Dev Academy
Yamunanagar (Haryana)

Unit - 1

Some Basic Concepts of Chemistry

Class 11

Revision Notes

CHEMISTRY

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Part – 1

1.1 Introduction

Chemistry is the branch of science which deals with the study of composition, structure, and properties of matter.

1.2 Branches of Chemistry

(a) Inorganic chemistry
(b) Organic chemistry
(c) Physical chemistry
(d) Analytical chemistry
(e) Soil and Agricultural chemistry
(f) Industrial chemistry
(g) Biochemistry
(h) Pharmaceutical chemistry
(i) Nuclear chemistry

1.3 Importance of Chemistry

1.3.1 Chemistry and Health Care

- Chemistry helps in fighting incurable disease like cancer and AIDS.
- **Cisplatin** and **taxol** are two important drugs which are used in **cancer therapy**.
- **AZT (azidothymidine)** is used in treating persons suffering from **AIDS**.
- **Antiseptic** such as dettol, bithional, iodoform, hydrogen peroxide, etc. are used to prevent infection of the wound.
- **Tranquilisers** such as diazepam, Estalopam, luminal, vernal, seconal, equanil etc. are used to reduce tension and depression.
- **Antipyretics and Analgesics** such as aspirin and paracetamol are used for lower down body temperature in fever as well as relieving pain.

1.3.2 Chemistry and Agriculture

- **Fertilisers:** Fertility of soil has been increased by the addition of chemical fertilizers such as urea, ammonium sulphate, calcium ammonium nitrate etc.
- **Insecticides and pesticides:** The crops are protected from pests and other disease using insecticides and pesticides like DDT and Atrazine.

1.3.3 Chemistry and Environment

Refrigerants like chlorofluorocarbon (CFCs) which destroy the ozone layer have been replaced by environment-friendly chemicals (i.e., HFOs, HFAs). However, Greenhouse gases like CH_4, CO_2 etc. are still posing a challenge to the chemists.

1.4 Matter

Matter can be defined as, anything that possesses mass, occupies space, and made up of tiny-particles (i.e., atoms and molecules).

1.4.1 Classification of Matter

The matter can be classified in two ways, depending upon its physical and chemical nature.

A. Physical Classification of Matter

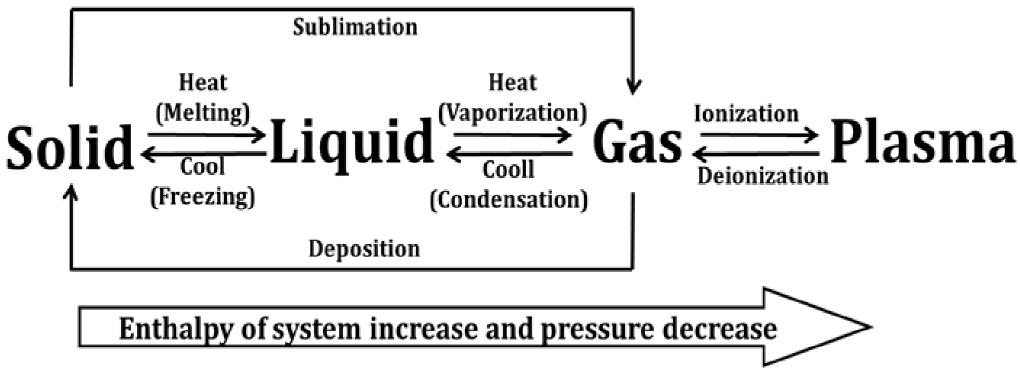
Matter can exist in three physical states i.e., solid, liquid, and gas

(i) **Solid State:** It has a definite shape and a definite volume. e.g., wood, coal etc.

(ii) **Liquid State:** It has definite volume but no definite shape. e.g., water, oil etc.

(iii) **Gaseous State:** It has neither definite volume nor a definite shape. e.g., air, oxygen, ammonia, carbon dioxide etc.

These three states of a matter are interconvertible, by changing the conditions of temperature and pressure.



NOTE 1 : Many other physical states are known to exist only in extreme situations, such as Plasma, Bose–Einstein condensates, neutron-degenerate matter, and quark-gluon plasma, which occur in situations of extreme cold, extreme density, and extremely high-energy etc.

NOTE 2 : Like a gas, plasma does not have definite shape or volume. Unlike gases, plasmas are electrically conductive, produce magnetic fields and electric currents, and respond strongly to electromagnetic forces.

NOTE 3 : Difference Between Gas and Vapour

Vapours represent a gaseous state of a substance which is liquid at room temperature.

Gas represents a gaseous state of a substance which is already gaseous at room temperature.

B. Chemical Classification of Matter

On the basis of chemical composition, matter can be classified as :

(I) Pure substance (II) Mixture

(I) Pure Substance

The substances which are composed of only one kind of particles but if contain more than one kind of particles then they are in fixed ratio by their mass, such substances are considered as pure substances.

They are further of two types: a) Elements b) Compounds

a) Elements: An element is defined as a pure substance that contain only one kind of particles (i.e., may be atoms or homoatomic molecules).

There are 118 elements known to date. They are further of three types:

i. **Metals:** They are those elements which possess shining luster. They are good conductor of heat and electricity. They are malleable and ductile. They possess high density and generally exist in solid state at room temperature (except mercury).

ii. **Non-Metals:** They are bad conductor of heat and electricity (except graphite). They are brittle and possess low density as compared to metals. They exist in all three states at room temperature (e.g. S_8, P_4, I_2 etc. are solids, Br_2 is liquid while O_2, N_2, H_2, Cl_2 etc. are gases).

iii. **Metalloids:** They are those elements which possess the characteristics of both metal as well as non-metals. (e.g. Arsenic, Antimony, Bismuth, Silicon, Germanium, etc.)

b) Compounds: A compound is a pure substance containing two or more than two elements combined in a definite proportion by mass.

They are of two types: **i.** Organic compounds **ii.** Inorganic compounds.

Common Inorganic Compounds and Their Names :

Potash alum = $K_2SO_4.Al_2(SO_4)_3.24H_2O$
Mohr's salt = $FeSO_4.(NH_4)_2SO_4.6H_2O$
Bleaching powder = $CaOCl_2$
Caustic soda = $NaOH$
Baking soda = $NaHCO_3$
Gypsum = $CaSO_4.2H_2O$
Plaster of Paris = $CaSO_4.1/2H_2O$
Blue vitriol = $CuSO_4.5H_2O$
White vitriol = $ZnSO_4$
Green vitriol = $FeSO_4$

(II) Mixture

A substance containing two or more particles (elements or compounds) in any proportion by mass is called mixture.

They are further of two types:

a.) Homogeneous mixture b.) Heterogeneous mixture.

a.) Homogeneous mixture: A mixture is said to be homogeneous if its composition is remained uniform throughout. A homogeneous mixture consists of only one phase. They are generally called solutions.

Example: Air is a homogeneous mixture of a number of gases.

b.) Heterogeneous mixture: A mixture is said to be heterogeneous if its composition is not uniform throughout. The components of a heterogeneous mixture are visible with naked eye. A heterogeneous mixture consists of more than one phase.

Example: Table salt, Milk, Sand in water etc.

Difference between a Compound and a Mixture	
Compound	**Mixture**
1. The components of a compound are always present in a fixed ratio by mass.	1. The components of a mixture may be present in any ratio by mass.
2. Compounds are always homogeneous in nature.	2. Mixture may or may not be homogeneous in nature.
3. Compounds are formed as a result of chemical change.	3. Mixtures are formed as a result of a physical change.
4. They possess sharp melting and boiling points.	4. They do not possess sharp melting and boiling points.

Question. Classify the following as pure substance or mixture. Also separate the pure substances into elements and compounds and divide the mixtures into homogeneous and heterogeneous :

(a) Air (b) Milk (C) Graphite (d) Diamond (e) Gasoline
(f) Tap water (g) Distilled water (h) Oxygen (i) 1-rupee coin
(j) 22 carat gold (k) Steel (l) Iron (m) Sodium chloride
(n) Iodised table salt.

1.4.2 Atoms and Molecules

- **Atoms:**
  An atom is the simplest particle of an element which may or may not be capable of independent existence.
  Example : Atoms of tungsten, Nickel, iron, copper, zinc, silver, gold etc. can exist freely whereas, atoms of hydrogen, chlorine, oxygen, nitrogen, sulphur etc. cannot exist freely. They exist as H_2, Cl_2, O_2, N_2, S_8 etc.

- **Molecules:**
  A molecule is formed when two or more atoms belongs to same or different elements combined together in fixed ratio by their mass. They are capable of independent existence. Molecules may be classified into two categories:
  (a) Molecules of element: They are made up of only one kind of atoms, so that also called Homoatomic or Homonuclear molecules.
  Example: H_2, Cl_2, O_2, N_2, S_8 etc.
  (b) Molecules of compounds: They are made up of atoms of different elements, so that also called Heteroatomic or heteronuclear molecules.
  Example: HF, HCl, H_2O, CO_2, NH_3, CH_4 etc.

- **Atomicity:**
  It represents the number of atoms present in one molecule of a substance.
  Example : H_2O is triatomic, NH_3 is tetratomic, H_2SO_4 is hepta-atomic molecule and so on.

Part – 2

2.1 Atomic and Molecular Masses

2.1.1 Atomic Mass

The atomic mass of an element may be defined as the average relative mass of its atoms as compared with an atom of carbon-12.

It is expressed in atomic mass unit (abbreviated as amu).
1 amu may be defined as a mass exactly equal to 1/12th the mass of one carbon-12 atom. **1 amu = 1.66056 $\times$ 10^{-24} g**

When an element exists in the form of its isotopes, then we use its average relative atomic mass. The average atomic mass may be calculated as

$$\overline{\mathbf{A}} = \sum \mathbf{f_i}\, \mathbf{A_i}$$

Where $\mathbf{A_i}$ is the atomic mass and $\mathbf{f_i}$ is fractional abundances.

[NOTE : In the periodic table, the atomic masses mentioned for different elements represented their average atomic masses.**]**

Difference between the atomic mass and the actual mass of an atom:

The atomic mass of an element is 6.022 $\times$ 10^{23} times the actual mass of its single atom.

The actual mass of 1 atom of an element $= \dfrac{\text{Gram atomic mass of the element}}{6.022 \times 10^{23}}$

Atomic Masses of Some Elements

Elements	Atomic mass (u)	Elements	Atomic mass (u)	Elements	Atomic mass (u)
H	1	Mg	24	Cu	63.5
He	4	Al	27	Zn	65
Li	7	P	31	Br	80
Be	9	S	32	Ag	108
B	11	Cl	35.5	Sn	119
C	12	Ar	40	Sb	122
N	14	K	39	I	127
O	16	Ca	40	Ba	137
F	19	Cr	52	W	184
Ne	20	Mn	55	Au	197
Na	23	Fe	56	Hg	200.5

2.1.2 Molecular Mass

It may be defined as the average relative mass of one molecule of the substance as compared to that of an atom of carbon-12.

Molecular mass of a substance is equal to the sum of the atomic masses (on amu scale) of various elements present in a single molecule of the substance.
Example : Molecular mass of Glucose ($C_6H_{12}O_6$) is :

$$C_6H_{12}O_6 = (6 \times 12 \text{ u}) + (12 \times 1 \text{ u}) + (6 \times 16 \text{ u}) = \mathbf{180\ u}$$

2.1.3 Formula Mass

The formula mass of a substance may be defined as the sum of the atomic masses of all the atoms present in a formula unit of an ionic compound.

Question 1.: Calculate the formula mass of the following ionic compounds :
(i) $K_4[Fe(CN)_6]$ (ii) $Na_2S_2O_3$ (iii) $(NH_4)_2SO_4$ (iv) $K_2[Fe(CO)_2(CN)_4]$

Question 2.: Calculate the molecular mass of the following compounds :
(i) CH_3COOH (ii) CO_2 (iii) $C_6H_5NH_2$ (iv) CCl_4 (v) C_4H_{10} (vi) $C_{12}H_{22}O_{11}$

2.2 Calculate the Percentage Composition of Elements in Compound

The percentage of any element in a compound is the number of parts of that element present in 100 parts by mass of the compound.

$$\text{Percentage of the element} = \frac{\text{No.of parts by mass of the element} \times 100}{\text{Molecular mass of the compound}}$$

Example: Percentage composition of various elements in H_2SO_4.
Molecular mass of H_2SO_4 = 1×2 + 32 + 16× 4 = 2 + 32 + 64 = 98

$$\text{Percentage of H} = \frac{2 \times 100}{98} = 2.041\ \%$$

$$\text{Percentage of S} = \frac{32 \times 100}{98} = 32.653\ \%$$

$$\text{Percentage of O} = \frac{64 \times 100}{98} = 65.306\ \%$$

Question 1.: Calculate the percentage composition of elements in following molecules: (i) $C_6H_{12}O_6$ (ii) H_2O (iii) Al_2O_3 (iv) C_2H_5OH (v) CH_3COOH
Question 2.: Calculate the percentage of water of crystallisation in the sample of: (i) Blue vitriol ($CuSO_4.5H_2O$) (ii) Washing soda ($Na_2CO_3.10H_2O$)

2.3 Empirical Formula and Molecular Formula

Empirical formula represents the simplest whole number ratio of atoms of the various elements present in one molecule of the compound.

Molecular formula represents the actual number of atoms of the various elements present in one molecule of the compound.

Example: The molecular formula of glucose is $C_6H_{12}O_6$ while the empirical formula is $C_1H_2O_1$. Similarly, the molecular formula of benzene is C_6H_6 while the empirical formula is C_1H_1.

Relation Between Empirical and Molecular Formula :

Molecular formula = n × Empirical formula

Where, n is any integer number such as 1, 2, 3....etc.

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

Molecular mass = 2 × Vapour density

$$n = \frac{2 \times \text{Vapour density}}{\text{Empirical formula mass}}$$

2.3.1 Calculation of the Empirical Formula

Calculate the empirical formula of compound whose percentage composition is : Na = 29.11 %, S = 40.51 % and O = 30.38 %

Answer :

Symbol	Percentage (%) of Element	Atomic mass of element	Moles = $\frac{\%}{\text{Atomic mass}}$	Simplest molar ratio	Simplest whole number ratio
Na	29.11	23	$\frac{29.11}{23} = 1.266$	$\frac{1.266}{1.266} = 1$	2
S	40.51	32	$\frac{40.51}{32} = 1.266$	$\frac{1.266}{1.266} = 1$	2
O	30.38	16	$\frac{30.38}{16} = 1.897$	$\frac{1.897}{1.266} = 1.5$	3

Thus, the Empirical formula of the compound is $\mathbf{Na_2S_2O_3}$.

2.3.2 Calculation of the Molecular Formula

An organic acid of molecular mass 90, contain 26.66 % carbon, 2.22 % hydrogen and rest is oxygen. Calculate molecular formula of the acid.

Answer: The percentage of oxygen = 100 – (26.66 + 2.22) = 71.12 %

Symbol	Percentage (%) of Element	Atomic mass of element	Moles = $\frac{\%}{\text{Atomic mass}}$	Simplest molar ratio	Simplest whole number ratio
C	26.66	12	$\frac{26.66}{12} = 2.22$	$\frac{2.22}{2.22} = 1$	1
H	2.22	1	$\frac{2.22}{1} = 2.22$	$\frac{2.22}{2.22} = 1$	1
O	71.12	16	$\frac{71.12}{16} = 4.44$	$\frac{4.44}{2.22} = 2$	2

Hence, the empirical formula of the compound = **CHO_2**

Now, calculations for molecular formula are:

Empirical formula mass = 12 + 1 + 32 = 45

Molecular mass = 90

Value of 'n' = $\frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{90}{45} = 2$

Hence,

molecular formula will be = n × empirical formula = 2 × CHO_2 = **$C_2H_2O_4$.**

Question 1.: A compound contains 4.07 % hydrogen, 24.27 % carbon and 71.65 % chlorine. If the molar mass is 98.96 g. What are its molecular and empirical formula.

Question 2.: A compound contain 10.06 % carbon, 0.84 % hydrogen and 89.10 % Chlorine. Find the molecular formula of its compound if its vapour density is 60.

Question 3. : A compound contain 14.3 % of Na, 9.97 % of S, 6.25 % of H and 69.47 % of Oxygen. If all the hydrogen atoms are present in the form of water of crystallisation, determine the empirical formula of the compound.

Question 4. : An organic compound contains carbon 34.6 % and hydrogen 3.84 %, the rest is oxygen. Its vapour density is 52. Calculate the molecular formula of the compound.

Part – 3

3.1 Properties of Matter and Their Measurement

Every substance has unique properties, which can be classified into two categories: **(a)** Physical properties **(b)** Chemical properties
Let us discuss these properties in brief:

(a) Physical properties

They are those which can be measured or observed without changing the identity or composition of the substance.

Example: Mass, Volume, Density, Melting point, Boiling point etc.

(b) Chemical properties

They are those in which identity or composition of the substance changes due to chemical reaction takes place.

Example: Acidity, Basicity, Combustibility, Precipitation etc.

3.1.1 Fundamental and Derived Units

- The units of measurement like mass, length, time, temperature, electric current, luminous intensity are considered as fundamental units. They are also called basic units. Because, these units cannot be described by other units.
- The units of area, volume, force, work, velocity, density etc. are considered as derived units. This is because they can be derived from fundamental units.

3.1.2 Systems of Measurement of Units

(a) **FPS system (English system):** In this system foot, pound and second were used for the measurement of length, mass, and time respectively.

(b) **MKS system (Metric system):** This system was described by committee of French Academy of Science in 1791. In this case length, mass and time were measured in meter, kilogram and second respectively.

(c) **CGS system:** In this system length, mass and time were measured in centimeter, gram and second respectively.

3.1.3 The International System of Units (SI Units)

The SI system consists of seven basic units by which the other physical quantities such as volume, density, speed etc. can by derived.

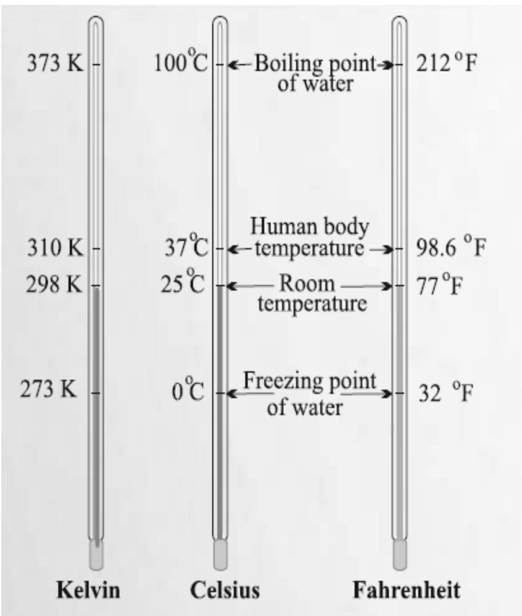
PHYSICAL QUANTITY	SI Unit
Mass	Kilogram (Kg)
Length	Metre (m)
Time	Second (s)
Temperature	Kelvin (K)
Electric current	Ampere (A)
Luminous intensity	Candela (cd)
Amount of the substance	Mole (mol)

3.1.4 Mass and Weight

Mass of the substance is the amount of matter present in it. Weight is the product of mass and force exerted by gravity on the object. The mass of a substance is constant whereas is weight may vary from one place to another due to change in gravity.

Weight = mass × gravity

3.1.5 Measurement of Temperature



There are three scales of temperature i.e., Degree Celsius (˚C), Degree Fahrenheit (˚F) and Kelvin (K).

Kelvin to Degree Celsius :-

$$K = {}^\circ C + 273.15$$

[It is interesting to note that temperature below 0˚ (i.e., negative value) are possible on Celsius scale but on Kelvin scale, negative temperature is not possible.]

Degree Celsius to Degree Fahrenheit :

$${}^\circ C = \frac{5}{9}({}^\circ F - 32) \quad \text{or} \quad {}^\circ F = \frac{9}{5}({}^\circ C) + 32$$

3.1.6 Measurement of Volume

Volume has the units of $(\text{Length})^3$ and can be expressed as :

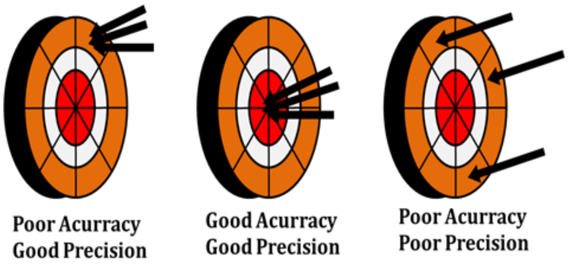
$1\ L = 1\ dm^3$ and $1\ mL = 1\ cm^3$

$1\ L = 1000\ mL = 1000\ cm^3 = 1\ dm^3$

$1\ m^3 = (100\ cm)^3 = 10^6\ cm^3 = 10^3\ dm^3 = 10^3\ L$

3.2 Uncertainty in Measurement and Significant Figures

3.2.1 Precision and Accuracy



- If the average value of different measurements is close to the correct value, the measurement is said to be accurate.
- If the values of different measurements are close to each other and close to their average value, then the measurement is said to be precise.

3.2.2 Significant Figures

The total number of digits in a number including the last digits whose value is uncertain is called the number of significant.

Rules for Determining the Number of Significant Figures :

Rule 1.: All non-zero digits as well as the zeros between the non-zero digits are significant.
Example: 489 has three significant figures, 0.254 has three significant figures
8006 has four significant figures, 5.028 has four significant figures.

Rule 2.: Zero to the left of the first non-zero digit in a number are not significant.
Example: 0.007 has one significant figure, 0.012 has two significant figures

Rule 3.: All trailing zeros or the zeros placed to the right of the decimal point are significant, when number is greater than one. But if the number is less than one, only zeros to the right to the first significant digit are significant.
Example: 23.00 has four significant figures, 0.023 has two significant figures,
0.26800 has five significant figures, 0.030 has two significant figures.

Rule 4.: If a number ends in zero but these zeros are not to the right of a decimal point, these zero may or may not be significant.
Example: 100 has only one significant figure, but 100.0 has four significant figures

Rule 5.: In case of exponential notation, the numerical portion gives the number of significant figures.
Example: 6.023×10^{23} has four significant figures.

3.3 Dimensional Analysis (Unit Factor Method)

Any calculation involving the use of the dimensions of the different physical quantities involved is called dimensional analysis.
It consists of the following steps :

- First determine the unit conversion factor:
  Example: for conversion of pound (lb) into kilogram (Kg)
  1 Kg = 2.205 lb so that $1 = \frac{2.205\ \text{lb}}{1\ \text{Kg}}$ or $1 = \frac{1\ \text{Kg}}{2.205\ \text{lb}}$
- Multiply the given physical quantity with the unit conversion factors, retaining the units of the physical quantity as well as that of the unit conversion factor in such a way that all units cancel out leaving behind only the required units.
- If the conversion involves a number of steps, each conversion factor is used in such a way that the units of the preceding factor cancel out.

Question: How many inches are there in 3.00 km ?
Given that 1 km = 1000 m, 1 m = 1.094 yd, 1yd = 36 in.

Answer: The unit conversion factors will be :

$$1 = \frac{1000\ \text{m}}{1\ \text{km}} = \frac{1\ \text{km}}{1000\ \text{m}} \qquad 1 = \frac{1.094\ \text{yd}}{1\ \text{m}} = \frac{1\ \text{m}}{1.094\ \text{yd}} \qquad 1 = \frac{36\ \text{in}}{1\ \text{yd}} = \frac{1\ \text{yd}}{36\ \text{in}}$$

Hence : $$3.00\ \cancel{\text{km}} \times \frac{1000\ \cancel{\text{m}}}{1\ \cancel{\text{km}}} \times \frac{1.094\ \cancel{\text{yd}}}{1\ \cancel{\text{m}}} \times \frac{36\ \text{in}}{1\ \cancel{\text{yd}}} = \mathbf{1.18 \times 10^5\ in}$$

Part – 4

4.1 Laws of Chemical Combination

4.1.1 Law of Conservation of Mass

This law was proposed by French chemist, Antonie Lavoisier in 1789.
According to this law, in all physical and chemical changes, the total mass of the reactant is equal to that of the product.
So, we can say, Matter can neither be created nor destroyed. Hence, this law is also called Law of Indestructibility of matter.

Question 1.: What mass of silver nitrate will react with 5.85 g of sodium chloride to produce 14.35 g silver chloride and 8.5 g of sodium nitrate, if the law of conservation of mass is true.
Question 2.: If 6.3 g of $NaHCO_3$ are added to 15.0 g of CH_3COOH solution, the residue is found to weight 18.0 g. What mass of CO_2 released in the reaction?

4.1.2 Law of Constant Composition or Law of Definite Proportion

This law was proposed by French chemist J.L. Proust in 1799.
According to this law, A chemical compound is always found to be made up of the same elements combined in the same fixed ratio by mass.
Example: Pure water obtained from any source (i.e., river, well, lake or sea) or any country will always be made up of only hydrogen and oxygen elements combined in the same fixed ratio of 1 : 8 by mass.

Question: 2.16 g of copper metal when treated with nitric acid followed by ignition of the nitrate gave 2.70 g of copper oxide. In another experiment 1.15 g of copper oxide upon reduction with hydrogen gave 0.92 g of copper. Show that the above data illustrate the Law of Constant composition.

4.1.3 Law of Multiple Proportions

This law was proposed by Dalton in 1804.
According to this law, when two elements combine to form two or more chemical compounds, then the masses of one of the elements which combine with a fixed mass of the other, bear a simple ratio to one another.
Example: In CO, the ratio of C:O is 12:16 but in CO_2, the ratio of C:O is 12:32. Thus the weight of oxygen combining with the fixed weight of carbon (12 parts) in two oxides, bear a simple ratio 1:2. This illustrate the law of multiple proportion.

Question: Two different compounds A and B, containing nitrogen and oxygen, were found to have 63.6 % and 46.7 % of nitrogen respectively. Shows that the data illustrates the law of multiple proportion.

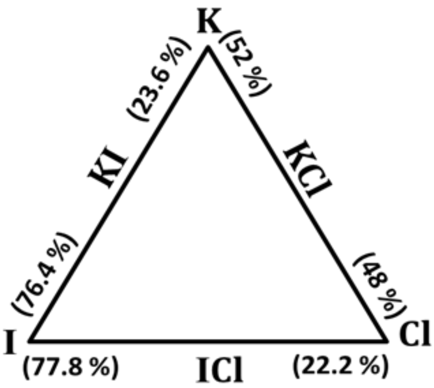
4.1.4 Law of Reciprocal Proportion or Law of Equivalent Proportions

This law was proposed by Richter in 1872.

According to this law, when two elements A and B combine directly with the fixed mass of a third element C, then the ratio in which they do so is either the same or some simple multiple of the ratio in which A and B combine themselves directly.

Question: Potassium chloride contain 52% potassium; Potassium iodide contains 23.6% potassium; and iodine chloride contain 77.8% iodine. Show that the above data illustrate the law of reciprocal proportions.

Explanation:



<u>In case of potassium chloride (KCl)</u>

52 gm K combines with = 48 g chlorine

1 gm K combines with = $\frac{48}{52}$ = 0.923 g chlorine

<u>In case of potassium iodide (KI)</u>

23.6 gm K combines with = 76.4 g iodide

1 gm K combines with = $\frac{76.4}{23.6}$ = 3.237 g iodide

<u>Ratio of I : Cl combining with 1 g potassium is</u>:

3.237 : 0.923 or 3.5 : 1 or 7 : 2

<u>Ratio of I : Cl in which they combine with each other is</u> :

77.8 : 22.2 or 3.5 : 1 or 7 : 2

Thus, the above data illustrate the law of reciprocal proportions.

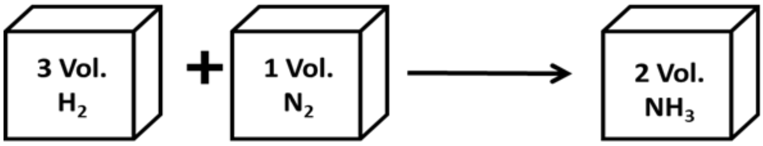
4.1.5 Gay Lussac's Law of Gaseous Volumes

This law was proposed by Gay Lussac.

According to this law, when gases react together, they do so in volumes, which bear a simple ratio to one another and to the volumes of the products, if gases and all the measurements of volumes are done under similar conditions of temperature and pressure.

Illustration : Reaction between nitrogen and hydrogen to produce ammonia.

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



The ratio between the volumes of the reactant and the product in this reaction is simple, 3 : 1 : 2

4.2 Dalton's Atomic Theory

The main points of this theory are as follows :

i. Matter is composed of indivisible particles called atoms.
ii. Atoms of the same elements are identical in all respects. i.e., size, shape, and mass.
iii. Atoms of the different elements have different masses and sizes
iv. Atoms of the different elements have different chemical properties.
v. Atoms of the same or different elements combine to form molecules.
vi. Atoms combine in whole numbers, so a whole number ratio exists between the atoms of different elements in compound
vii. Atom is the smallest unit that take part in a chemical reaction.
viii. Atoms can neither be created or nor destroyed (Atom is indestructible).

Limitations of Dalton's Atomic Theory

i. It failed to explain the law of gaseous volumes
ii. It could not explain why atoms of different elements have different masses and sizes.
iii. It could not explain the nature of binding force between atoms and molecules.
iv. It could not explain why atoms of same or different elements combine to form molecules.

4.3 Modern Atomic Theory

i. Atom is no longer considered to be indivisible. It has complex structure and is made up of lots of small particles specially electrons, protons, and neutrons.
ii. Atoms of the same elements may have different atomic masses. Such as hydrogen has three isotopes having atomic masses 1 amu, 2 amu or 3 amu.
iii. Atoms of different elements may have same atomic masses. Such as calcium and argon have the same atomic masses (i.e., 40 amu)
iv. The ratio in which the different atoms combine with one another may be fixed and integral but may not always be simple. Such as ratio of C : H : O in molecule of sucrose is 12 : 22 : 11 which by no mean is simple.
v. Atom is the smallest particle that takes part in a chemical reaction.
vi. Atom is destructible. By carrying nuclear reactions, atom of the one element may be changed to another.

4.4 Avogadro’s Law / Hypothesis / Principle

Amedo Avogadro, an Italian Physicist stated his hypothesis in 1811.
According to Avogadro’s Law, Equal volume of all gases under similar conditions of temperature and pressure contain equal number of molecules.

Importance or Applications of Avogadro’s Law :

i. To determine the relationship between molecular mass and the vapour density of a gas
ii. To determine the molecular formula of a gas
iii. To determine the relationship between the mass and the volume of a gas.

Part – 5

5.1 Mole Concept

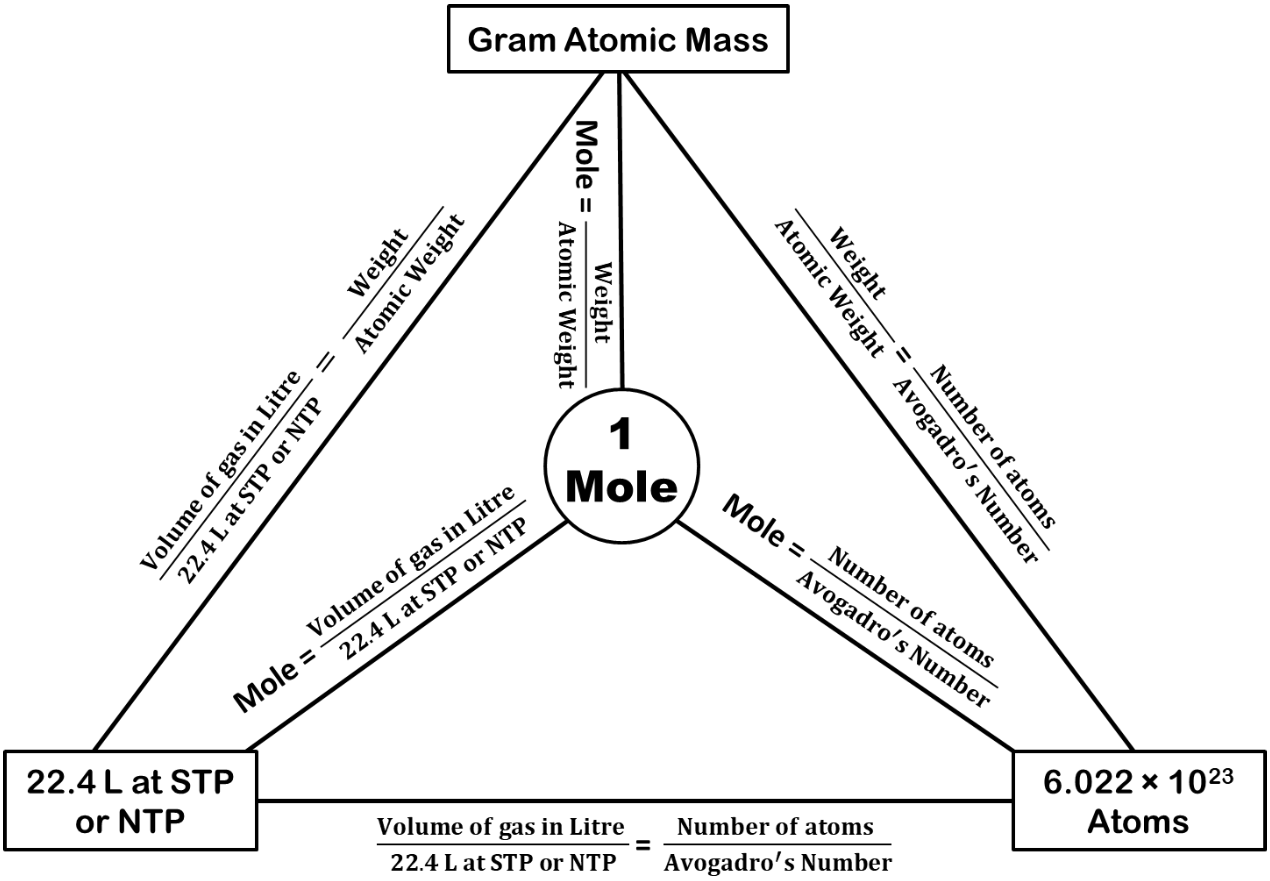
One mole is defined as the amount of substance that contains as many particles as there are in exactly 12 g of the carbon-12 isotope.

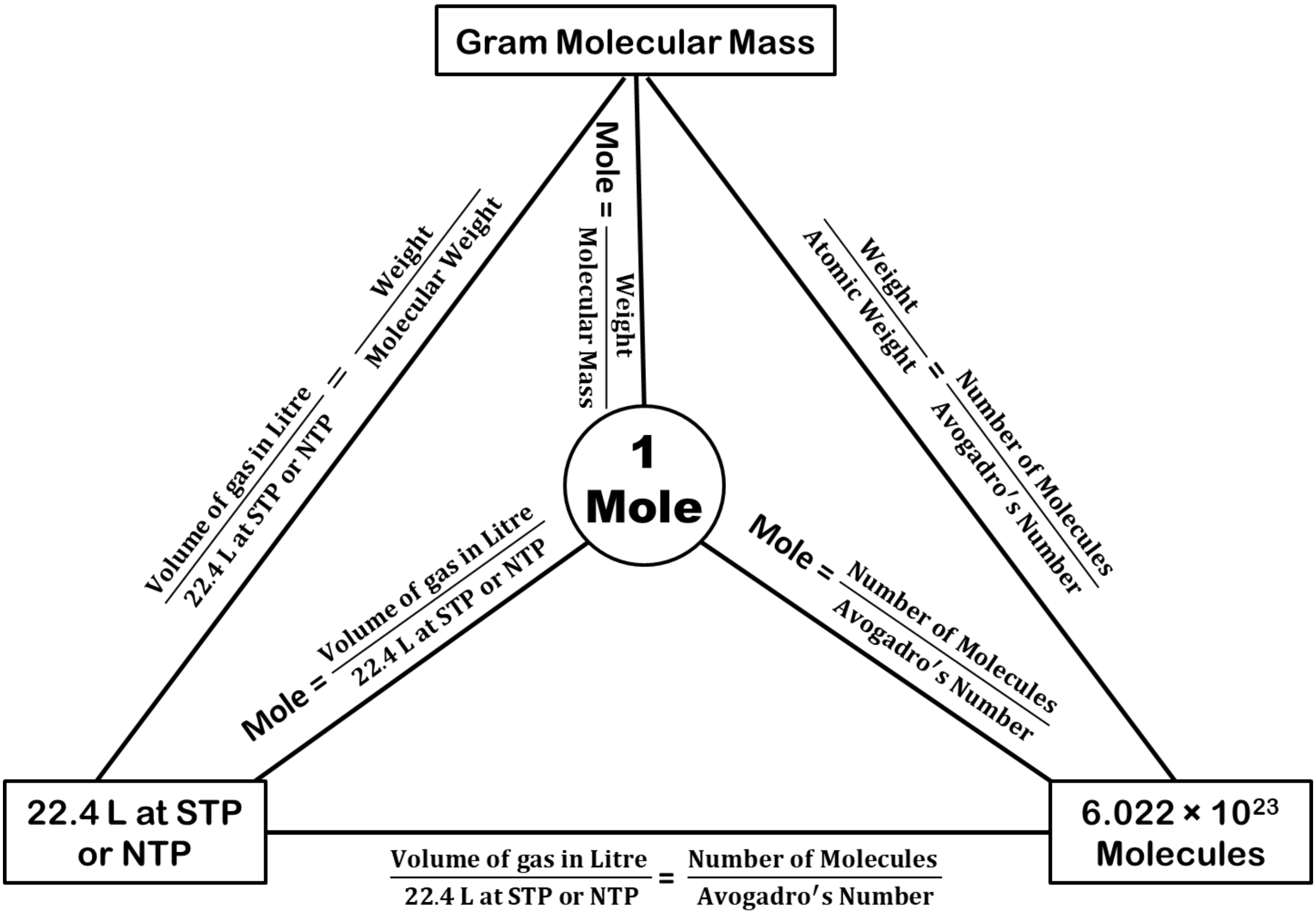
The number of carbon atoms in one mole of carbon is found to be equal to:

1.992648×10^{-23} g is the actual mass of = 1 carbon atom.

12 g is the mass of = $\frac{12\ \text{g/mole}}{1.992648 \times 10^{-23}\ \text{g}}$ = 6.0221367×10^{23} C-atoms/mole

Where, 6.022×10^{23} = Avogadro's Number (N_A)





Question 1.: Calculate the mass of (i) an atom of silver (ii) a molecule of CO_2.
Question 2.: Calculate the number of moles for the following: (i) 4.6 g of sodium (ii) 3.011×10^{24} atoms of Silver (iii) 5.6 litres of CO_2 at S.T.P.
Question 3.: Calculate the number of moles contained in: (i) 115 g of ethyl alcohol (ii) 66.5 g of Magnesium chloride.
Question 4.: Calculate the mass of 0.6 mole of each of the following: (i) $CaCO_3$ (ii) $Na_2SO_4.10H_2O$ (iii) $CaCl_2$ (iv) $MgSO_4$.
Question 5.: Which of the following weighs the most ? (i) 50 g of Fe (ii) 5 g atoms of nitrogen (iii) 0.1 g of Ag (iv) 1×10^{23} atoms of C.
Question 6.: Calculate the number of atoms of the constituent elements in 49 g of H_2SO_4.
Question 7.: Which of the following contains the highest number of atoms? (i) 80 g of Na (ii) 100 g Fe (iii) 5.6 L of oxygen at S.T.P.
Question 8. : Cost of table salt (NaCl) and table sugar ($C_{12}H_{22}O_{11}$) is Rs. 6 per Kg and Rs. 20 per Kg respectively. Calculate their cost per mole.
Question 9.: Chlorophyll, the green matter of plants responsible for photosynthesis, contains 2.68 % of magnesium by mass. Calculate the number of magnesium atoms in 2.00 g of chlorophyll.

5.2 Stoichiometry and Stoichiometric Calculations

Stoichiometry is referring to that branch of chemistry, which deals with the calculations (involving masses, moles, volume, molecules) of the reactants and the products involved in a chemical reaction. The coefficients of the balanced chemical equation are called stoichiometric coefficients.

$CaCO_3$	+	$2HCl \longrightarrow$	$CaCl_2$	+	H_2O	+	CO_2
1 Mole		2 Mole	1 Mole		1 Mole		1 Mole
100 g		2(36.5) = 73 g	111 g		18 g		44 g

Question 1.: How many moles of methane are required to produce 22 g CO_2 after combustion ?
Question 2.: Calculate the mass of iron which will be converted into its oxide.
Question 3.: What mass of $Ca(OH)_2$ or Slaked lime would be required to decompose completely 4 g of ammonium chloride and what would be the mass of each product?

5.3 Limiting Reagent

The reactant which is consumed completely in the reaction is called limiting reagent or limiting reactant. The amount of the product formed then depends upon limiting reagent / reactant.

Question 1.: 50.0 Kg of N_2 and 10.0 Kg of H_2 are mixed to produce NH_3. Calculate the NH_3. Calculate the NH_3 formed. Identify the limiting reagent in the production of NH_3 in this situation.
Question 2.: 3.0 g of H_2 react with 29.0 g of O_2 to form H_2O. Identify the limiting reagent and calculate the maximum amount of H_2O that can be formed. Also determine the amount of the reactant left unreacted.

5.4 Methods for Expressing the Concentration of a Solution

- The concentration of a solution may be defined as the **amount of solute present in the given quantity of the solution or solvent**. The concentration of solution can be expressed in various common ways as discussed below:

(a) Mass %	(b) Volume %	(c) Mass by volume %
(d) parts per million	(e) Strength	(f) Molarity
(g) Molality	(h) Mole fraction	(i) Normality

- Points to Remember:

Conc. = $\frac{\text{Amount of } \textbf{Solute}}{\text{Amount of } \textbf{Solution}}$	Conc. = $\frac{\text{Amount of } \textbf{Solute}}{\text{Amount of } \textbf{Solvent}}$
• Mass Percentage **(w/w)** • Volume Percentage **(v/v)** • Mass-Volume Percentage **(w/v)** • Parts per million **(w/w , v/v)** • Strength **(gm/L)** • Molarity **(mol/L)** • Mole fraction **(mol/mol)** • Normality **(gm eq.wt / L)**	• Molality **(mol/Kg)**

- Symbols Used:

A = Solvent	**B** = Solute	**W** = Weight	**M** = Molar mass	**V** = Volume
n = Moles	**Sol^n** = Solution	**D** = Density	**Eq.wt.** = Equivalent Weight	

- **Mass Percentage (w/w)**

$$\text{Mass \% of B} = \frac{\text{Mass of B} \times 100}{\text{Total mass of solution}} = \frac{W_B \times 100}{W_{sol}} = \frac{W_B \times 100}{W_A + W_B}$$

- **Volume Percentage (v/v)**

$$\text{Volume \% of B} = \frac{\text{Volume of B} \times 100}{\text{Total volume of solution}} = \frac{V_B \times 100}{V_{sol}} = \frac{V_B \times 100}{V_A + V_B}$$

- **Mass-Volume Percentage (w/v)**

$$\text{Mass-Volume \% of B} = \frac{\text{Mass of B} \times 100}{\text{Total volume of solution}} = \frac{W_B \times 100}{V_{sol}}$$

- **Parts per million (ppm)**

(In terms of **w/w**) ppm of B = $\dfrac{\text{Mass of B} \times 10^6}{\text{Total Mass of solution}} = \dfrac{W_B \times 10^6}{W_{sol}}$

(In terms of **v/v**) ppm of B = $\dfrac{\text{Volume of B} \times 10^6}{\text{Total Volume of solution}} = \dfrac{V_B \times 10^6}{V_{sol}}$

- **Strength (gm/L)**

Strength (**S**) = $\dfrac{\text{Mass of the solute in grams}}{\text{Volume of solution in litres}} = \dfrac{W_B}{V_{\text{sol in L}}} = \dfrac{W_B \times 1000}{V_{\text{sol in ml}}}$

- **Molarity (mol/L)**

Molarity (**M**) = $\dfrac{\text{Number of moles of solute}}{\text{Volume of the solution in Litres}} = \dfrac{n_B}{V_{\text{sol in L}}} = \dfrac{n_B \times 1000}{V_{\text{sol in mL}}}$

$$\boxed{\text{Moles of solute } (n_B) = \dfrac{\text{Weight of solute}}{\text{Molecular mass of solute}} = \dfrac{W_B}{M_B}}$$

Molarity (**M**) = $\dfrac{W_B \times 1000}{M_B \times V_{\text{sol in mL}}} = \dfrac{\text{Strength}}{M_B}$

- **Molality (mol/Kg)**

Molality (**m**) = $\dfrac{\text{Number of moles of solute}}{\text{Weight of the solvent in Kg}} = \dfrac{n_B}{W_{\text{A in Kg}}} = \dfrac{n_B \times 1000}{W_{\text{A in g}}}$

Molality (**m**) = $\dfrac{W_B \times 1000}{M_B \times W_{\text{A in g}}}$

- **Mole Fraction (χ : Chi) (mol/mol)**

Mole fraction of B (χ_B) = $\dfrac{\text{Moles of B}}{\text{Moles of solution}} = \dfrac{n_B}{n_{sol}} = \dfrac{n_B}{n_A + n_B}$

Mole fraction of A (χ_A) = $\dfrac{\text{Moles of A}}{\text{Moles of solution}} = \dfrac{n_A}{n_{sol}} = \dfrac{n_A}{n_A + n_B}$

⇨ The sum of mole fraction of all components in solution is always equal to 1.

$$\chi_A + \chi_B = 1$$

⇨ So that: $\chi_A = 1 - \chi_B$ or $\chi_B = 1 - \chi_A$

- **Normality (gm equivalent/L)**

Normality (**N**) = $\frac{\text{Gram equivalent weight of solute}}{\text{Volume of solution in Litre}} = \frac{\text{gm eq.wt}_B}{V_{sol}\text{ in L}} = \frac{\text{gm eq.wt}_B \times 1000}{V_{sol}\text{ in mL}}$

$$\boxed{\text{Gram equivalent weight of solute} = \frac{\text{Weight of solute in gm}}{\text{Equivalent mass of solute}} = \frac{W_B}{\text{Eq.wt.}_B}}$$

Normality (**N**) = $\frac{W_B \times 1000}{\text{Eq.wt.}_B \times V_{\text{Sol in mL}}}$

$$\boxed{\text{Eq.wt. of B} = \frac{M_B}{\text{n\_factor}}}$$

Normality (**N**) = $\frac{W_B \times 1000 \times \text{n\_factor}}{M_B \times V_{\text{Sol in mL}}}$ = Molarity × n_factor

How to determine n_factor of solute ?

n_factor of **acid** = Basicity of the acid (or number of H^+)

n_factor of **base** = Acidity of the base (or number of OH^-)

n_factor of **ion** = Charge on the ion

n_factor of **salt** = Total cationic charge or Total anionic charge per molecule

n_factor of **an element** = Valency of the element

n_factor of **oxidising or reducing agent** = No. of e_s^- gain/loss by a molecule

- **Points to Remember**
  - Molarity, Strength, and Normality change with temperature because of expansion or contraction of the liquid.
  - Molality does not change with temperature because mass of the solvent does not change with change in temperature.
  - Normality equation (Dilution Formula): $N_1V_1 = N_2V_2$
  - Normality of a Mixture: $N_{mixture} = \frac{N_1V_1 + N_2V_2}{V_1 + V_2}$ or $\frac{N_1V_1 + N_2V_2 \ldots + N_nV_n}{V_1 + V_2 \ldots + V_n}$
  - Molarity equation (Dilution Formula): $M_1V_1 = M_2V_2$
  - Molarity equation (for Reacting components): $\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2}$
  - Molarity of a Mixture: $M_{mixture} = \frac{M_1V_1 + M_2V_2}{V_1 + V_2}$ or $\frac{M_1V_1 + M_2V_2 \ldots + M_nV_n}{V_1 + V_2 \ldots + V_n}$
  - Relationship between Molarity and Molality:
    - Molality = $\frac{\text{Molarity} \times 1000}{(\text{density} \times 1000) - (\text{Molarity} \times M_B)}$
    - Molarity = $\frac{\text{Molality} \times 1000 \times \text{density}}{1000 + (\text{Molality} \times M_B)}$

Numerical - 1

2.46 gm of NaOH is dissolved in water and the solution is made to 250 mL in the volumetric flask. Calculate the molarity of the solution.

Explanation:

- Given Values: Solute = NaOH , M_B = 40 , W_B = 2.46 gm , V_{sol} = 250 mL
- Formula Used: $\text{Molarity } (\mathbf{M}) = \dfrac{W_B \times 1000}{M_B \times V_{\text{sol in mL}}}$
- Calculations: $\text{Molarity } (\mathbf{M}) = \dfrac{2.46 \times 1000}{40 \times 250} = \dfrac{2.46}{10} = \mathbf{0.246\ mol/L}$

Numerical - 2

Calculate the molality of a solution containing 20.7 gm of K_2CO_3 dissolved in 500 mL of the solution (assuming density of solution = 1.05 g/ml).

Explanation:

- Given Values: Solute = K_2CO_3 , M_B = 138 , W_B = 20.7 gm , V_{sol} = 500 mL
  D_{sol} = 1.05 gm/mL.
- Formula Used: $\text{Molality } (\mathbf{m}) = \dfrac{W_B \times 1000}{M_B \times W_{\text{A in g}}}$
- Let's find other values:
  $W_{sol} = D_{sol} \times V_{sol} = 1.05 \times 500 = 525$ gm
  $W_A = W_{sol} - W_B = 525 - 20.7 = 504.3$ gm
- Calculations: $\text{Molality } (\mathbf{m}) = \dfrac{20.7 \times 1000}{138 \times 504.3} = \dfrac{20700}{69593.4} = \mathbf{0.297\ mol/Kg}$

Numerical - 3

Battery acid is 4.27 M H_2SO_4 and has density of 1.25 g/ml. What is the molality of H_2SO_4 in the solution?

Explanation:

- Given Values: Solute = H_2SO_4 , M_B = 98 , D_{sol} = 1.25 gm/mL ,
  Molarity = 4.27 mol/L
- Formula Used: $\text{Molality} = \dfrac{\text{Molarity} \times 1000}{(\text{density} \times 1000) - (\text{Molarity} \times M_B)}$
- Calculations: $\text{Molality} = \dfrac{4.27 \times 1000}{(1.25 \times 1000) - (4.27 \times 98)}$ = **5.135 mol/Kg**

Exercises

1. If 16 gm of oxalic acid is dissolved in 750 ml of solution, what is the mass% of oxalic acid in solution? (Density of solution is 1.1 g cm^{-3}).
2. A litre of public supply water contain 3×10^{-3} gm of chlorine. Calculate the mass % and ppm of chlorine.
3. A litre of sea water (about 1030 gm) contains 6×10^{-3} gm of dissolved oxygen. Calculate the mass % and ppm of dissolved oxygen.
4. Calculate the mass percentage of aspirin ($C_9H_8O_4$) in acetonitrile (CH_3CN) when 6.5 g of $C_9H_8O_4$ is dissolved in 450 g of CH_3CN.
5. Calculate the percentage composition in terms of mass of a solution obtained by mixing 300 gm of 25% and 400 gm of 40% solution by mass.
6. Calculate the molarity of a solution of $CaCl_2$ if on chemical analysis it is found that 500 mL of $CaCl_2$ solution contain 1.505×10^{23} Cl^- ions.
7. 100 ml of a solution containing 5 g of NaOH are mixed with 200 ml of M/5 NaOH solution. Calculate the molarity of the resulting solution.
8. The density of a 2.05 M acetic acid in water is 1.02 g/ml. Calculate the molality of the solution.
9. Calculate the mole fraction of ethylene glycol ($C_2H_6O_2$) in a solution containing 20% of $C_2H_6O_2$ by mass.
10. 2.82 gm of glucose is dissolved in 30 gm of water. Calculate (a) molality (b) mole fraction of glucose and water.
11. A solution is 25% water, 25% ethanol, and 50% acetic acid by mass. Calculate the mole fraction of each component.
12. A solution of glucose in water is labelled as 10 % (w/w). The density of the solution is 1.20 g/ml. Calculate the (a) molality (b) molarity and (c) mole fraction of each component in solution.
13. A sugar syrup of weight 214.2 gm contains 34.2 gm of sugar ($C_{12}H_{22}O_{11}$). Calculate: (a) Molal concentration (b) mole fraction of sugar in solution.
14. A commercially available sample of sulphuric acid is 15% H_2SO_4 by weight (density = 1.10 g/mL). Calculate (a) molarity (b) normality (c) molality.
15. Concentrated sulphuric acid has a density 1.9 gm/mL and is 99% H_2SO_4 by mass. Calculate the molarity of sulphuric acid.