). What is the molar ratio of the two gases in the mixture ? **[AIPMT 2015]**
(A) 4 : 1 (B) 16 : 1
(C) 2 : 1 (D) 1 : 4

**Q.16** The number of water molecules is maximum in – **[RE-AIPMT 2015]**
(A) 18 gram of water (B) 18 moles of water
(C) 18 molecules of water (D) 1.8 gram of water

**Q.17** If avogadro number $N_A$, is changed from $6.022 \times 10^{23} mol^{-1}$ to $6.022 \times 10^{20}\ mol^{-1}$, this would change : **[RE-AIPMT 2015]**
(A) the ratio of chemical species to each other in a balanced equation.
(B) the ratio of elements to each other in a compound.
(C) the definition of mass in units of grams.
(D) the mass of one mole of carbon.

**Q.18** 20.0 g of a magnesium carbonate sample decomposes on heating to give carbon dioxide and 8.0g magnesium oxide. What will be the percentage purity of magnesium carbonate in the sample ? (Atomic weight : Mg = 24)
(A) 60 (B) 84 **[RE-AIPMT 2015]**
(C) 75 (D) 96

**Q.19** What is the mole fraction of the solute in a 1.00m aqueous solution ? **[RE-AIPMT 2015]**
(A) 0.0354 (B) 0.0177
(C) 0.177 (D) 1.770

**Q.20** What is the mass of the precipitate formed when 50 mL of 16.9% solution of $AgNO_3$ is mixed with 50 mL of 5.8% NaCl solution ? (Ag = 107.8, N = 14, O = 16, Na = 23, Cl = 35.5) **[RE-AIPMT 2015]**
(A) 7 g (B) 14 g
(C) 28 g (D) 3.5 g

**Q.21** Suppose the elements X and Y combine to form two compounds $XY_2$ and $X_3Y_2$. When 0.1 mole of $XY_2$ weighs 10 g and 0.05 mole of $X_3Y_2$ weighs 9 g, the atomic weights of X and Y are **[NEET 2016 PHASE 1]**
(A) 40, 30 (B) 60, 40
(C) 20, 30 (D) 30, 20

**Q.22** Which of the following is dependent on temperature? **[NEET 2017]**
(A) Molarity (B) Mole fraction
(C) Weight percentage (D) Molality

**Q.23** A mixture of 2.3 g formic acid and 4.5 g oxalic acid is treated with conc. $H_2SO_4$. The evolved gaseous mixture is passed through KOH pellets. Weight (in g) of the remaining product at STP will be **[NEET 2018]**
(A) 2.8 (B) 3.0
(C) 1.4 (D) 4.4

**Q.24** In which case is number of molecules of water maximum? **[NEET 2018]**
(A) 0.00224 L of water vapours at 1 atm & 273K
(B) 0.18 g of water
(C) 18 mL of water
(D) $10^{-3}$ mol of water

**Q.25** The number of moles of hydrogen molecules required to produce 20 moles of ammonia through Haber's process is : **[NEET 2019]**
(A) 10 (B) 20
(C) 30 (D) 40

# ANSWER KEY

### EXERCISE - 1

| Q | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 | 16 | 17 | 18 | 19 | 20 | 21 | 22 | 23 | 24 | 25 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| A | C | B | B | D | B | C | B | B | C | C | C | D | A | A | B | B | D | B | A | A | B | A | B | C | B |
| Q | 26 | 27 | 28 | 29 | 30 | 31 | 32 | 33 | 34 | 35 | 36 | 37 | 38 | 39 | 40 | 41 | 42 | 43 | 44 | 45 | 46 | 47 | 48 | 49 | 50 |
| A | B | C | D | C | A | D | B | D | D | B | D | C | C | B | B | D | C | A | C | A | C | A | C | B | A |
| Q | 51 | 52 | 53 | 54 | 55 | 56 | 57 | 58 | 59 | 60 | 61 | 62 | 63 | 64 | 65 | 66 | 67 | 68 | 69 | 70 | 71 | 72 | 73 | 74 | 75 |
| A | B | C | C | B | C | B | A | D | C | D | C | B | B | B | A | C | D | C | A | C | B | C | B | B | C |
| Q | 76 | 77 | 78 | | | | | | | | | | | | | | | | | | | | | | |
| A | A | B | C | | | | | | | | | | | | | | | | | | | | | | |

### EXERCISE - 2

| Q | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 | 16 | 17 | 18 | 19 | 20 | 21 | 22 | 23 | 24 | 25 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| A | A | D | C | C | A | D | A | D | A | D | D | A | B | C | D | C | B | C | B | D | A | C | C | B | D |
| Q | 26 | 27 | 28 | 29 | 30 | 31 | 32 | | | | | | | | | | | | | | | | | | |
| A | A | A | A | D | D | B | C | | | | | | | | | | | | | | | | | | |

### EXERCISE - 3

| Q | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 |
|---|---|---|---|---|---|---|---|---|---|---|
| A | 2 | 4 | 2 | 3 | 3 | 5 | 6 | 8 | 4 | 8 |

### EXERCISE - 4

| Q | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 | 16 | 17 | 18 | 19 | 20 | 21 | 22 | 23 | 24 | 25 | 26 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| A | A | A | C | B | A | D | C | C | D | D | D | C | D | B | D | B | D | B | C | D | C | A | 27 | 5 | 14 | D |

### EXERCISE - 5

| Q | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 | 16 | 17 | 18 | 19 | 20 | 21 | 22 | 23 | 24 | 25 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| A | A | C | D | D | A | B | B | B | B | C | C | C | A | A | A | B | D | B | B | A | A | A | A | C | C |

# SOLUTIONS
# SOME BASIC CONCEPTS IN CHEMISTRY

## TRY IT YOURSELF - 1

**(1)** (i) Three (ii) Four (iii) Three

**(2)** (i) 55.05 (ii) $1.755 \times 10^{10}$

(iii) 0.002343

**(3)** (i) $\dfrac{32.4 \times 0.0867}{4.238} = 0.6628 = 0.663$

As 32.4 has three significant figures so result should be expressed as 0.663 after rounding off the last digit.

(ii) $0.42 + 452.32 = 452.74$

It is correct answer as it is reported upto two decimal places.

**(4)** Contribution of Boron-10 = $10.0 \times 0.196 = 1.96$ u

Contribution of Boron-11 = $11.0 \times 0.804 = 8.844$ u

Average atomic mass of Boron = $1.96 + 8.844 = 10.804$ u

**(5)** (i) Weight of Cu = 2.16g

Weight of CuO = 2.70g

$$\% \text{ of Cu in CuO} = \frac{2.16 \times 100}{2.70} = 80\%$$

% of oxygen = 20%

(ii) Weight of CuO = 1.15g

Weight of Cu = 0.92g

$$\% \text{ of Cu in CuO} = \frac{0.92 \times 100}{1.15} = 80\%$$

% of oxygen = 20%

As the percentage composition of Cu and O in two samples is same, law is verified.

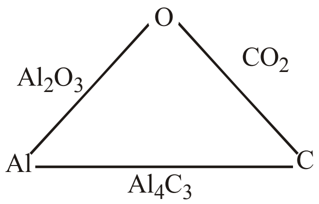
**(6)** From reaction, 2L $NH_3$ is given by 3 litre $H_2$

$\therefore$ 20 L $NH_3$ is given by $= \dfrac{3}{2} \times 20$ litre $H_2$ = 30 litre

**(7)** Atomic mass $= \dfrac{78.99 \times 24 + 10 \times 25 + 11.01 \times 26}{100} = 24.32$

**(8)** Let us fix 1g of oxygen (O) as the fixed weight.

In $Al_2O_3$, 47.1g of O are combined with Al = 52.9g




$\therefore$ 1.0g of O is combined with Al $= \dfrac{52.9}{47.1}$ g = 1.123g

In $CO_2$, 72.73g of O are combined with C = 27.27g

$\therefore$ 1.0g of O is combined with C $= \dfrac{27.27}{72.73}$ g = 0.375g

The ratio by weight of aluminium and carbon combining with oxygen in the two oxides

1.123 : 0.375.

Since the law of reciprocal proportions is true, aluminium and carbon in aluminium carbide will combine either in the same ratio or in simple multiple ratio of their weights.

$\therefore$ % of Al in $Al_4C_3$

$$= \frac{1.123}{1.123 + 0.375} \times 100 = \frac{1.123}{1.498} \times 100 = 74.97\%$$

**(9)** Let x be the percentage abundance of Carbon-12 then (100 – x) will be the percentage abundance of Carbon-13.

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

$12x + 1300 - 13x = 1201.1$

$x = 98.9$

$\therefore$ Abundance of Carbon-12 is 98.9%

**(10)** Percentage of Oxygen, $100 - (54.54 + 9.09) = 36.37\%$

| Element | % | At. wt. | Relative no. of atoms | Ratio |
|---|---|---|---|---|
| C | 54.54 | 12 | $\frac{54.54}{12} = 4.53$ | $\frac{4.53}{2.27} = 2$ |
| H | 9.09 | 1 | $\frac{9.09}{1} = 9.09$ | $\frac{9.09}{2.27} = 4$ |
| O | 36.37 | 16 | $\frac{36.37}{16} = 2.27$ | $\frac{2.27}{2.27} = 1$ |

Empirical formula = $C_2H_4O$

Empirical formula weight = 44

Molecular weight = $2 \times 88 = 176$, n = 176/44 = 4

So, molecular formula = $4 \times$ E.F. = $4(C_2H_4O) = C_8H_{16}O_4$

## TRY IT YOURSELF - 2

**(1)** Molecular mass of $CO_2$ = 44 amu

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

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

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

**(2)** Gram molecular weight of $H_2SO_4$ is 98g.

So, number of moles = 4.9/98 = 0.05

**(3)** $2KClO_3 \rightarrow 2KCl + 3O_2$

(2 × 122.5g) (3 × 32g)

245g of $KClO_3$ on heating produces 96g of oxygen

10g of $KClO_3$ will give $= \frac{96 \times 10}{245} = 3.92$ g

**(4)** 22400 ml of $O_2$ at NTP contains $= 6.02 \times 10^{23}$ molecules

1ml of $O_2$ at NTP contains $= \frac{6.02 \times 10^{23}}{22400}$ molecules

$= 2.69 \times 10^{19}$ molecules

**(5)** (i) 1 mole of a gas at NTP occupies 22.4 L volume

So, 2.5 moles will occupy = 22.4 × 2.5 L = 56 L

(ii) 1 mole of nitrogen i.e. 28g of nitrogen occupies 22.4 L

So, 14g of nitrogen will occupy 11.2 L

**(6)** From given equation,

2 mol NaOH gives $Na_2SO_4$ = 1 mole

1 mol NaOH gives $Na_2SO_4 = \frac{1}{2}$ mole

**(7)** Mass of Fe $= \frac{5.6}{100} = 5.6 \times 10^{-2}$ g

Number of atoms $= \frac{5.6 \times 10^{-2}}{56} \times 6.02 \times 10^{23}$

$= 6.02 \times 10^{20}$ atoms

**(8)** $CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + H_2O + CO_2$

(100g) (98g)

100g marble requires 98g of $H_2SO_4$

25g will need $= \frac{98}{100} \times 25 = 25.4$g of $H_2SO_4$

But $H_2SO_4$ provided is only 59% so amount required

will be $= \frac{24.5 \times 100}{50} = 49$g

**(9)** Molecular mass of $H_2O$ = 18

Weight of $6.02 \times 10^{23}$ molecules of water = 18g

Weight of one molecule $= \frac{18}{6.02 \times 10^{23}} = 2.99 \times 10^{-23}$g

**(10)** $CaCO_3 (s) \rightarrow CaO + CO_2 \uparrow$

100g 1 mole = 44g = 22.4 litre

100g of lime stone gives $CO_2$ = 22.4 litre

10g of lime stone gives $CO_2 = \frac{22.4 \times 10}{100} = 2.24$ litre

**(11)** $CaCO_3 \rightarrow CaO + CO_2$

1 mole (100g) 1 mol (44g)

Pure $CaCO_3 = \frac{40 \times 20}{100} = 8$g

From equation,

100g $CaCO_3$ gives $CO_2$ = 44g

$\therefore$ 8g $CaCO_3$ gives $CO_2 = \frac{44}{100} \times 8 = 3.52$g

**(12)** $2Pb(NO_3) \rightarrow 2PbO + 4NO_2 + O_2$

1 mol or 22.4 L $O_2$ at NTP ≡ 2 mol $Pb(NO_3)_2$

224 L $O_2$ at NTP $= \frac{2 \times 224}{22.4} = 20$ mol $Pb(NO_3)_2$

**(13)** According to the statement given above we write chemical equation as $KClO_3 (s) \rightarrow KCl (s) + O_2 (g)$

Now, balancing the chemical equation by inspection we get $2KClO_3 (s) \rightarrow 2KCl (s) + 3O_2 (g)$

From the above equation, we find that for obtaining 3 mol of oxygen, we require 2 mol of $KClO_3$.

For 2.4 mol of oxygen, we need,

2.4 mol of oxygen $\left(\frac{2\text{mol of } KClO_3}{3 \text{ mol of oxygen}}\right) = 1.6$ mol of $KClO_3$.

Molar mass of $KClO_3$ = (39g for potassium + 35.5g for chlorine + 48.0g for oxygen) = 122.5 g $mol^{-1}$.

Therefore, mass of $KClO_3$ required

= 1.6 mol × 122.5 g $mol^{-1}$ = 196.0g

**(14)** $2H_2 + O_2 \rightarrow 2H_2O$

(4g) (32g)

10g of hydrogen requires $= \frac{32 \times 10}{4} = 80$g

But oxygen provided is only 50g. The amount of oxygen product can thus react with 4 × (50/32) = 6.25g of hydrogen and water produced will be $\frac{50\times 36}{32}=56.25g$

Unreacted hydrogen $= 10-6.25 = 3.75g$

**(15)** Formula units $= \frac{\text{Mass of species}}{\text{Formula mass}}\times 6.02\times 10^{23}$

$= \frac{3}{60}\times 6.02\times 10^{23} = 3.0\times 10^{22}$ (approx.)

Number of oxygen atoms = Number of formula units × 3

$= 3.0\times\ 3\times 10^{22}$ [because one unit has 3 oxygen atom]

$= 9.0\times 10^{22}$

Charge $= 3.0\times 10^{22}\times 3.2\times 10^{-19} = 9.6\times 10^{3}$ coulomb

(One formula unit has charge $= 1.6\times 2\times 10^{-19}$

$= 3.2\times 10^{-19}$ coulomb)

## TRY IT YOURSELF - 3

**(1)** Gram molecular mass of NaOH = 40g

Molarity $= \frac{W_B}{M_B}\times\frac{1000}{V_{solution}\,(mL)} = \frac{1}{40}\times\frac{1000}{250} = 0.1\ M$

**(2)** Gram equivalents = N × volume of solution in litre

$= \frac{1}{10}\times\frac{200}{1000} = \frac{1}{50} = 0.02$ gram equivalent

**(3)** Mass of metal = x g

Mass of chlorine = mass of metal chloride – mass of metal

$= y - x$ g

$E = \frac{\text{mass of metal}}{\text{mass of } Cl_2}\times 35.5 = \frac{x}{y-x}\times 35.5$

**(4)** Molecular mass of HCl = 36.5

Molarity = 0.5 M = 0.5 Mol $L^{-1}$

Volume of solution = $250cm^3$

So, number of moles in $250cm^3 = MV_L = \frac{0.5\times 250}{1000} = \frac{0.5}{4}$

$= 0.125$ moles

Weight of HCl dissolved = number of moles of HCl × 36.5

$= 0.125\times 36.5 = 4.5625g$

**(5)** $NV = N_1V_1 + N_2V_2$

$N\times 200 = \frac{1}{10}\times 100 + \frac{1}{2}\times 100$

$N = \frac{60}{200} = 0.3\ N$

**(6)** $N\times 500 = \frac{1}{10}\times 200 - \frac{1}{100}\times 300 = \frac{17}{500} = 0.03\ N$

**(7)** $m = \frac{5.3/106}{400/1000} = \frac{0.05}{0.4} = 0.125m$

**(8)** (a) Equivalent weight of $CaCO_3$

$= \frac{\text{Mol. wt.}}{\text{Total charge on +ve ion}} = \frac{100}{2} = 50$

(b) Equivalent weight of $K_2SO_4\times Al_2(SO_4)_3\times 24\,H_2O$

$= \frac{\text{Mol. wt.}}{8}$

**(9)** Equivalent weight of $H_3PO_4 = (98/1) = 98$

(n = 1 because one H is replaced in given reaction)

Equivalent weight of $Ca(OH)_2 = (74/1) = 74$

(n = 1 because one $OH^-$ is produced by $Ca(OH)_2$ in given reaction)

**(10)** $M = \frac{1.26}{126\times(250/1000)} = 0.04\ M$

# CHAPTER-1 : SOME BASIC CONCEPTS IN CHEMISTRY

## EXERCISE-1

**(1)** **(C).** Physical and chemical properties of a compound are different from those of its constituent elements.

**(2)** **(B).** A – Kelvin, B – Celsius, C – Fahrenheit

**(3)** **(B).** Components in a mixture are not present in a fixed ratio.

**(4)** **(D).** Significant figures are meaningful digits which are known with certainty.

**(5)** **(B).** A: Average reading $= \frac{3.01+2.99}{2} = 3.0$ g

B : Average reading $= \frac{3.05+2.95}{2} = 3.0$ g

For both the students A and B, average reading is close to the correct reading (i.e., 3.0 g). Hence, both recorded accurate readings. But the readings recorded by student A are more precise as they differ only by $\pm$ 0.01, whereas readings recorded by the student B are differ by $\pm$ 0.05. Thus, the results of student A are both precise and accurate.

**(6)** **(C).** $1 \times 10^2$ has **one** significant figure.

$1.00 \times 10^2$ has three significant figures.

**(7)** **(B).** Statement I should be every experimental measurement has some amount of uncertainty associated with it.

**(8)** **(B).** $0.010100 \times 10^3$ contains 5 significant figures.

**(9)** **(C).** Significant figures are 4, 4, 4

**(10)** **(C).** $2.5 \times 1.25 = 3.125$. Since, 2.5 has two significant figures, the result should not have more than two significant figures. Hence, the answer will be 3.1

**(11)** **(C).** The amount of products formed is equal to the amount of the reactants reacted.

**(12)** **(D).**

**(13)** **(A).** According to law of conservation of mass, total mass of elements in reactants is equal to the total mass of elements in products.

**(14)** **(A).** $2KClO_3 \rightarrow 2KCl + 3O_2$

4.88 g  2.96 g  1.92 g

Since, mass of the products (2.96 + 1.92) is equal to the mass of the reactants, this illustrates the law of conservation of mass.

**(15)** **(B).**

**(16)** **(B).** 1: 2, the masses of oxygen (i.e., 16 g and 32g) which combine with a fixed mass of hydrogen (2 g) bear a simple ratio i.e.,

16 : 32 or 1 : 2.

**(17)** **(D).**

**(18)** **(B).** The reaction $C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$ is not a balanced equation hence, is not correct according to the law of conservation of mass.

**(19)** **(A).** Average atomic mass

$$= \frac{\%\text{ of I isotope} \times \text{its atomic mass} + \%\text{ of II isotope} \times \text{its atomic mass}}{100}$$

$$= \frac{75.53 \times 34.969 + 24.47 \times 36.96}{100} = 35.5 \text{ amu.}$$

**(20)** **(A).** 'amu' has been replaced by 'u' which is known as unified mass.

**(21)** **(B).** Average atomic weight

$$= \frac{54 \times 5 + 56 \times 90 + 57 \times 5}{100} = 55.95$$

**(22)** **(A).** The molar mass of the naturally occurring argon is equal to the weighted arithmetic mean of the various isotopes present in it.

$$\text{Molar mass of argon} = \frac{0.337 \times 35.96755 \text{g/mol} + 0.063 \times 37.96272 \text{g/mol} + 99.6 \times 39.9624 \text{g/mol}}{100}$$

$$= \frac{12.121 \text{g/mol} + 2.392 \text{g/mol} + 3980.255 \text{g/mol}}{100} = \frac{3994.768 \text{g/mol}}{100}$$

= 39.948 g/mol

**(23)** **(B).** The mass of one mole of a substance in grams is called its molar mass.

**(24)** **(C).** 1 mol of $C_6H_{12}O_6$ has = 6 N atoms of C

$\therefore$ 0.35 mol of $C_6H_{12}O_6$ has

= $6 \times 0.35$ N atoms of C = 2.1 N atoms

= $2.1 \times 6.023 \times 10^{23} = 1.26 \times 10^{24}$ carbon atoms

**(25)** **(B).** $\because$ 180 gm glucose has = N molecules

$\therefore$ 5.23 gm glucose has

$$= \frac{5.23 \times 6.023 \times 10^{23}}{180} = 1.75 \times 10^{22} \text{ molecules}$$

**(26)** **(B).** $\because$ $6.023 \times 10^{23}$ molecules of $NH_3$ has weight = 17 gm

$\therefore$ $3.01 \times 10^{23}$ molecules of $NH_3$ has weight

$$= \frac{17 \times 3.01 \times 10^{23}}{6.023 \times 10^{23}} = 8.50 \text{ gm}$$

**(27)** **(C).** Number of molecules of gas at STP

$$= \frac{6.023 \times 10^{23} \times 2.8}{22.4} = 7.5 \times 10^{22} \text{ molecules}$$

Number of atoms in diatomic molecule

$= 2 \times 7.5 \times 10^{22} = 15 \times 10^{22}$ atoms

**(28)** **(D).** (A) 1 mol of He = 4g = $N_A$ atoms

or 4 g of He = 4/4 mol = 1 $N_A$ atom

(B) 1 mole of Na = 23g = $N_A$ atoms

$\left[\text{Number of moles} = \frac{\text{Given mass}}{\text{Atomic mass}}\right]$

$\therefore$ 46 g of Na $= \frac{46}{23}$ mol $= \frac{46}{23} N_A$ atoms

$= 2 N_A$ atoms

(C) 1 mole of Ca = 40g = $N_A$ atoms

$\therefore$ 0.40 g of Ca $= \frac{0.40}{40}$ mol $= \frac{0.40}{40} N_A$ atoms

$= 0.01 N_A$ atoms

(D) 1 mole of He = 4g = $N_A$ atoms

$\therefore$ 12 g of He $= \frac{12}{4}$ mol $= \frac{12}{4} N_A$ atoms $= 3 N_A$ atoms

**(29) (C).**

(i) 1 g of Au (s) = $\frac{1}{197}$ mol of Au (s)

$= \frac{6.022 \times 10^{23}}{197}$ atoms of Au (s)

$= 3.06 \times 10^{21}$ atoms of Au (s)

(ii) 1 g of Na (s) = $\frac{1}{23}$ mol of Na (s)

$= \frac{6.022 \times 10^{23}}{23}$ atoms of Na (s)

$= 26.2 \times 10^{21}$ atoms of Na (s)

(iii) 1 g of Li (s) = $\frac{1}{7}$ mol of Li (s)

$= \frac{6.022 \times 10^{23}}{7}$ atoms of Li (s)

$= 86.0 \times 10^{21}$ atoms of Li (s)

(iv) 1 g of $Cl_2$ (g) = $\frac{1}{71}$ mol of $Cl_2$ (g)

$= \frac{6.022 \times 10^{23}}{71}$ atoms of $Cl_2$ (g)

$= 8.48 \times 10^{21}$ atoms of $Cl_2$ (g)

(Molar mass of $Cl_2$ molecule = $35.5 \times 2 = 71$ g $mol^{-1}$)

**(30) (A).** Mass of $Al_2O_3 = 2 \times 27 + 3 \times 16 = 102$

0.051 g of $Al_2O_3 = \frac{0.051}{102} = 0.0005$ mol

1 mol of $Al_2O_3$ contains $2 \times 6.023 \times 10^{23}$ $Al^{3+}$ ions

0.0005 mol of $Al_2O_3$ contains

$2 \times 0.0005 \times 6.023 \times 10^{23}$ $Al^{3+}$ ions

$= 6.023 \times 10^{20}$ $Al^{3+}$ ions

**(31) (D).** Amount of ethane ($C_2H_6$) = 3 mol

1 mol of $C_2H_6$ = 2 mol of C atoms = 6 mol of H atoms

(i) No. of moles of C atoms $= 2 \times 3$ mol = 6 mol

(ii) No. of moles of hydrogen atoms = $6 \times 3$ mol = 18 mol

(iii) No. of molecules of ethane

$= 3$ mol $\times 6.023 \times 10^{23}$ molecules/mol

$= 18.069 \times 10^{23}$ molecules.

**(32) (B).** $2H_2O \rightarrow 2H_2 + O_2$

$2 \times 18 = 36$ g

36 g of water produces 1 mole of $O_2$ gas.

180 g of water will produce 180/36=5moles of $O_2$ gas.

**(33) (D).** The empirical formula of boron trifluoride is $BF_3$.

Grams = (moles) (MW)

Grams of boron = (1 mol) (10.8 g/mol) = 10.8 g

Grams of fluorine = (3 mol)(19.0 g/mol) = 57.0 g

So the mass ratio is about 57 to 11, which is about 5.3 to 1.

**(34) (D).** Moles = $\frac{\text{grams}}{\text{MW}}$

Moles of calcium hydroxide = $\frac{148 \text{ g}}{74 \text{ g/mol}}$ = 2 moles

Every mole of $Ca(OH)_2$ contains 2 moles of oxygen.

So, there are (B) (B) = 4 moles of oxygen

Grams = (moles) (MW)

So, grams of oxygen = (4 mol)(16 g/mol) = 64 grams

**(35) (B).** You might be able to do this one in your head just from knowing that sulphur's molecular weight is twice as large as oxygen's. If not, let's say you have 100 grams of the compound. So you have 50 grams of sulphur and 50 grams of oxygen.

Moles = $\frac{\text{grams}}{\text{MW}}$

Moles of sulphur = $\frac{50 \text{ g}}{32 \text{ g/mol}}$ = a little more than 1.5

Moles of oxygen = $\frac{50 \text{ g}}{16 \text{ g/mol}}$ = a little more than 3

The molar ratio of O to S is 2 to 1, so the empirical formula must be $SO_2$.

**(36) (D).** Let's say we have 100 grams of the compound.

Moles = $\frac{\text{grams}}{\text{MW}}$

So moles of carbon = $\frac{80\text{g}}{12 \text{ g/mol}}$ = 6.7 moles

and moles of hydrogen = $\frac{20\text{g}}{1 \text{ g/mol}}$ = 20 moles

According to calculation, there are about three times as many moles of hydrogen in the compound as there are moles of carbon, so the empirical formula is $CH_3$. The molar mass for the empirical formula is 15 g/mol, so we need to double the moles of each element to get a compound with a molar mass of 30 g/mol. That makes the molecular formula of the compound $C_2H_6$.

**(37) (C).** Let's say we have 100 grains of the compound.

Moles = $\frac{\text{grams}}{\text{MW}}$

So moles of carbon = $\frac{75 \text{ g}}{12 \text{ g/mol}}$ = 6 moles

and moles of hydrogen = $\frac{25 \text{ g}}{1 \text{ g/mol}}$ = 25 moles

According to our rough calculation, there are about four times as many moles of hydrogen in the compound as there are moles of carbon, so the empirical formula is $CH_4$.

**(38) (C).**

| Element | % | %/At. wt. | Ratio |
|---|---|---|---|
| N | 30.5 | $\frac{30.5}{14} = 2.18$ | 1 |
| O | 69.5 | $\frac{69.5}{16} = 4.34$ | 2 |

Empirical formula = $NO_2$

Empirical formula weight = 46

$n = \frac{92}{46} = 2$. Molecular formula = $(NO_2)_2 = N_2O_4$

**(39) (B).**

| Element | %age | Atomic mass | Molar ratio | Simpler molar ratio |
|---|---|---|---|---|
| C | 10.06% | 12 | $\frac{10.06}{12}=0.84$ | $\frac{0.84}{0.84}=1$ |
| H | 0.84% | 1 | $\frac{0.84}{1}=0.84$ | $\frac{0.84}{0.84}=1$ |
| Cl | 89.10% | 35.5 | $\frac{89.10}{35.5}=2.5$ | $\frac{2.5}{0.84}=3$ |

Thus, the empirical formula of the substance is $CHCl_3$.

**(40) (B).** Molecular formula of ethanol is : $C_2H_5OH$
Molar mass of ethanol is :
$(2\times 12.01+6\times 1.008+16.00)\text{ g}=46.068\text{g}$

Mass per cent of carbon$=\frac{24.02\text{ g}}{46.068\text{g}}\times 100=52.14\%$

**(41) (D).**

| Element | % | No. of moles | Mole ratio | Whole number ratio |
|---|---|---|---|---|
| P | 27.3 | 27.3/12 = 2.27 | 1 | 1 |
| Q | 72.7 | 72.7/16 = 4.54 | 2 | 2 |

Empirical formula = $PQ_2$

**(42) (C).** 25.4 g $I_2$ combines with 8 g oxygen
$\therefore$ 254 g iodine will combine with 80 g oxygen
$\therefore$ Formula of oxide iodine would be $I_2O_5$

**(43) (A).** Weight of empirical formula
$CH_2=12+(1\times 2)=12+2=14$
Mass of one mole of the compound = its molecular weight = 42

$$n=\frac{\text{mol. wt.}}{\text{empirical formula wt.}}=\frac{42}{14}=3$$

$\therefore$ Mol. formula = (empirical formula) × n
$=(CH_2)\times 3=C_3H_6$

**(44) (C).** $n=\frac{\text{Molecular mass}}{\text{Empirical formula mass}}=\frac{180}{30}=6$

$\because$ Empirical formula mass of $CH_2O=30)$
$\therefore$ Molecular formula $=6\times CH_2O=C_6H_{12}O_6$

**(45) (A).** First, you have to remember that 1 liter of water has a mass of 1 kg.
Moles = (molarity) (liters)
Moles of NaCl = (0.5 M) (1 L) = 0.5 moles
Grams = (moles) (MW)
Grams of NaCl = (0.5 mol) (59 g/ mol)
= 30 g

**(46) (C).** The molecular weight of $CuSO_4$ is 160g/mol, so we have only 1 mole of the hydrate. The lost mass was due to water, so 1 mole of the hydrate must have contained 90 grams of $H_2O$.

$$\text{Moles}=\frac{\text{grams}}{\text{MW}}$$

$$\text{Moles of water}=\frac{90\text{ g}}{18\text{g/mol}}=5\text{ moles}$$

So if 1 mole of hydrate contains 5 moles of $H_2O$, then the formula for the hydrate must be $CuSO_4 . 5H_2O$.

**(47) (A).** $\text{Moles}=\frac{\text{grams}}{\text{MW}}$

$$\text{Moles of }ZnSO_3=\frac{145\text{ g}}{145\text{ g/mol}}=1\text{ mole}$$

From the balanced equation, when 1 mole of $ZnSO_3$ is consumed, 1 mole of $SO_2$ will be produced.
So about 1 mole of $SO_2$ is produced.
Liters = (moles) (22.4 L/mol)
Liters of $SO_2$ = (about 1 mol) (22.4 L)
= 23 liters

**(48) (C).** $\text{Moles}=\frac{\text{grams}}{\text{MW}}$

$$\text{Moles of }CaCO_3=\frac{150\text{ g}}{100\text{ g/mol}}=1.5\text{ moles}$$

From the balanced equation, for every mole of $CaCO_3$ consumed, one mole of $CO_2$ is produced.
So 1.5 moles of $CO_2$ are produced.
At STP, volume of gas = (moles)(22.4 L)
So volume of $CO_2$ = (1.5) (22.4) = 34 L

**(49) (B).** $H_2 + Cl_2 \rightarrow 2\,HCl$

Volume before reaction
8 lit 6 lit 0
Volume after reaction
2 0 12
$\therefore$ Volume after reaction
= Volume of $H_2$ left + Volume of HCl formed
= 2 + 12 = 14 lit

**(50) (A).** $X_{C_2H_5OH}=0.040$ ; $X_{H_2O}=0.96$

$$\frac{X_{C_2H_5OH}}{X_{H_2O}}=\frac{0.04}{0.96}=\frac{1}{24}\ \therefore\ \frac{n_{C_2H_5OH}}{n_{H_2O}}=\frac{1\text{ mol}}{24\text{ mol}}$$

$\therefore$ Total wt. of solution = 46 + 24 × 18 = 478g
d = 1 $\therefore$ V = 478ml

$$\text{Molarity}=\frac{n_{C_2H_5OH}}{V}=\frac{1\times 1000}{478}=2.09M$$

**(51) (B).** $\underset{1\text{mol}}{M}+\underset{6\text{mol}}{6F}\rightarrow MF_6$

Given wt. of $MF_6$ = 0.547 g
Weight of M = 0.250 g
$\therefore$ Weight of F combined with
M = 0.547 – 0.250 = 0.297 g
$\because$ 0.297 g of F combine with 0.250 g of M
$\therefore$ 6 mol (6 × 19 g) of F combine with

$$\frac{0.250}{0.297}\times(6\times 19)=95.96\approx 96$$

and According to stoicheometry 6 mol combine with 1 mol of M i.e. at. wt of M, therefore At. wt of M is 96.

**(52)** **(C).** The normality of 1(M) $H_2SO_4 = 2(N)$
The normality of 1(M) $H_3PO_3 = 2(N)$
The normality of 1(M) $H_3PO_4 = 3(N)$
The normality of 1(M) $HNO_3 = 1(N)$

**(53)** **(C).** $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$

No. of moles of $AgNO_3 = \frac{3.4}{170} = 0.02$

No. of moles of $NaCl = \frac{5.85}{58.5} = 0.1$

Limiting reagent = $AgNO_3$
1 mole of $AgNO_3$ produces 1 mole of AgCl
0.02 mole of $AgNO_3$ will produce 0.02mole of AgCl
Weight of AgCl produced = 0.02 × 143.5 = 2.870 g

**(54)** **(B).** 69 g acid in 100 g,
volume of 100 g sample = 100/d

$$\text{Molarity} = \frac{69}{63} \times \frac{1}{100/d} \times 1000$$

$$15.44 = \frac{69}{63} \times \frac{1}{100/d} \times 1000$$

$$d = \frac{63 \times 15.44}{69 \times 10} = 1.409 \text{ g/cc}$$

**(55)** **(C).** $\text{Molarity} = \frac{W_B \text{ (in g)}}{M_B \text{ (in g mol}^{-1}) \times \text{Volume of solution (in L)}}$

$$= \frac{\text{Conc. (in g L}^{-1})}{M_B \text{ (in g mol}^{-1})} = \frac{0.9}{180} = 0.005 \text{ M}$$

(Molar mass of $C_6H_{12}O_6 = 180 \text{ g mol}^{-1}$)

**(56)** **(B).** $M_1V_1 = M_2V_2$
$0.5 \times 100 = 0.1 \times V_2 ; V_2 = 500 \text{ cm}^3$
Volume of water to be added to $100\text{cm}^3$ of solution = $500 - 100 = 400 \text{ cm}^3$

**(57)** **(A).** $\text{Molality} = \frac{\text{weight of solute}}{\text{mol. weight of solute}} \times \frac{1000}{\text{weight of solvent}}$

$$m = \frac{10}{40} \times \frac{1000}{100} = 2.5 \text{ m}$$

**(58)** **(D).** 200 kg of 95% pure means

$$\frac{95}{100} \times 200 = 190 \text{ kg}$$

$$\underset{\substack{(40+12+3\times16)\\100}}{CaCO_3} \xrightarrow{\Delta} \underset{\substack{(40+16)\\56}}{CaO} + CO_2$$

∵ 100 g of$CaCO_3$ on heating gives lime = 56g
∴ 190 g of $CaCO_3$ on heating will give lime

$$= \frac{56 \times 190}{100} = 106.4\text{g}$$

**(59)** **(C).** $\text{Molarity (M)} = \frac{\text{No. of moles of solute}}{\text{Volume of solution (in L)}}$

or $\text{Molarity} = \frac{W_B \times 100}{M_B \times \text{Volume of solution (in mL)}}$

$$= \frac{5.85 \times 1000}{58.5 \times 500} = 0.2 \text{ mol L}^{-1}$$

(∵ Molar mass of NaCl = $58.5 \text{ g mol}^{-1}$)

**(60)** **(D).** $\text{Molality (m)} = \frac{\text{No. of moles of solute}}{\text{Mass of solvent (in kg)}}$

$$\text{Molality} = \frac{W_B \times 1000}{M_B \times W_A \text{ (in g)}} = \frac{18.25 \times 1000}{36.5 \times 500} = 1\text{m}$$

(Molar mass of HCl = $36.5 \text{ g mol}^{-1}$)

**(61)** **(C).** A solution is prepared by adding 2 g of a substance A to 18g of water.

$$\text{Mass \% of A} = \frac{\text{Mass of A}}{\text{Mass of solution}} \times 100$$

$$= \frac{2\text{g}}{2 \text{ g of A} + 18 \text{ g of water}} \times 100$$

$$= \frac{2\text{g}}{20 \text{ g}} \times 100 = 10\%$$

**(62)** **(B).** $V_{strength} = 56$, ∴ $M = \frac{28}{11.2} = 2.5$

∴ 1L contain 2.5 moles of $H_2O_2$
or $2.5 \times 34 = 85\text{g } H_2O_2$
Weight of 1 litre solution = 265g
(∵ d = 265g/L)

∴ $W_{H_2O} = 180\text{g}$ or moles of $H_2O = 10$

$$x_{H_2O_2} = \frac{2.5}{2.5 + 10} = 0.2 ; \%\frac{w}{v} = \frac{2.5 \times 34}{1000} \times 100 = 8.5$$

$$m = \frac{2.5}{180} \times 1000 = 13.88$$

**(63)** **(B).** Molar mass of $CuSO_4 = 63.5 + 32 + 64 = 159.5$

$$\text{Moles of } CuSO_4 = \frac{80}{159.5} = 0.50$$

Volume of solution = 3 L

$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Volume of solution in L}} = \frac{0.50}{3}$$
$$= 0.167 \text{ mol L}^{-1}$$

**(64)** **(B).** According to the chemical equation
$CH_4 (g) + 2O_2 (g) \rightarrow CO_2 (g) + 2H_2O (g)$
44 g $CO_2$ (g) is obtained from 16 g $CH_4$(g)
[∵ 1 mol $CO_2$(g) is obtained from 1 mol of $CH_4$(g)]

$$\text{Mole of } CO_2 \text{ (g)} = \frac{22 \text{ g } CO_2\text{(g)}}{44 \text{ g } CO_2\text{(g)}} = 0.5 \text{ mol } CO_2 \text{ (g)}$$

Hence, 0.5 mol $CO_2$ (g) would be obtained from 0.5mol $CH_4$ (g) or 0.5 mol of $CH_4$(g) would be required to produce 22 g $CO_2$(g).

**(65)** **(A).** Molarity

$$= \frac{\text{Wt. of solute}}{\text{Mol. wt. of solute}} \times \frac{1000}{\text{Volume of soln. (mL)}}$$

$$= \frac{49}{98} \times \frac{1000}{250} = 2M$$

**(66)** **(C).** The molarity of pure water is 55.6 M which can be calculated by $\frac{1000\ (\text{Vol. in mL})}{18\ (\text{molar mass})}$.

**(67)** **(D).** $C_xH_y(g) + O_2(g) \to CO_2(g) + H_2O(g)$
(Hydrocarbon) ↓ ↓
10 mL 40 mL 50 mL
or 1 vol. 4 vol. 5 vol.
$C_xH_y + 13O_2 \to 4CO_2 + 5H_2O$
Comparing both sides, $x = 4, y = 10$

**(68)** **(C).** $2H_2S + SO_2 \to 3S + 2H_2O$
2 mol 1 mol
Given 0.025 mol 0.025 mol

After reaction 0 $(0.025 - \frac{0.025}{2})$

so $SO_2$ will remains in excess

Given $SO_2 \to \frac{1.6}{64} = 0.025$ mol

$$H_2S = \frac{1.5 \times 10^{22}}{6.02 \times 10^{23}} = 0.025 \text{ mol}$$

**(69)** **(A).** 0.25 mole of $O_2 = 0.25 \times 32 = 8$ gm of $O_2$

gm equivalent $= \frac{W}{E} = \frac{8}{8} = 1$

**(70)** **(C).** In $S_2$ $Cl_2$
Ratio of wt.= $2 \times 32$ $2 \times 35.5$
32 35.5
∴ E (s) = 32

**(71)** **(B).** Here gm equivalent of metal = gm equivalent of $H_2$

$$\frac{W_1}{E_2} = \frac{W_2}{E_1}\ ;\ \frac{12}{24} = \frac{W_2}{1}\ ;\ W_2 = 0.5 \text{ gm}$$

2 gm $H_2$ = 22.4 L ; 0.5 gm $H_2$ = 5.6 L

**(72)** **(C).** gm equivalent of metal carbonate
= gm equivalent of $H_2SO_4$

$$\frac{W}{E} = N \times V\,(L)\ ;\ \frac{0.84}{E} = \frac{1}{2} \times \frac{40}{1000}$$

**(73)** **(B).** Equivalent weight of Br = 80
8.89 gm Br is combines with 1gm of a metal

∴ 80 gm Br is combines with $\frac{1}{8.99} \times 80$ of a metal
(Equivalent weight of metal)
(Elements combines in the ratio of their volumes)

**(74)** **(B).** gm equivalent of $NaH_2PO_4 \equiv$ gm equivalent of NaOH

$\frac{W}{E} = N \times V$ [$NaH_2PO_4$ is acidic salt & its V.F = 2,

$E = \frac{120}{2} = 60$]

**(75)** **(C).** Weight of metal hydride = 0.84 gm
Weight of hydrogen = 0.04 gm
∴ Weight of metal = 0.84 – 0.04 = 0.80 gm.

$$\text{Equivalent weight of metal} = \frac{\text{weight of metal}}{\text{weight of } H_2}$$

**(76)** **(A).** If weight of oxide = 100
then weight of oxygen = 20
Weight of element = 100 – 20 = 80

$$\text{Equivalent weight of element} = \frac{\text{Weight of element}}{\text{Weight of oxygen}} \times 8$$

**(77)** **(B).** One gm of hydrogen is combined with 80gm of Br.
∴ Equivalent weight of Br = 80
and we know that element combines in the ratio of their equivalent weight
4 gm of Br combines with 1 gm of Ca
∴ 80 (Equivalent weight) of Br combines with 20 gm of Ca (Equivalent weight)

**(78)** **(C).** $\text{Equivalent weight} = \frac{\text{weight of metal}}{\text{weight of oxygen}} \times 8$

## EXERCISE-2

**(1)** **(A).** ∵ N atoms of Ag weigh 108 gm
∴ 1 atom of Ag weigh

$$= \frac{108}{N} = \frac{108}{6.023 \times 10^{23}}$$

$= 17.93 \times 10^{-23}$ gm.

**(2)** **(D).** ∵ 22.4 litre water vapour at STP has
$= 6.023 \times 10^{23}$ molecules
∴ $1 \times 10^{-3}$ litre water vapours at STP has

$$= \frac{6.023 \times 10^{23}}{22.4} \times 10^{-3} = 2.69 \times 10^{+19}$$

**(3)** **(C).** ∵ $10^6$ rupees are spent in 1sec.
∴ $6.023 \times 10^{23}$ rupees are spent in

$$= \frac{1 \times 6.023 \times 10^{23}}{10^6} \text{ sec} = \frac{1 \times 6.023 \times 10^{23}}{10^6 \times 60 \times 60 \times 24 \times 365} \text{ years}$$

$= 19.098 \times 10^9$ year

**(4)** **(C).** ∵ Weight of 1 atom of element $= 6.644 \times 10^{-23}$ gm
∴ Weight of 'N' atoms of element
$= 6.644 \times 10^{-23} \times 6.023 \times 10^{23} = 40$ gm
∵ 40 gm of element has 1 gm atom.

∴ $40 \times 10^3$ gm of element has $\frac{40 \times 10^3}{40} = 10^3$ gm atom.

**(5)** **(A).** ∵ Molecular weight of $CaCl_2$ = 111 g
∵ 111 g $CaCl_2$ has = N ions of $Ca^{+2}$

∴ 222g of $CaCl_2$ has $\frac{N \times 222}{111} = 2N$ ions of $Ca^{+2}$

Also 111 g $CaCl_2$ has = 2N ions of $Cl^-$

∴ 222 g $CaCl_2$ has $= \frac{2N \times 222}{111}$ ions of $Cl^-$
= 4N ions of $Cl^-$.

**(6)** **(D).** $\because$ 100 kg impure sample has pure $CaCO_3 = 95$ kg

$\therefore$ 200 kg impure sample has pure $CaCO_3$

$$= \frac{95 \times 200}{100} = 190 \text{ kg.}$$

$CaCO_3 \to CaO + CO_2$

$\because$ 100 kg $CaCO_3$ gives CaO = 56 kg.

$\therefore$ 190 kg $CaCO_3$ gives $CaO = \frac{56 \times 190}{100} = 106.4$ kg.

**(7)** **(A).** $\because$ 1.429 gm of $O_2$ gas occupies volume = 1 litre.

$\therefore$ 32 gm of $O_2$ gas occupies $= \frac{32}{1.429} = 22.4$ litre/mol.

**(8)** **(D).**

(a) 1 mole of He contain $6.02 \times 10^{23}$ atoms

$\therefore$ 52 mole of He contain $= 52 \times 6.02 \times 10^{23}$

$= 31.3 \times 10^{24}$ atoms

(b) Atomic weight of He = 4amu

$\therefore$ 52 amu of He contain $= \frac{52}{4} = 13$ atoms of He

(c) Number of moles of He in 52g $= \frac{52}{4} = 13$ moles

$\therefore$ Number of atoms in 52g of He i.e. 13 moles

$= 13 \times 6.02 \times 10^{23}$ atoms $= 78.26 \times 10^{23}$ atoms

**(9)** **(A).** Given mole $= \frac{160}{64} = 2.5$

Remove number of moles $= \frac{1.2046 \times 10^{24}}{6.023 \times 10^{23}} = 2$

So, remaining moles = 0.5

Remaining volume at STP = 0.5 × 22.4 = 11.2 litre.

**(10)** **(D).** 0.25 M $Ba(OH)_2 = 0.5\ Ba(OH)_2$

25mL of 0.5 N $Ba(OH)_2$ neutralize acid = 1.25g

$\therefore$ 1000mL of 1N $Ba(OH)_2$ will neutralize acid

$$= \frac{1.25}{25} \times \frac{1000}{0.5} = 100\text{g}$$

$\therefore$ eq. mass of the acid = 100

and mol. mass = 100 × 2 = 200

**(11)** **(D).** % of $C = \frac{12}{44} \times \frac{0.44}{0.30} \times 100 = 40 = \frac{40}{12} = 3.33 = 1$

% of $H = \frac{2}{18} \times \frac{0.18}{0.30} \times 100 = 6.66 = 2$

% of O = 100 = (%C + %H) = 100 – (40 + 6.66)

$= 53.34\% \quad = \frac{53.34}{16} = 3.33 = 1$

$E.F = CH_2O = M.F = n \times E.F$ ; n = 60/30 = 2

$M.F. = 2 \times CH_2O = C_2H_4O_2$

**(12)** **(A).** At NTP vol. of 1 mol of $O_2 = 22.4$ L

& 1 mol of $O_2 = 32$ g

Density $= \frac{32}{22.4}$ g / L

**(13)** **(B).** Mol. wt of AgCl = 108 + 35.5 = 143.5

$\because$ 143.5 g AgCl contains 108 g of Ag

$\therefore$ 7.20 g AgCl contains $\frac{108}{143.5} \times 7.20$ g of Ag

= 5.42 g of Ag

It means wt of pure Ag in silver with is 5.42g

$\therefore$ % of Ag in coin $= \frac{5.42}{5.82} \times 100 = 93.12\%$

**(14)** **(C).** Per mole of the formation

$NaBrO_3$ the x factor = + 5

$\therefore$ Equivalent weight of $NaBrO_3 = \frac{\text{Mol. wt}}{5}$

**(15)** **(D).** Molecular weight of $NaHCO_3 = 23 + 1 + 12 + 48 = 84$

Molecular weight of $Na_2CO_3 = 46 + 12 + 48 = 106$

Hence, total weight = 84 + 106 = 190

In 190 g of a mixture, weight of $Na_2CO_3$ is = 106

$\therefore$ In 19 g of a mixture weight of

$$Na_2CO_3 = \frac{106 \times 19}{190} = 10.6 \text{ g}$$

**(16)** **(C).** $\underset{1\text{ mol}}{C\ (s)} + O_2(g) \longrightarrow \underset{\substack{1\text{ mol} \\ = 6.023 \times 10^{23}}}{CO_2(g)}$

$\because$ 1 mole of graphite on complete combustion gives $CO_2 = 6.023 \times 10^{23}$ molecules

$\therefore$ 0.1 mole of graphite will give $CO_2$

$$= \frac{6.023 \times 10^{23} \times 0.1}{1} = 6.023 \times 10^{22}$$

**(17)** **(B).** Lower the number of moles, down is the pressure

$\therefore$ Number of moles in 0.071 g $= \frac{0.071}{71} = 0.001$

Number of moles in 0.0355 g $= \frac{0.0355}{71} = 0.0005$

Number of mole = 0.02

Number of moles with $6.023 \times 10^{21}$ molecules

$$= \frac{6.023 \times 10^{21}}{6.023 \times 10^{23}} = 0.01$$

**(18)** **(C).** Molarity $= \frac{\text{wt. of solute}}{\text{mol. wt.}} \times \frac{100}{\text{vol. of solution (in ml.)}}$

(Volume of solution $= \frac{\text{mass}}{\text{density}} = \frac{100}{1.84} = 54.34$ ml)

Now molarity $= \frac{98}{98} \times \frac{1000}{54.34} = 18.4$ M

**(19)** **(B).** Number of equivalents of HCl remaining after adding

$50\text{cm}^3$ of 0.1 N NaOH $= \frac{0.2 \times 50 - 0.1 \times 50}{100} = \frac{0.5}{100}$

$\therefore$ Volume of 0.5 N KOH required $\frac{0.5}{100}$ eq $\equiv \frac{V \times 0.5}{1000}$

$$V = \frac{0.5}{100} \times \frac{1000}{0.5} = 10 \text{ cm}^3$$

**(20)** **(D).** $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(\ell)$
20 ml 40 ml 20 ml
Volume of oxygen unreacted = 50 – 40 = 10 ml
Total volume of the gas left after cooling
= 10 ml of $O_2$ + 20 ml of $CO_2$ = 30 ml

**(21)** **(A).** $CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$
0.1 0.1
$2CH_3COONa \rightarrow C_2H_6 + NaOH + CO_2 + H_2$
0.01 0.005
∴ Number of gram equivalent of

$CH_3COONa = \frac{100 \times 0.1}{1000} = 0.01$

∴ 0.01 mole of $CH_3COONa$ gives 0.005 mole of $C_2H_6$
1 mole of $C_2H_6$ = 22400 $cm^3$ at STP
∴ 0.005 mole of $C_2H_6$ = ?

∴ Volume of $C_2H_6 = \frac{0.005 \times 22400}{1}$
$= 112\ cm^3$

**(22)** **(C).** 18 ml of water contains $10 \times 6.022 \times 10^{23}$ electrons
$= 6.022 \times 10^{24}$ electrons

**(23)** **(C).** $V_1N_1 = V_2N_2$
Oxalic acid KMnO$_4$
0.1 M = 0.2 N
Oxalic acid
$V_1 \times 0.2 = 20 \times 0.125$ 0.025 = 0.125 N
$KMnO_4$
Volume of oxalic acid = 12.5 ml

**(24)** **(B).** Atomic mass = 20 × 2 = 40

$\text{Valency} = \frac{\text{Atomic mass}}{\text{Eq. mass}} = \frac{40}{20} = 2$

Molecular Formula of its chloride = $MCl_2$
∴ Molecular Mass = 40 + 71 = 111

**(25)** **(D).** Number of atoms $= \frac{6.022 \times 10^{23} \times 0.018}{18}$
$= 6.022 \times 10^{20}$

**(26)** **(A).** Empirical formula of all the 4 options is $CH_2O$. Since molecular mass is 90.
Molecular formula is $C_3H_6O_3$.

**(27)** **(A).** Mass of 22400 $cm^3$ of $NH_3$ = 17

Man of 112 $cm^3$ of $NH_3 = \frac{17 \times 112}{22400} = 0.085$ g

**(28)** **(A).** Let the mass of CaO = x g ; BaO = 10 – xg

$\therefore \frac{10 - x}{76.5} + \frac{x}{28} = \frac{100 \times 2.5}{1000}$
[As the Eq mass of BaO = 153/2
= 76.5 × CaO = 56/2 = 28]
280 – 28x + 76.5x = 0.25 (76.5) (28)
48.5x = 535.5 – 280

$48.5x = 255.5;\ x = \frac{255.5}{48.5} = 5.26$

∴ Percentage of CaO $= \frac{5.26}{10} \times 100 = 52.6$

**(29)** **(D).** Oxalic acid, $\frac{25 \times N}{1000} = \frac{0.064}{40}$

$N = \frac{0.064 \times 1000}{40 \times 25} = 0.064$

∴ Molarity $= \frac{0.064}{2} = 0.032$

**(30)** **(D).** 5.5 mg in 180 g → 1 atm
∴ 5 atm pressure requires 5.5 mg × 5 = 27.5 mg of $N_2$

∴ Mole fraction of $N_2 = \frac{\frac{27.5 \times 10^{-3}}{28}}{\frac{180}{18}} = \frac{10^{-3}}{10} = 1 \times 10^{-4}$

**(31)** **(B).** 0.04 M $K_2Cr_2O_7$ = 0.24 N $K_2Cr_2O_7$
0.03 M $KMnO_4$ = 0.15 N $KMnO_4$
$0.24 \times 50 = 0.15 \times V_2$

$V_2 = \frac{0.24 \times 50}{0.15} = 80$ml

**(32)** **(C).** Molarity of urea solution $= \frac{0.06 \times 10}{60} = 0.01$ M

## EXERCISE-3

**(1)** **2.** $PbO_2 + 4KI \longrightarrow PbI_2 + I_2 + 2K_2O$

$N_2H_4 + 2I_2 \longrightarrow N_2 + 4HI$

moles of $N_2$ liberated $= \frac{1.12}{22.4} = 5 \times 10^{-2}$ mole

sol mole of $I_2$ reacted $= 5 \times 10^{-2} \times 2 = 10^{-1}$ mole

$PbO_2 + 2H_2O \longrightarrow Pb(OH)_4$

$Pb(OH)_4 + 2NaOH \longrightarrow Na_2[Pb(OH)_6]$
or $Na_2PbO_2$ mole of NaOH required = $10^{-1} \times 2$

$V \times \frac{1}{10} \times 10^3 = 2 \times 10^{-1}$

$V = \frac{2}{10^{-3}} = 2000$ mili litre = 2 litre

**(2)** **4.** Let the m. mole of $Fe_2(SO_4)_3$ and $FeC_2O_4$ are a and b m. mole respectively
M.e. of $FeC_2O_4$ = m.e. of $KMnO_4$

$b \times 3 = 40 \times \frac{1}{16}$ .............. (1)

Only $FeC_2O_4$ can oxidise
After reduction we will get 2a + b moles of $Fe^{2+}$ ion which will oxidise by $KMnO_4$
M.e. of $Fe^{2+}$ = M.e $KMnO_4$

$(2a + b) \times 1 = 60 \times \frac{1}{16}$ .............. (2)

a : b = 7 : 4

(3) 2. $P_{N_2} = 0.001$ atm, T = 298 K, V = 2.46 cm$^2$
By ideal gas, PV = nRT

$$n_{N_2} = \frac{PV}{RT} = \frac{0.001 \times 2.46 \times 10^{-3}}{0.0821 \times 298} = 1.0 \times 10^{-7}$$

Now molecules of $N_2 = 6.023 \times 10^{23} \times 1 \times 10^{-7}$
$= 6.023 \times 10^{16}$
Now total surface sites available
$= 6.023 \times 10^{14} \times 1000 = 6.023 \times 10^{17}$

$\therefore$ Surface site used to adsorb $N_2 = \frac{20}{100} \times 6.023 \times 10^{17}$
$= 12.04 \times 10^{16}$

$\therefore$ Sites occupied per molecule of $N_2 = \frac{12.04 \times 10^{16}}{6.02 \times 10^{16}} = 2$

(4) 3. $Be_nAl_2Si_6O_{18}$.
The value of n = 3 by charge balancing.
$(Be^{2+})_3 (Al^{3+})_2 (Si_6O_{18})^{12-}$ [2 × n + 2 (+3) + (–12) = 0
$\therefore$ n = 3

(5) 3. Average titre value $= \frac{25.2 + 25.25 + 25.0}{3} = \frac{75.45}{3}$
$= 25.15 = 25.2$ mL
Number of significant figures will be 3.

(6) 5. $3Br_2 + 3Na_2CO_3 \longrightarrow 5NaBr + NaBrO_3 + CO_2$

(7) 6. m moles of $[Cr(H_2O)_5Cl]Cl_2 = 0.01 \times 30 = 0.3$.
$\Rightarrow$ m mole of $Cl^- = 0.3 \times 2 = 0.6$
$\Rightarrow$ m mole of $Ag^+$ = mmoles of $Cl^-$
$\Rightarrow$ $0.1 \times V = 0.6 \Rightarrow V = 6$ mL.

(8) 8. Molarity of given solution

$$= \frac{\left(\frac{w}{w}\%\right) \times \text{density} \times 10}{\text{mol. wt.}} = \frac{29.2 \times 1.25 \times 10}{36.5} = 10$$

During dilution, $M_1V_1 = M_2V_2$
$10 \times V_1 = 0.4 \times 200$
$V_1 = 8$mL

(9) 4. $6.023 \times 10^{23} \times 1.380 \times 10^{-23} = 8.312$
It has four significant figure.

(10) 8. $m = \frac{w_2 \times 1000}{m_2 \times w_1}$

1 mL solvent having mass 0.4 g.
1000 mL solvent having mass 400 g
1000 mL solution contain 3.2 × 80g solute = 256 g

$\therefore \quad m = \frac{256 \times 1000}{80 \times 400} = 8$

## EXERCISE-4

(1) **(A).** Wt. of $6.023 \times 10^{23}$ molecule of CO = 289gm.
So wt. of $2.01 \times 10^{23}$ molecule of CO

$$= \frac{28 \times 2.01 \times 10^{23}}{6.023 \times 10^{23}} = 9.3\text{gm.}$$

(2) $C_6H_8N_2$ (mol. wt. = 72 + 8 + 28 = 108)
Ratio of C, H and N = 72 : 8 : 28 = 9 : 1 : 3.5

(3) **(C).** 70g of $N_2$ = 5 mole of $N_2$
20g of $H_2$ = 20 mole of $H_2$

(4) **(B).** $n = \frac{\text{molecules of urea}}{\text{Avogadro no.}} = \frac{6.02 \times 10^{20}}{6.02 \times 10^{23}}$
$= 10^{-3}$ mole in 100 ml.

$\therefore$ Concentration $= \frac{n}{\text{litre}} = \frac{10.3}{100} \times 1000 = 10^{-2}$ M=0.01 M

(5) **(A).** $Mg_3(PO_4)_2$, mole = 8 oxygen mole.
So, 8 mole of $O^-$ atoms are contained by
= 1 mole of $Mg_3(PO_4)_2$
0.25 moles of $O^-$ atoms are contained by

$$= \frac{1}{8} \times 0.25 = 3.125 \times 10^{-2}$$

(6) **(D).** $2Al(s) + 6HCl(aq) \longrightarrow 2Al^{+3} + 6Cl^- + 3H_2(g)$
6 mole — 3×22.4 lit.
1 mole — $\frac{3}{6} \times 22.4$ lit. = 11.2 lit.

(7) **(C).** HCl being stronger reducing agent reduces $MnO_4^-$ to $Mn^{2+}$ and result of the titration becomes unsatisfactory.

(8) **(C).** Xethyl alcohol $= \frac{5.2}{5.2 + \frac{1000}{18}} = 0.086$

(9) **(D).** 18g $H_2O$ contains 2g H
$\therefore$ 0.72 g $H_2O$ contains 0.08 gH.
44 g $CO_2$ contains 12g C
$\therefore$ 3.08 g $CO_2$ contains 0.84 g C

$\therefore$ $C : H = \frac{0.84}{12} : \frac{0.08}{1} = 0.07 : 0.08 = 7 : 8$

$\therefore$ Empirical formula = $C_7H_8$

(10) **(D).** Moles of $O_2$ = w/32 ; Moles of $N_2$ = 4w/28

$$\frac{n_{O_2}}{n_{N_2}} = \frac{w}{32} \times \frac{28}{4w} = \frac{7}{32}$$

(11) **(D).** $C_xH_y(g) + \left(x + \frac{y}{4}\right) O_2(g) \to xCO_2(g) + \frac{y}{2} H_2O(\ell)$

15ml — $15\left(x + \frac{y}{4}\right)$ ml — 15 x ml

$O_2$ used = 20% of 375 = 75 ml
Inert part of air = 80% of 375 = 300 ml
Total volume of gases=$CO_2$ + Inert part of air = 330ml
Volume of $CO_2$ = 30 ml, Two equations are x = 2, y = 12

**(12)** **(C).** $M_2CO_3 + 2HCl \rightarrow 2MCl + H_2O + CO_2$

No. of moles of $M_2CO_3$ = No. of moles of $CO_2$ evolved

$1/M = 0.01186$ (M = molar mass of $M_2CO_3$)

$$M = \frac{1}{0.01186} = \frac{10^5}{1186} = 84.3$$

**(13)** **(D).** Mass of $^1H$ in body $= 75 \times 10^3 \times \frac{10}{100} g = 7.5 \times 10^3 g$

No. of moles of $^1H$ replaced by $^2H = 7.5 \times 10^3$

So mass increased $= 7.5 \times 10^3$ g = 7.5 kg

**(14)** **(B).** $C_xH_yO_z$ has z oxygen atom

$$C_xH_y + \left(x + \frac{y}{4}\right) O_2 \rightarrow xCO_2 + \frac{y}{2} H_2O$$

O atoms required for combustion $= 2\left(x + \frac{y}{4}\right)$

$$z = \frac{1}{2}\left[2\left(x + \frac{y}{4}\right)\right] \; ; \; z = x + \frac{y}{4}$$

**(15)** **(D).** $n_{Na^+} = \frac{92}{23} = 4$. So molality = 4

**(16)** **(B).** $n_{eq.}$ $KMnO_4$

$= n_{eq.}[FeC_2O_4 + Fe_2(C_2O_4)_3 + FeSO_4]$

or $n \times 5 = 1 \times 3 + 1 \times 6 + 1 \times 1$ $\quad \therefore n = 2$

**(17)** **(D).** % by mole of carbon $= \frac{1 \text{ mol atom}}{5 \text{ mol atom}} \times 100 = 20\%$

**(18)** **(B).** Volume strength = 11.2 × molarity = 11.2

Molarity = 1 M

Strength = 34 g/L

$$\% \text{ w/w} = \frac{34}{1000} \times 100 = 3.4\%$$

**(19)** **(C).** $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

(A) 0.5 mol 2 mol
(LR)
(B) 1 mol 3 mol (completion)
(C) 2 mol 5 mol
(LR)
(D) 1.25 mol 4 mol
(LR)

**(20)** **(D).** $C_xH_y + \left(x + \frac{y}{4}\right) O_2 \rightarrow xCO_2 + \frac{y}{2} H_2O$

$10 \qquad 10\left(x + \frac{y}{4}\right) \qquad 10x$

By given data, $10\left(x + \frac{y}{4}\right) = 55$ .... (1)

$10x = 40$ .... (2)

$\therefore$ $x = 4, y = 6 \Rightarrow C_4H_6$

**(21)** **(C).** Option (C) is according to Gaylussac's law of volume combination.

**(22)** **(A).** 2 × mole of Urea ≡ mole of $NH_3$ ......(1)

Mole of $NH_3$ = mole of HCl ......(2)

$\therefore$ Mole of HCl = 0.02 mole

**(23)** **26.92**

$[ML_6]Cl_3 + 3AgNO_3 \rightarrow 3AgCl$

0.3 g v ml, 0.125 M

$$\frac{0.3}{267.46} \times 3 = 0.125 \times V \times 10^{-3}$$

or, $V = \frac{0.3 \times 3 \times 1000}{267.46 \times 0.125} = 26.92$ ml.

**(24)** **04.96**

$10 = \frac{\text{Mass of Fe (in g)}}{100 \times 100} \times 10^6$ or mass Fe = 1g

$FeSO_4.7H_2O$ (M = 278)

56 g in 1 mole

$1 \text{ g} — \frac{1}{56}$ mole ; $\frac{1}{56} \times 278g = 4.96$ g

**(25)** **14.**

100 gm soln → 63 gm $HNO_3$

$\frac{100}{1.4}$ mL → 1 mole $HNO_3$

$$\text{Molarity} = \frac{1}{\frac{100}{1.4} \times \frac{1}{1000}} = 14 \text{ M}$$

**(26)** **(D).** $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

$Zn + 2NaOH \rightarrow Na_2ZnO_2 + H_2$

The ratio of the volume of $H_2$ is 1 : 1.

## EXERCISE-5

**(1)** **(A).** $\overset{+7}{KMnO_4} \xrightarrow{OH^-} K_2\overset{+6}{MnO_4}$

Change in oxidation number of Mn in basic medium is 1. Hence mole of KI is equal to mole of $KMnO_4$.

**(2)** **(C).** $2Al_2O_3 + 3C \longrightarrow Al + 3CO_2$

Gram equivalent of $Al_2O_3$ ≡ gm. equivalent of C

Equivalent weight of C $= \frac{12}{4} = 3$ $(\overset{0}{C} \rightarrow \overset{+4}{C}O_2)$

No. of gram equivalent of Al $= \frac{270 \times 10^3}{9} = 30 \times 10^3$

No. of gram equivalent of C $= 30 \times 10^3$

No. of gram equivalent of C $= \frac{\text{mass in gram}}{\text{gram equivalent weight}}$

$\Rightarrow 30 \times 10^3 = \frac{\text{mass}}{3} \Rightarrow$ Mass $= 90 \times 10^3$ g $= 90$ kg

**(3)** **(D).** Average isotopic mass of

$$X = \frac{200 \times 90 + 199 \times 8 + 202 \times 2}{90 + 8 + 2}$$

$$= \frac{18000 + 1892 + 404}{100} = \frac{19996}{100} = 199.96 \text{ amu.}$$

**(4)** **(D).** $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

1 mol $\rightarrow 5 \times 22.4$ ltr

22.4 ltr $\rightarrow 5 \times 22.4$ ltr $\therefore$ 1 ltr $\rightarrow$ 5 ltr

**(5)** **(A).** $PbO + 2HCl \longrightarrow PbCl_2 + H_2O$

223 +73 278gm

Given 6.5gm 3.2gm

eq. of PbO $= \frac{6.5}{223} \times 2 = 0.058$ (limiting reagent)

eq. of HCl $= \frac{3.2}{36.5} \times 1 = 0.08767$

Now, eq. of PbO = eq. of $PbCl_2$

$0.058 = n \times z$; $n = 0.058/2 = 0.029$

**(6)** **(B).** $C \rightarrow 38.71$ $\frac{38.71}{12} = 3.226$ 1

$O \rightarrow 51.62$ $\frac{51.62}{16} = 3.226$ 1

$H \rightarrow 9.67$ $\frac{9.67}{1} = 9.67$ 3

Thus empirical formula $= CH_3O$

**(7)** **(B).** $H_2 + \frac{1}{2}O_2 \longrightarrow H_2O$

10/2 = 5 mol 64/32 = 2 mol

Oxygen is the limiting agent. Hence 4 mole of water formed

**(8)** **(B).** Number of atoms

= number of moles $\times N_A \times$ atomicity

$= 0.1 \times 6.02 \times 10^{23} \times 3 = 1.806 \times 10^{23}$ atoms.

**(9)** **(B).** Molarity $= \frac{\text{number of moles of solute}}{\text{volume of solution (in mL)}} \times 1000$

$= \frac{25.3 \times 1000}{106 \times 250} = 0.9457 \approx 0.955$ M

$Na_2CO_3$ in aqueous solution remains dissociated as

$Na_2CO_3 \rightleftharpoons 2Na^+ + CO_3^{2-}$

x 2x x

Since, the molarity of $Na_2CO_3$ is 0.955 M, the molartiy of $CO_3^{2-}$ is also 0.955 M and that of $Na^+$ is $2 \times 0.955 = 1.910$M

**(10)** **(C).** Moles of $CO_2 = \frac{44}{44} = 1$; Moles of $O_3 = \frac{48}{48} = 1$

Moles of $H_2 = \frac{8}{2} = 4$ ; Moles of $SO_2 = \frac{64}{64} = 1$

**(11)** **(C).** Number of moles

$$= \frac{\text{Number of molecules}}{N_A} = \frac{6.02 \times 10^{20}}{6.02 \times 10^{23}} = 10^{-3} \text{mol}$$

Molar conc. $= \frac{n \times 1000}{V_{solution}(\text{mL})} = \frac{10^{-3} \times 1000}{100}$

Molar conc. = 0.01 M

**(12)** **(C).** Ratio or moles (volume)

$\Rightarrow \frac{W}{2} : \frac{W}{32} : \frac{W}{16} \Rightarrow 16 : 1 : 2$

**(13)** **(A).**

| | $H_2$ + | $Cl_2$ | $\rightarrow$ 2HCl |
|---|---|---|---|
| Initial | 22.4 L | 11.2 L | 0 |
| Final | 11.2 L | 0 | 22.4 L = 1 mole |

**(14)** **(A).** 24 g Mg requires 16 g oxygen

$\because$ 0.56 g oxygen requires 0.84 g Mg

$\therefore$ Mg left = 0.16 g

**(15)** **(A).**

| | $H_2$ | : | $O_2$ |
|---|---|---|---|
| Mass | 1 | | 4 |
| Mole | 1/2 | | 4/32 |

Mole ratio $= \frac{1}{2} \times \frac{32}{4} = \frac{32}{8} = 4 : 1$

**(16)** **(B).** $\because$ 1 mole water $= 6.02 \times 10^{23}$ molecules

$\therefore$ 18 mole water $= 18 \times 6.02 \times 10^{23}$ molecules

So, 18 mole water has maximum number of molecules.

**(17)** **(D).** $\because$ Mass of 1 mol ($6.022 \times 10^{23}$ atoms) of carbon = 12g

If Avogadro Number ($N_A$) is changed

than mass of 1 mol ($6.022 \times 10^{20}$ atom) of carbon

$$= \frac{12 \times 6.022 \times 10^{20}}{6.022 \times 10^{23}} = 12 \times 10^{-3} \text{ g}$$

The mass of 1 mol of carbon is changed.

**(18)** **(B).** $MgCO_3\ (s) \rightarrow MgO\ (s) + CO_2\ (g)$

Moles of $MgCO_3 = 20/84 = 0.238$ mol

From above equation

1 mole $MgCO_3$ gives 1 mole MgO

$\therefore$ 0.238 mole $MgCO_3$ will give 0.238 mole MgO

$= 0.238 \times 40$ g $= 9.523$ g MgO

Practical yield of MgO = 8 g MgO

$\therefore$ % purity $= \frac{8}{9.523} \times 100 = 84\%$

**(19)** **(B).** 1.00 m solution means 1 mole solute is present in 1000g water.

$$n_{H_2O} = \frac{1000}{18} = 55.5 \text{ mol } H_2O$$

$$X_{solute} = \frac{n_{solute}}{n_{solute} + n_{H_2O}} = \frac{1}{1+55.5} = 0.0177$$

**(20)** **(A).** 16.9 g $AgNO_3$ is present in 100 mL solution.

$\therefore$ 8.45 g $AgNO_3$ is present in 50 mL solution

5.8 g NaCl is present in 100 mL solution

$\therefore$ 2.9 g NaCl is present in50 mL solution

$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$

$\frac{8.45}{170}$ mol $\quad \frac{2.9}{58.5}$

$= 0.049$mol $\quad = 0.049$mol $\quad 0$

Mass of AgCl precipitated= $0.049 \times 143.5$ g $= 7$g AgCl

**(21)** **(A).** For $XY_2$, $\because$ 0.1 mole $XY_2 \equiv 10$ g

$\therefore$ 1 mole $XY_2 \equiv 100$ g ; $X + 2Y = 100$ .... (1)

For $X_3Y_2$, $\because$ 0.05 mole $X_3Y_2 \equiv 9$ g

$\therefore$ 1 mole $X_3Y_2 \equiv 180$ g

and $3X + 2Y = 180$ ..... (2)

On solving, X = 40 and Y = 30.

**(22)** **(A).** Molarity has volume term in its expression and volume is temperature dependent.

**(23)** **(A).** $\underset{\substack{2.3 \text{ or} \\ (1/20 \text{ mol})}}{HCOOH} \xrightarrow{\text{Conc. } H_2SO_4} \underset{(1/20 \text{ mol})}{CO\ (g)} + H_2O\ (\ell)$

$$\underset{4.5 \text{ g or } (1/20 \text{ mol})}{\begin{matrix} COOH \\ | \\ COOH \end{matrix}} \xrightarrow{\text{Conc. } H_2SO_4} \underset{(1/20 \text{ mol})}{CO\ (g)} + \underset{(1/20 \text{ mol})}{CO_2\ (g)} + H_2O\ (\ell)$$

Gaseous mixture formed is CO and $CO_2$ when it is passed through KOH, only $CO_2$ is absorbed. The remaining gas is CO. So, weight of remaining gaseous product CO is $\frac{2}{20} \times 28 = 2.8$ g.

**(24)** **(C).**

(A) Moles of water $= \frac{0.00224}{22.4} = 10^{-4}$

Molecules of water = mole × $N_A = 10^{-4} N_A$

(B) Molecules of water = mole × $N_A = \frac{0.18}{18} N_A = 10^{-2} N_A$

(C) Mass of water = 18 × 1 = 18 g

Molecules of water = mole × $N_A = \frac{18}{18} N_A$
$= N_A$

(D) Molecules of water = mole × $N_A = 10^{-3} N_A$

**(25)** **(C).** Haber's process $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

20 moles need to be produced

2 moles of $NH_3 \rightarrow$ 3 moles of $H_2$

Hence 20 moles of $NH_3 \rightarrow \frac{3 \times 20}{2} = 30$ moles of $H_2$