9/11/2020
MONDAY

UNIT - 1

SOME BASIC CONCEPTS OF CHEMISTRY

ROLE OF CHEMISTRY

- Supply of Food
- Play a key role in understanding everything from viral structure to Pathogenesis.
- Isolation of vaccine and Therapies.
- Life saving drugs such as cis Platin and Taxol are effective in cancer therapy.
- saving the Environment.
- Application in Industry.

Definition of Chemistry

chemistry is a branch of science which deals with the composition, structure and properties of matter.

MATTER

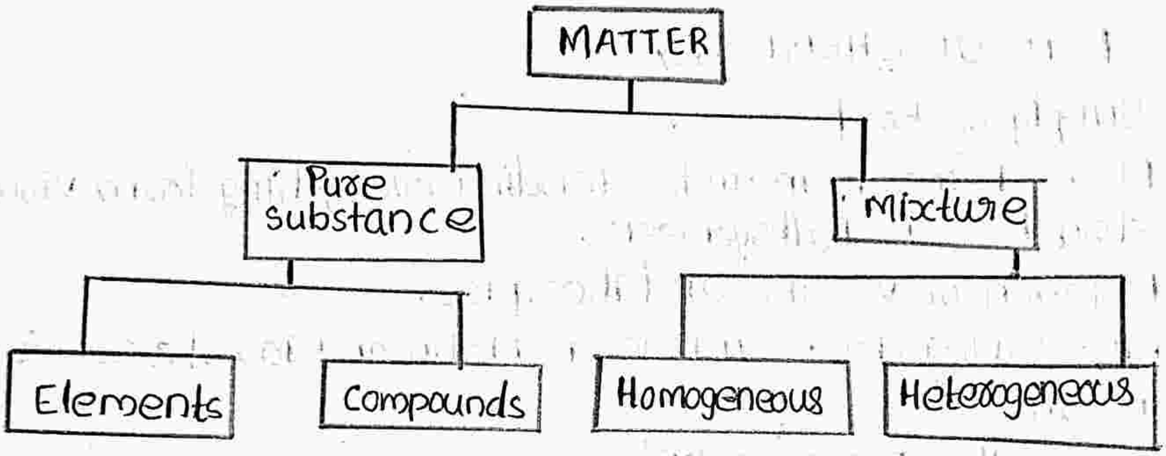
Any thing which has mass and occupies space is called matter.

Three states of Matter :- Solid, Liquid, Gas

SOLIDS	LIQUIDS	GASES
Definite volume, Definite shape, Definite size	Indefinite shape, Take the shape of container, Definite volume	Indefinite shape and volume, Take the shape and volume of container
Particles are closely packed	Particles are not so close, mobile in Nature	Particles are far apart
Particles moves very slowly	Particles moves very slowly	Particles moves very fast

$$\text{SOLID} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{LIQUID} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{GAS}$$

* Classification of Matter



- **Pure substance** :- Materials containing only one substance is called a Pure substance

* **Elements** :- • An element consists of only one type of particle.
  • These particles may be Atoms or Molecules.
  Eg :- Oxygen, Hydrogen, chlorine, etc.

* **Compounds** :- * When two or more Atoms of different elements combine together to form a compound.
  Eg :- water, hydrogen chloride etc.

- **Mixtures** :- Materials containing two or more substance present in it.

* **Homogeneous mixtures** :- Here composition is uniform throughout.
  Eg :- Bronze, Sugar solution

* **Heterogeneous mixtures** :- Here composition is not uniform throughout.
  Eg :- oil in water.

* Properties of Matter And their Measurements

Physical properties	Chemical properties
Properties which can be measured or observed without changing the identity or composition of substance. Eg :- volume, mass, melting point, boiling point etc.	Properties in which a chemical change in the substance occurs Eg :- Burning, Rusting etc.

* Two Different System of Measurements

- English system
- Metric system

Fundamental quantity		S.I. Unit	
Name	Symbol	Name	Symbol
Mass	m	kilogram	kg
Length	l	metre	m
Time	t	second	s
current	I	ampere	A
Temperature	T	kelvin	K
Amount of Substance	n	mole	mol
Luminous intensity	Iv	candela	cd

* Prefix in S.I. system

Factor	Symbol	Prefix
10^{-1}	d	deci
10^{-2}	c	centi
10^{-3}	m	milli
10^{-6}	μ	micro
10^{-9}	n	nano
10^{-12}	p	pico
10^{-15}	f	femto
10^{-18}	a	atto
10^{-21}	z	zepto
10^{-24}	y	yocto

Factor	Symbol	Prefix
10	da	deca
10^{2}	h	hecto
10^{3}	k	kilo
10^{6}	M	mega
10^{9}	G	giga
10^{12}	T	tera
10^{15}	P	peta
10^{18}	E	exa
10^{21}	Z	zetta
10^{24}	Y	yotta

* ## Mass and weight

Mass of a substance is the amount of matter present in it. weight is the force exerted by a gravity on an object.

* ## Volume

Space that a substance or shape occupies or contains.

SI unit of volume = m^3

* ## Density

It is the amount of mass per unit volume. Its unit is kg/m^3

* ## Temperature

It is a physical quantity that express hot and cold.

Three common scales

1. Degree celsius - $0^\circ C$
2. Fahrenheit scale - $0^\circ F = \frac{9}{5}(^\circ C) + 32$
3. Kelvin scale. $K = {}^\circ C + 273.15$

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9/11/2020

Assignment

1) write 5 examples of each on Homogeneous and Heterogeneous mixtures.

Ans)

Homogeneous
- sugar solution
- water and salt
- Bronze
- Soda
- water and yeast
- lemonade
- milk and coffee

Heterogeneous
- Iron ore
- water and oil
- Fried rice
- Granite
- water and sand
- muddy water
- sugar and salt

2) write the definitions of S.I. Basic unit.

Branches of Chemistry

1) organic chemistry - deals with the study of hydrocarbons.
2) Inorganic chemistry - deals with the synthesis and behaviour of inorganic and organometallic compound.
3) Physical chemistry - Theoretical explanation of physical and chemical properties.
4) Analytical chemistry - Studies and uses instruments and methods used to separate, identify and quantify matter.
5) Industrial chemistry (Applied chemistry) - Production and application of industrial products.
6) Nuclear chemistry - study of nuclear reactions.
7) Biochemistry - is both life science and a chemical science. It explores the chemistry of living organisms.
8) Environmental chemistry - is the scientific study of the chemical and biochemical phenomena that occur in natural places.
9) Industrial chemistry - Manufacturing art concerned with the transformation of matter in to useful materials in useful amounts.
10) Polymer chemistry - chemical synthesis, structure and chemical and physical properties of solids.

* ## Uncertainity in measurement

- **Precision** – Refers to the closeness of a set of values obtained from identical measurements of quantity
- **Accuracy** – Reflects how close a measurement is to a known or accepted value.

eg:- If you weigh a given substance 5 times, and get 3.2 kg each time then your measurement is very precise. Precision is independent of accuracy.

Accuracy is close to the true value. If we measure 0.500 g weighed object, then the value got be 0.501 g of 0.500 g, the measurement can be considered as both accurate and precise.

* ## Scientific notation

Scientific notation is a way of expressing numbers that are too large or too small to be conveniently written in decimal form.

In general scientific notation a number is expressed in the form of $N \times 10^n$ where N is called digit term between 1.000 and 9.999 and 'n' is a number called exponent.

Scientific notation is mathematically equal to original number.

eg:- $5325.76 = 5.32576 \times 10^3$

$0.000532576 = 5.32576 \times 10^{-4}$

$10^0 = 1$	$10^{-1} = 0.1 \left(\frac{1}{10}\right)$
$10^1 = 10$	$10^{-2} = 0.01 \left(\frac{1}{100}\right)$
$10^2 = 100$	$10^{-3} = 0.001 \left(\frac{1}{1000}\right)$
$10^3 = 1000$	
$10^4 = 10,000$	$10^{-4} = 0.0001 \left(\frac{1}{10000}\right)$

* ## Significant Figures

• Rules for determining significant figures

1) All digits are significant except zero at the beginning of the numbers.

2) The zeros to the right of the decimal points are significant.
eg:- 200.0 has 4 significant figures.

3) The zeroes to the left of the first non zero digit in a number are not significant.
eg:- 0.0012 has 2 significant figures.

4) All zeroes between two non zero digits are significant.
eg:- 3.005 has 4 significant figures.

5) All zeroes at the end or right of a number are ~~signifa~~ significant provided they are on the right side of the decimal point.
eg:- 1) 0.00430 has three significant figures.
2) 3000 - 1 significant figure (no decimal point)
3) 3.0×10^3 - 2 significant figure

6) Number written in scientific notations are significant.
eg:- 3.02×10^4 = 3 significant figures

* ## LAWS OF CHEMICAL COMBINATIONS

The combination of elements to form compound is governed by the following laws, known as Laws of chemical combinations.

1) The law of conservation of mass.

2) The law of definite proportions.

3) The law of multiple proportions.

4) Gay-Lussac's law of gaseous volumes.

5) Avogadro's Law

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1) Law of conservation of mass. (Antoine Lavoisier)

It states that matter can neither be created nor be destroyed.

or

In a chemical reaction total mass of the reactant is equal to the total mass of products.

eg:- $C_{(g)} + O_{2(g)} \longrightarrow CO_{2(g)}$

12g 32g 44g

$12 + 32 = 44$

In the above case 12g 'C' combines with 32 g of oxygen to form 44 g of CO_2. That is the weight of the reactant (44g) is equal to the weight of products. (44 g).

2) Law of Definite Proportion (Joseph Proust)

A given compound always contains exactly the same proportion of elements by weight. (Irrespection of the source). eg:- CO_2 can be prepared by different methods. Irrespective of the method of preparation. carbon & oxygen in carbon dioxide in the ratio (12 : 32) 3 : 8 by weight.

3) Law of Multiple Proportion (John Dalton)

If two elements can combine to form more than one compound, the masses of one element that combines with a fixed mass of the other element are in the ratio of small whole numbers.

Sl. No	Compound	Combining elements		Ratio of masses	
1	CO_2	C	O	12	32
2	CO	C	O	12	16

The weights of oxygen that combine with the same weight of carbon (12 g) are in the simple ratio. $16 : 32 = 1 : 2$

4) <u>Gay - Lussac's law of Gaseous volumes.</u>

When gases combine to form gaseous product a simple ratio exists between the volume of the reactants and the products at constant temperature and pressure.

eg :- $H_{2(g)} + Cl_{2(g)} \longrightarrow 2HCl_{(g)}$

The ratio of the volumes of $H_2 : Cl_2 : HCl = 1 : 1 : 2$ which is a simple ratio.

5) <u>Avogadro's hypothesis</u>

It states that 'under similar conditions of temperature and pressure equal volumes of all gases contains the same number of molecules.

Avogadro's hypothesis is helpful in.

1. deriving molecular formula of gases.
2. deriving the relation, molecular mass of a gas = 2 × vapour density.
3. determining the atomicity of gases.

Sample 1 $\boxed{1L\ O_2}$ Sample 2 $\boxed{1L\ H_2}$ Sample 3 $\boxed{1L\ CO_2}$ } All contains same no. of molecules at constant Temperature & Pressure

Equal volumes of all gases under the similar conditions of temperature and pressure contain equal no. of molecules.

or

$V \propto n$ at constant T, P

★ ## Dalton's Atomic Theory

Important prostulates are -

1. Matter is made up of extremely small, indivisible particles called atoms.
2. Atoms of the same elements are identical in mass and other properties.
3. Atoms of different elements have different properties.
4. Atoms of same or different elements combine in a simple whole number ratio to form 'molecules'.
5. Atoms can neither be created nor be destroyed.

Assignment of victers class 2

1) Express the following in the scientific notation.

(i) 0.0048

Ans) 4.8×10^{-3}

(ii) 234,000

Ans) 2.34×10^{5}

(iii) 8008

Ans) 8.008×10^{3}

(iv) 500.0

Ans) 5.000×10^{2}

(v) 6.0012

Ans) 6.0012×10^{0}

2) How many significant figures are present in the following?

(i) 0.0025

Ans) 2

(ii) 208

Ans) 3

(iii) 5005

Ans) 4

(iv) 12600

Ans) 3

(v) 500.0

Ans) 4

(vi) 2.0034

Ans) 5

★ **Modifications of Dalton's Atomic Theory**

1. Atom is divisible.
2. All atoms of an element are not identical in mass.
   concept of isotope, where atomic number is same mass number is different.
   eg :- Isotope of H, $^{1}_{1}H$, $^{2}_{1}H$, $^{3}_{1}H$
3. Atoms of different elements may possess same Atomic mass.
   (concept of Isobar where atomic number is different, mass number is same.)
   eg :- $^{14}_{6}C$, $^{14}_{7}N$. (In both case, mass number, Atomic mass is same, but they are different elements)

* **Atomic and Molecular mass**

Atom - smallest particle of an element that retain all its properties and enters into a chemical reaction.

Molecule - smallest particle of matter that has independent existance and retains all the properties of substance.

Atomic mass unit (amu)

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

$$= \frac{1}{12} \times \frac{12}{6.022 \times 10^{23}}$$

$$= \frac{1.9924 \times 10^{-23}}{12}$$

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

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

Atomic mass unit is also expressed as u (unified mass)

1 amu or u = Mass exactly equal to $\frac{1}{12}$th the weight of a C-12 isotope.

Note :- [Atomic masses are expressed relative to that of a standard reference atom. C-12 is chosen as standard because in such case the masses of most of the elements are whole numbers or nearest to the whole number.]

Molecular mass

Molecular mass is a number which express how many times the mass of a molecule is heavier than $\frac{1}{12}$th the mass of a C - 12 atom.

Molecular mass is the sum of atomic masses of the element present in a molecule.

- Calculate the molecular mass of .

a) H_2O

b) CO_2

c) $C_6H_{12}O_6$

Ans) a) Molecular mass of H_2O = (2 x atomic mass of H) + (1 x atomic mass of O).

$= 2 \times 1 + 1 \times 16 = \underline{\underline{18\,u}}$

b) Molecular mass of CO_2 = (1 x atomic mass of C) + (2 x atomic mass of O)

$= 1 \times 12 + 2 \times 16 = \underline{\underline{44\,u}}$

c) Molecular mass of $C_6H_{12}O_6$

= 6 x atomic mass of C + 12 x atomic mass of H + 6 x atomic mass of O

$= 6 \times 12 + 12 \times 1 + 6 \times 16 = \underline{\underline{180\,u}}$

Homework

1) C_2H_5-OH (ethanol)

2) $C_{12}H_{22}O_{11}$ (sucrose)

Ans) 1) Molecular mass of ethanol $= (2 \times 12) + (5 \times 1) + (1 \times 16) + (1 \times 1)$

$= 24 + 5 + 16 + 1$

$= \underline{\underline{46\,u}}$

2) Molecular mass of ~~C_6H~~ $C_{12}H_{22}O_{11}$

$$= (12 \times 12) + (1 \times 22) + (16 \times 11)$$
$$= \underline{\underline{342\ u}}$$

★ GRAM ATOMIC MASS and GRAM MOLECULAR MASS

GAM & GMM

GAM - Atomic mass expressed in grams.

eg :- one gram atom of H = 1.008 g

one gram atom of O = 16 g

One gram atom of Sulphur = 32 g

one gram atom of Silver = 108 g

GMM - Molecular mass of a substance expressed in grams.

eg :- One gram mole of water = 18 g

One gram mole of CO_2 = 44 g

One gram mole of glucose ($C_6H_{12}O_6$) = 180 g

★ Atomic mass - The relative atomic mass is a number which express how many times the mass of an atom is heavier than $\frac{1}{12}$th the mass of a C-12 atom. (or 1 amu)

eg :- Atomic mass of Oxygen is 16 u. means one oxygen atom is 16 times heavier than $\frac{1}{12}$th the mass of a C-12 atom.

Atomic mass of Silver (Ag) is 108 u. Means one silver atom is 108 times heavier than $\frac{1}{12}$th the mass of a C-12 atom.

Average atomic mass

★ The average atomic mass of an element is the sum of the masses of its isotopes, each multiplied by its natural abundance.

$$\text{Average atomic mass} = \frac{\text{Mass}\ (\%\ \text{abundance}) + \text{Mass}\ (\%\ \text{abundance})}{100}$$

- Calculate the average atomic mass of oxygen if its abundance in nature is 99.76 % ^{16}O, 0.04 % ^{17}O, and 0.20 % ^{18}O.

Ans) $$\text{Average atomic mass} = \frac{16(99.76) + 17(0.04) + 18(0.20)}{100}$$

$$= \underline{16.00\ \text{amu}}$$

★ **Formula mass** – Substances like sodium chloride (NaCl) do not exist as discrete molecules. In such case Formula mass is used.

Formula mass of an ionic compound

= (no. of cations × its atomic mass) + (no. of anions × its atomic mass)

eg :– Formula mass of NaCl

= Atomic mass of sodium + Atomic mass of chlorine

= 23 u + 35.5 u

= $\underline{58.5\ u}$

Home work,

1) Calculate the formula mass of KCl. What do you mean by formula mass?

2) Boron has two isotopes.

B-10 (mass 10.013 amu) 19.8 % abundance

B-11 (mass 11.009 amu) 80.2 % abundance

Calculate the average atomic mass

1. Ans) Formula mass of KCl =

Atomic mass of potassium + Atomic mass of chlorine

$= 39.0983\ u + 35.5\ u$

$= \underline{74.5983\ u}$

Formula mass - The formula mass of a substance is the sum of the average atomic masses of each atom represented in the chemical formula and is expressed in amu.

2. Ans) Average atomic mass of B-10 and B-11

$$= \frac{\text{mass}(\%\ \text{abundance}) + \text{Mass}(\%\ \text{abundance})}{100}$$

$$= \frac{(10.013 \times 19.8) + (11.009 \times 80.2)}{100}$$

$= \underline{10.811792}$ amu

Mole Concept

Mole in chemistry is a standard scientific unit for measuring large quantities of very small entities such as atoms, molecules or other specified particle.

1 mole $= 6.022 \times 10^{23}$, This number is called Avogadro number and is represented as N_A.

Mole is also defined as the amount of any substance while contains Avogadro number of particles. (atoms, ions or molecules).

Molar mass

The mass of one mole of any substance is called molar mass.

eg:- Gram atomic mass (GAM)

Gram molecular mass (GMM)

Gram atomic mass - Atomic mass expressed in gram.

concept → 1 gram atom of all elements contain Avogadro number of atoms.

1 g Hydrogen, 4 g Helium, 12 g carbon, 32 g Sulphur, 108 g Silver.

All contains 6.022×10^{23} atoms.

Gram molecular mass - molecular mass expressed in gram.

2 g Hydrogen, 32 g Oxygen, 44 g carbon dioxide, 180 g glucose ($C_6H_{12}O_6$)

All contains 6.022×10^{23} molecules

Molar volume

The volume occupied by 1 mole of any gas at STP is called molar volume.

The molar volume of any ideal gas is 22.4 L at STP

STP (P = 1 atm, T = 273 K) 22.4L

NTP (22.4 L - V)

Atomic mass of few common Elements

H - 1
He - 4
Li - 7
Be - 9
B - 10
C - 12
N - 14
O - 16
Ne - 20
Na - 23
Mg - 24
Al - 27
P - 31

S - 32
Cl - 35.5
K - 39
Ca - 40
Cr - 52
Mn - 55
Fe - 56
Cu = 63.5
Ag - 108
Hg - 200
Pb - 208

Calculation of mass of an atom

1) Calculate the mass of an atom of 'C'.

Ans) From mole concept it is clear that mass of 6.022×10^{23} C atom = 12 g.

$\therefore$ Mass of 1 atom of C = $\frac{12\,g}{6.022 \times 10^{23}}$

$= \underline{1.992 \times 10^{-23}}\,g$

2) Calculate the mass of a silver atom (Ag).

Ans) Gram atomic mass of silver = 108 g

i.e mass 6.022×10^{23} silver atom = 108 g

$\therefore$ Mass of one atom of silver = $\frac{108}{6.022 \times 10^{23}}$

$= \underline{1.793 \times 10^{-22}}\,g$

3) Calculate the mass of one molecule of water.

Ans) Gram molecular mass of H_2O = 18 g

i.e mass of 6.022×10^{23} water molecule = 18 g

Mass of one molecule of water = $\frac{18}{6.022 \times 10^{23}}$

$= \underline{2.99 \times 10^{-23}}\,g$

Generally,

Mass of one atom of any element = $\frac{GAM}{N_A}$

Mass of one molecule of any substance = $\frac{GMM}{N_A}$

Formulas to calculate number of moles. 'n'

* For atomic substances,

$$n = \frac{\text{Given mass}}{GAM} = \frac{m}{GAM}$$

* For molecules,

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

* For gaseous substances,

$$n = \frac{V \text{ at STP in litres}}{22.4\,L}$$

* If number of particle is given then: $n = \frac{\text{no. of Particles}}{\text{Avogadro number}}$

$$n = \frac{N}{N_A}$$

Problems

1. Calculate number of moles in 392 g of H_2SO_4.

Ans) $n = \frac{m}{GMM} = \frac{392}{98} = \underline{\underline{4 \text{ mole}}}$

2. Calculate number of moles in 44.8 L of CO_2 at STP.

Ans) $n = \frac{V}{22.4\,L} = \frac{44.8\,L}{22.4\,L} = \underline{\underline{2 \text{ mole}}}$

3. Calculate number of moles in 11 g of CO_2.

Ans) $n = \frac{m}{GMM} = \frac{11}{44} = 0.25 \text{ mole}$

4. Calculate the number of moles in 3.01×10^{22} molecules of CO_2.

Ans) $n = \frac{N}{N_A}$ (no. of particle)

$= \frac{3.01 \times 10^{22}}{6.022 \times 10^{23}} = 0.0499$ mol

$= \underline{0.05 \text{ mol}}$

5. Calculate the no. of atoms of the, constituent element in 53 g of Na_2CO_3.

Ans) $n = \frac{53}{106} = 0.5$ mole

number of molecules in 0.5 mole

$= 0.5 \times 6.022 \times 10^{23}$ molecule of Na_2CO_3

$= 3.011 \times 10^{23}$ molecule

1 molecule of Na_2CO_3 contains,

2 Na atom, 1 C atom and 3 oxygen atom

No. of sodium atom in 53 g $Na_2CO_3 = 2 \times 3.011 \times 10^{23}$ Na atom

No. of carbon atom in 53 g $Na_2CO_3 = 3.011 \times 10^{23}$ C atoms

No. of oxygen atoms $= 3 \times 3.011 \times 10^{23}$ O atoms

6. Calculate the number of atoms present in 25 g of Carbon.

Ans) $n = \frac{25}{12} = \underline{2.08 \text{ mol}}$

no. of atoms in 2.08 mol $= 2.08 \times 6.022 \times 10^{23}$

$= \underline{12.52 \times 10^{23}}$

7. Find out the volume of 14 g of Nitrogen at NTP.

Ans) $n = \frac{14}{28} = 0.5$ mole

1 mole N_2 occupy 22.4 L.

0.5 mole N_2 occupy $\underline{11.2\ L}$.

8. The oxygen obtained from 72 kg of water.

Ans) $2H_2O \longrightarrow 2H_2 + O_2$

$36\ g\ H_2O \longrightarrow 32\ g\ O_2$

$36\ kg\ H_2O \longrightarrow 32\ kg\ O_2$

1 kg H_2O produce $= \frac{32}{36}$ kg O_2

72 kg H_2O produces $= \frac{32}{36} \times 72$

$= \underline{64\ kg\ O_2}$

Method - 2

$2H_2O \longrightarrow 2H_2 + O_2$

$36\ g\ H_2O \longrightarrow 32\ g\ O_2$

$36\ kg\ H_2O \longrightarrow 32\ kg\ O_2$

$\downarrow$

$72\ kg\ H_2O \longrightarrow 64\ kg\ O_2$

(exact double of 36 = 72)

9. Number of molecules in 1 L of water is close to.

Ans) 1 L of water means = 1000 g

then 'n' $= \frac{1000}{18}$ mol $= 55.55$ mol

1 mole H_2O contains 6.022×10^{23} molecules

$\therefore$ 55.55 mol of H_2O contains $= 55.55 \times \underline{6.022 \times 10^{23}}$ molecules

10. The mass $112\ cm^3$ of CH_4 gas at STP is:

Ans) At STP $22400\ cm^3$ of $CH_4 = 12 + 4 = 16\ g$

At STP $112\ cm^3$ of $CH_4 = \frac{16}{22400} \times 112$

$= \underline{\underline{0.08}}\ g$

Assignment

1. Calculate the molecular mass of glucose ($C_6H_{12}O_6$) molecule.

Ans) molecular mass $= (6 \times 12) + (1 \times 12) + (6 \times 16)$

$= \underline{\underline{180\ u}}$

2. Calculate the molar mass of the following:

i) H_2O

ii) CO_2

iii) CH_4

Ans) i) $H_2O = 2 + 16 = \underline{\underline{18\ g/mol}}$

ii) molar mass of $CO_2 = 12 + (16 \times 2)$

$= \underline{\underline{44\ g/mol}}$

iii) molar mass of $CH_4 = 12 + (4 \times 1)$

$= \underline{\underline{16\ g/mol}}$

PERCENTAGE COMPOSITION

★ The mass percentage of each constituent element present in a compound is called its percentage composition.

mass % of element 'x' in a compound is

$$= \frac{\text{Mass of element 'x'}}{\text{Total mass of the compound}} \times 100$$

Eg :- mass % of hydrogen in water (H_2O) $= \frac{2 \times 1}{18} \times 100$

$= \underline{11.18\ \%}$

Mass % of oxygen in water (H_2O) $= \frac{16}{18} \times 100$

$= \underline{88.88\ \%}$

Mass % of each element in $CaCO_3$

$CaCO_3$

Molecular mass $= 40 + 12 + (16 \times 3)$

$= \underline{100\,g}$

$$\text{Mass \% of element} = \frac{\text{Mass of element}}{\text{Mass of molecule}} \times 100$$

Mass percentage of calcium (Ca) $= \frac{40\,g}{100\,g} \times 100 = \underline{40\ \%}$

" " carbon (C) $= \frac{12\,g}{100\,g} \times 100 = \underline{12\ \%}$

" " Oxygen (O_3) $= \frac{16 \times 3\,g}{100\,g} \times 100 = \underline{48\ \%}$

★ EMPIRICAL AND MOLECULAR FORMULA

EMPIRICAL FORMULA :-

- The empirical formula represents the smallest ratio of atoms present in a compound.

MOLECULAR FORMULA :-

- The molecular formula gives the total number of atoms of each element present in one molecule of a compound.

The empirical formula is the simplest formula and the molecular formula is the "true" formula

★ Relation between Empirical Formula and Molecular Formula

$$\text{Molecular formula} = \text{Empirical formula} \times n$$

$$\therefore n = \frac{\text{molecular formula}}{\text{empirical formula}}$$