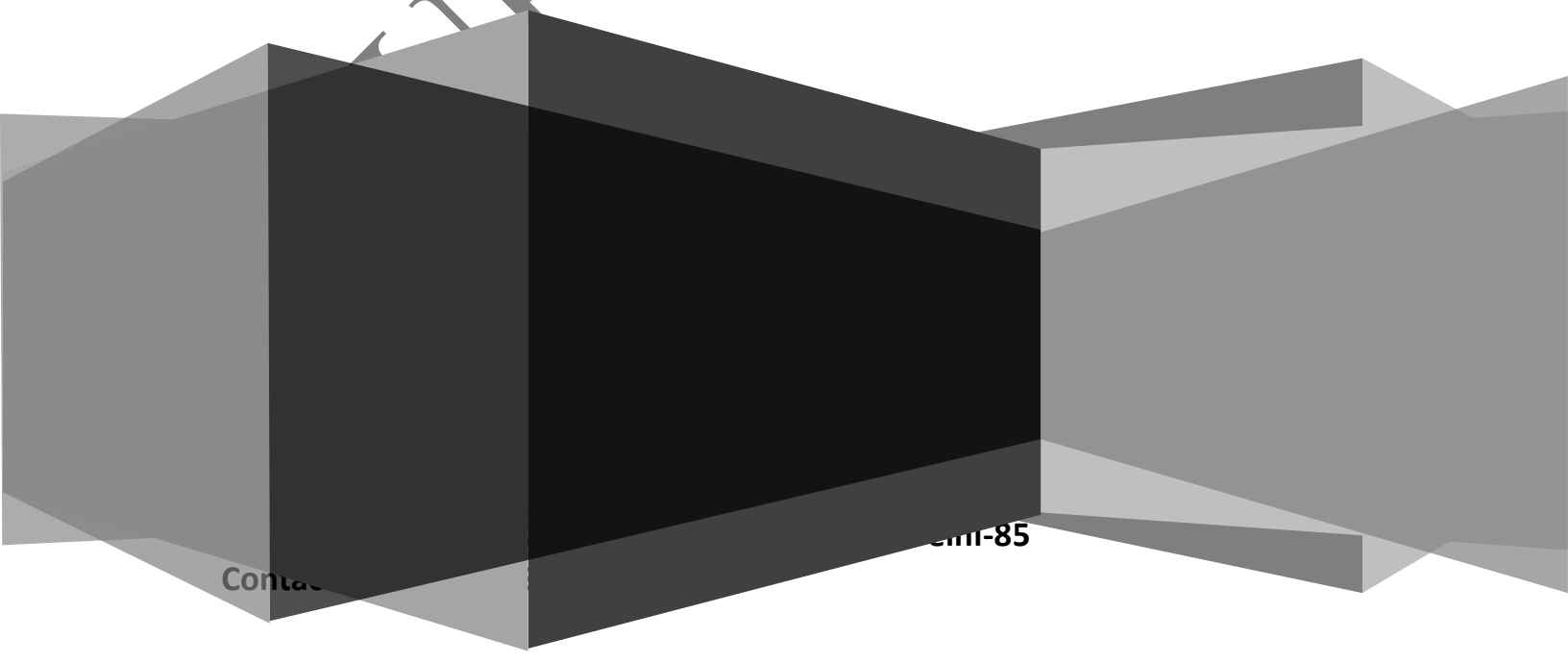
BASIC CONCEPT IN CHEMISTRY

GUPTA CLASSES

For any help contact:
9953168795, 9268789880



BASIC CONCEPTS IN CHEMISTRY

Matter : Anything that occupies space, possess mass and can be felt by any of our senses is known as matter

Physical classification of matter

Solid:- have definite shape and volume, they are rigid eg. Iron, gold wood etc

Liquids:- have definite volume but they take the shape of container in which they are put.

Gases: Neither possesses definite shape nor definite volume. Shape and volume of gas is the shape and volume of the container

Homogeneous:- It has uniform composition and identical properties through out. It has only one phase

Heterogeneous:- It has more than one phase

Element:- Simplest form of matter which can neither be broken into nor built from simpler substances by any chemical or physical methods. It has only one type of atoms. eg. Hydrogen, chlorine, sulphur, ozone, oxygen etc.

Compounds:- It has more than one type of atoms. A compound is obtained by combination of two or more elements in a fixed ratio by mass properties of compound are different from properties of its constituent elements. The constituents of a compounds can not be separated by simple physical methods

Mixture:- A mixture is obtained by mixing two or more substances (elements or compound) in any ratio by mass. Properties of mixture are properties of its components. The constituents of a mixture can be separated by physical methods. A mixture may be heterogeneous or homogeneous

Metals:- Metals have lustre, they are good conductors of heat and electricity they are malleable and ductile, possess high density eg. Copper, silver, gold, aluminium, iron etc. Majority of elements are metals

Non-metals:- Do not possess lustre, they are brittle, poor conductors of heat and electricity and exist in all the three states ie solid, liquid and gas. Carbon, Sulphur, Phosphorus, Iodine are solids Bromine is a liquid and Chlorine, Nitrogen are gases.

Metalloids:- They have properties of both metals and non-metals, Arsenic, Antimony, Silicon.

Atom:- Smallest part of matter (element) which can not be divided into simpler parts. It may or may not have independent existence. It is the smallest part of element which can take part in a chemical reaction

Molecule:- Smallest part of matter (element or compound) which can exist freely.

Allotropy:- An element may exist in two or more different forms which have different physical properties but same chemical properties. Such different forms of same element are known as allotropes and this phenomenon is known as allotropy. eg. Diamond and Graphite are allotropes of carbon.

Polymorphism:- When same compound exists in two or more different crystalline forms then this phenomenon is known as polymorphism. Different forms are known as polymorphs. ZnS has two polymorphs Zinc Blende & Wurtzite

Isomorphism:- When different compounds have similar chemical composition and are in same crystalline form, they are known as isomorphs and this phenomenon as isomorphism eg. $ZnSO_4.7H_2O$ and $FeSO_4.7H_2O$ are isomorphs.

Hygroscopic:- Which absorb moisture from atmosphere eg. P_2O_5, CaO, Anhydrous $CuSO_4$.

Deliquescent :- Solids which absorb large amount of moisture from atmosphere and become wet or pass into solution eg NaOH, $CaCl_2$, $MgCl_2$

Efflorescent:- Solids which when exposed to air lose their water of hydration partly or wholly are known as efflorescent.

(S.I. Units)

Unit:- Standard of reference chosen to measure any physical quantity

Basic Physical quantities and their S.I. units

Units

1.	Length (l)	metre (m)
2.	Mass (m)	kilogram (kg)
3.	Time (t)	second (s)
4.	Electric current (I)	Ampere (A)
5.	Thermodynamic Temperature (T)	Kelvin (K)
6.	Amount of Substance (n)	mole (mole)
7.	Luminous Indensity (Iv)	candela (cd)

Some derived units

1.	Area	m^2
2.	Volume	m^3

3.	Density		Kgm^{-3}
4.	Velocity		ms^{-1}
5.	Acceleration		ms^{-2}
6.	Force		$N=kgms^{-2}$
7.	Pressure		Pascal(Pa) = Nm^{-2} = $kgm^{-1}s^{-2}$
8.	Work, energy		Joule (J) = Nm = kgm^2s^{-2}
9.	Frequency		Hertz (Hz) = s^{-1}
10.	Electric charge		c = As
11.	Potential difference	(volt)	= $JA^{-1}s^{-1}$
			= JC^{-1}
12.	Electric resistance	ohm	= VA^{-1}
13.	Electric conductance		ohm^{-1}(S)

Subsidary units

Deci (d)	10^{-1}	deca (da)	10^{1}
Centi (c)	10^{-2}	Kelo (k)	10^{3}
Milli (m)	10^{-3}	mega (M)	10^{6}
Micro (υ)	10^{-6}	giga (G)	10^{9}
Nano (n)	10^{-9}	Tera (T)	10^{12}
Pico (p)	10^{-12}	peta (P)	10^{15}
Femto (f)	10^{-15}	exa (E)	10^{18}
Atto (a)	10^{-18}	Zeta (Z)	10^{21}
Zepto (z)	10^{-21}	yotta (Y)	10^{24}
Yocto (y)	10^{-24}		

Measurement of volume

1 l = 1000 ml = 1000 cm^3 = 1 dm^3

1 m^3 = 10^3 dm^3 = 10^3 l=10^6 cm^3

Measurement of pressure

1 atm = 1.01325 bar = 1.01325x10^5 Pa

1 bar = 10^5 Pa

1 Atm = 76 cm Hg = 760 mm Hg = 760 torr

Precision and Accuracy

Precision:- Values of different measurements are close to each other hense close to their average value. **Accuracty:-** Average value of different measurements is close to the correct value (Individual measurements may not be close to each other.

Significant Figures:-

Total number of digits in a number including the last digit which is uncertain.

Rules for determining the number of significant figures:-

1. All non-zero digits are significant
2. Zeros between the non-zero digits are significant.
3. Zeros to the left of the first non zero digits are not significant.
4. If a number ends in zeros but these zeros are to the right of decimal point then these zeros are significant.
5. If a number ends in zeros but these zeros are not to the right of decimal point then these zeros may or may not be significant.
6. In case of addition and subtraction the answer should be reported in same number of decimal places as are in the number with least number of decimal places.
7. In case of multiplication or division the result should be reported in the same number of significant figures as possessed by the term with least no. of significant figures

8. If calculation involves a number of steps. The result should contain the same number of significant figures as that of the least precise number involved, other than exact numbers.

$$\frac{42.967 \times 0.02433}{0.34 \times 4} = 0.769298$$

As leaving exact number most precise number has two significant numbers result should be expressed with two significant figures

9. Rounding off.
(i). If the digit next to the last digit to be retained is less than 5 last digit is taken as such and other digits on its right are dropped.
(ii) If the digit is greater than 5. Last digit to be retained is increased by 1
(iii) If the digit is equal to 5 then last significant figure is unchanged if it is even and increased by 1 if is odd.
Scientific notation of a number
Or Exponential notation

$N \times 10^n$

N lies between 1-10 n = exponent
Which may be positive or negative
eg. 1. $4683.507 = 4.683507 \times 10^3$
2. $0.000256 = 2.56 \times 10^{-4}$

Questions

1. What is the difference between 5.0g and 5.00g?
2. How many significant figures are there in each of the following numbers?
   (i) 6.005 (ii) 6.022×10^{23}
   (iii) 8000 (iv) 0.0025
   (v) π (vi) the sum 18.5 + 0.4235
   (vii) the product 14x6.345.
3. Express the following to four significant figures:
   (i) 6.45372 (ii) 48.38250
   (iii) 70000 (iv) 2.65986×10^3
   (v) 0.004687.
4. A sample of nickel weighs 6.5425g and has a density of $8.8 g/cm^3$. What is the volume? Report the answer to correct decimal place.
5. Express the result of the following calculation to the appropriate number of significant figures

$$\frac{3.24 \times 0.08666}{5.006}$$

6. How many significant figures are there in each of the following numbers?
   (i) 6.200 (ii) 0.052
   (iii) 7.5×10^4 (iv) 0.00050
   (v) 67.32-6.3 (vi) 4.2+7.589
   (vii) $(5.56)^2(8.24)/(3.6)$ (viii) 18.567 /(8.1x2)
7. What is the number of significant figures in
   (i) Avogadro's number (6.0×10^{23}) and (ii) Planck's constant (6.62×10^{-34} J s)?
8. Express the number 45000 in exponential notation to show
   (i) two significant figures (ii) four significant figures.
9. A man weighs 175 1b. Express his weight in kg. Given that 1 kg = 2.205 1b.
10. How many inches are there in 3.00 km? Given that 1 km = 1000 m=1.094 yd, 1 yd=36 in.
11. Express the following in S.I. base units using power of 10 notation (Q. No. 2.54 mm = 2.54×10^{-3} m)
    (a) 1.35 mm (b) 1 day (c) 6.45 mL
    (d) 48μg (e) 0.0426 in
12. What is the mass (in grams) of an aluminium block whose dimensions are 2.0 in. x 3.0 in. x4.0 in. and whose density is 2.7 g/cm^3? Given that 1 in. = 2.54 cm.

Nishant Gupta, D-122, Prashant vihar, Rohini, Delhi-85
Contact: 9953168795, 9268789880

13. The mass of precious stones is expressed in terms of 'carat'. Given that 1 carat=3.168 grains and 1 gram = 15.4 grains, calculate the total mass of a ring in grams and kilograms which contains 0.500 carat diamond and 7.00 gram gold.
14. Convert 16.1 km to miles using the following units equivalents:
    1 km = 1000 m, 1 ft = 12 inches
    1 m = 100 cm, 1 mile =1760 yd
    1 inch = 2.54 cm, 1 yd = 3 ft
15. Convert the following into metre
    (i) 40 Em (thickness of Milky way galaxy)
    (ii) 1.4 Gm (diameter of Sun)
16. Vanadium metal is added to steel to impart strength. The density of vanadium is 5.96 g/cm^3. Express this in S.I. units (kg/m^3).
17. A piece of metal is 3 inch (represented by in) long. What is its length in cm?
18. A jug contains 2 L of milk. Calculate the volume of milk in m^3.
19. How many seconds are there in 2 days?

Laws of Chemical Combination

1. Law of Conservation of mass

Proposed by Antoine Lavoisier in 1789 and verified by Landolt

1 "In a chemical change mass is neither created nor destroyed?
2 In a chemical change total mass of the reactants is equal to total mass of products.
3 Total mass of reactants and products at any time of chemical change is constant.

Questions

1. 4.88 g of $KClO_3$ when heated produced 1.92 g of oxygen and the residue (KCl) left behind weighs 2.96g. Show that these results illustrate the law of conservation of mass.
2. When 4.2 g of $NaHCO_3$ is added to a solution of acetic acid (CH_3COOH) weighing 10.0 g, it is observed that 2.2 g of CO_2 is released into the atmosphere. The residue left behind is found to weigh 12.0 g. Show that these observations are in agreement with the law of conservation of mass.

2. Law of constant composition or Definite Proportions

Proposed by J.L. Proust in 1799.

"A chemical compound may be prepared by number of methods but it always contains same elements combined together in the same fixed ratio by mass".

Questions

1. 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 Definite Proportions.
2. 6.488 g of lead combine directly with 1.002 g of oxygen to form lead peroxide (PbO_2). Lead peroxide is also produced by heating lead nitrate and it was found that the percentage of oxygen present in lead peroxide is 13.38 percent. Use these data to illustrate the law of constant composition.

3. Law of Multiple Proportions

Proposed by Dalton in 1808.

" When two elements combine to give two or more compounds then masses of one element which combine with fixed mass of other element are in a simple whole number ratio.

For example:-

Carbon combines with oxygen to give two compounds

	C	O
CO	12 g	16 g
CO_2	12 g	32 g

The mass of oxygen that combine with fixed mass of carbon (12 g) are in the ratio 16:32 ie 1:2 a simple whole number ratio.

Questions

1. Carbon is found to form two oxides, which contain 42.9% and 27.3% of carbon respectively. Show that these figures illustrate the law of multiple proportions.
2. Two oxides of a metal contain 27.6% and 30.0% of oxygen respectively. If the formula of the first oxide is M_3O_4, find that of the second.
3. Copper gives two oxides, On heating 1.0 g of each in hydrogen gas, 0.888 g and 0.799 g of the metal are produced. Show that the results agree with the Law of Multiple Proportions.
4. A metal forms two oxides. One contains 46.67% of the metal and another, 63.94% of the metal. Show that these results are in accordance with the law of multiple proportions.
5. Elements X and Y form two different compounds. In the first, 0.324 g of X is combined with 0.471 g of Y. In the second, 0.117 g of X is combined with 0.509g of Y. Show that these date illustrate the Law of Multiple Proportions.

Law of Reciprocal proportions:-

Law was proposed by Richter in 1792

"When two element 'A' and 'B' combine with a third element 'C' separately then the ratio between the masses of 'A' & 'B' which combine with fixed mass of 'C' is either same or simple multiple of the ratio of masses in which 'A' and 'B' combine directly with each other".

eg. Element 'C' and 'O' combine separately with third element 'H' to give CO_2 and methane.

Ratio between Masses of H and Oxygen that combine with 12 g of carbon separately is 4:32 ie 1:8

Ratio in which Hydrogen combines with oxygen directly is 2:16 ie 1:8 as both ratios are same so it illustrates law of reciprocal proportions.

Questions

1. Ammonia contains 82.35% of nitrogen and 17.65% of hydrogen. Water contains 88.90% of oxygen and 11.10% of hydrogen. Nitrogen trioxide contains 63.15% of oxygen and 36.85% of nitrogen. Show that these data illustrate the law of reciprocal proportions.
2. Carbon dioxide contains 27.27% of carbon, carbon disulphide contains 15.79% of carbon and sulphur dioxide contains 50% of sulphur. Are these figures in agreement with the law of reciprocal proportions?
3. 61.8 g of A combine with 80 g of B. 30.9 g of A combine with 106.5 g of C. B and C combine to form compound CB_2. Atomic weights of C and B are respectively 35.5 and 6.6. Show that the law of reciprocal proportions is obeyed.

Gay Lussac's Law of Gaseous volumes

"When gases combine, they do so in volumes which bear a simple ratio to one another and to the volumes of products, if gases. Provided all the measuremnets are made under similar conditions of temperature and pressure.

eg. $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

1 vol 1 vol 2 vol

one volume of H_2 combines with one volume of Cl_2 to give two volumes of HCl $\therefore$ ratio is

1:1:2 a simple ratio.

Dalton's Atomic Theory

1. Matter is made up of extremely small indivisible particles called atoms.
2. Atoms of same element are identical in all respects ie. Size, Shape and mass.
3. Atoms of different elements are different in mass and size and have different chemical properties.
4. Atoms of same or different elements combine with one another in simple whole number ratio to give compound atoms (now called molecules)
5. Atom of two elements may combine in different ratios to give more than one compound.
6. Atom is the smallest particle that can take part in a chemical reaction
7. Atom can neither be created nor destroyed.

Modern Atomic Theory

1. Atom is no longer considered to be indivisible. It is made up of number of smaller particles like electron, proton and neutron.
2. Atoms of same element may have different atomic masses ie isotopes

3. Atoms of different elements may have same atomic mass ie isobars.
4. Ratio in which different atoms combine may not always be simple.
5. Atom is smallest part of matter which may take part in a chemical reaction

Avogadrow's Law

'Under similar conditions of temperature and pressure equal volumes of all gases contain equal number of molecules'

Application of Avogadrow's Law

1. **Atomicity of gases:-**

Atomicity:- Number of atoms present in one molecule of the element (gas).

Calculation of atomicity of oxygen:-

2 volumes of hydrogen combine with 1 volume of oxygen to give two volumes of water vapours
At same temp. & pressure

Hydrogen + Oxygen $\rightarrow$ Water vap
2 vol 1 vol 2 vol

Let the number of molecules in one volume of Oxygen is n

Applying Avogadrow's Law

Hydrogen + Oxygen $\rightarrow$ Water vapours
2n Molecules n. Molecules 2 n molecules
2 Molecules 1 Molecules 2 Molecules
1 Molecules ½ molecule 1 Molecule
$\therefore$ One molecule of water contains ½ molecule of oxygen but one molecule of water contains one atom of oxygen
$\therefore$ ½ Molecule of oxygen = 1 atom
1 Molecule of oxygen =2 atom
ie atomicity of oxygen is 2

2. **Relation between Molecular mass and vapour density of a gas ie**

M. M. = 2xV.D.

$$\text{V.D. of a gas} = \frac{\text{Mass of certain volume of a gas}}{\text{Mass of same volume of } H_2 \text{ at same temp and pressure}}$$

Applying Avogadrow's law

$$\text{V.D. of Gas} = \frac{\text{Mass of 'n' Molecules of gas}}{\text{Mass of 'n' Molecules of } H_2}$$

$$= \frac{\text{Mass of one Molecule of gas}}{\text{Mass of one Molecule of } H_2}$$

One molecule of Hydrogen = 2 atoms

$$= \frac{\text{Mass of one Molecule of gas}}{\text{Mass of 2 Atoms of 'H'}}$$

$$\text{M.M. of Gas} = \frac{Mass\ of\ one\ Molecule\ of\ gas}{Mass\ of\ 1\ Atoms\ of\ 'H'}$$

$$\therefore \quad \frac{\text{M.M}}{\text{V.D}} = \frac{\text{Mass of one Molecule of gas}}{\text{Mass of 1 Atom of 'H'}} \times \frac{\text{Mass of 2 atoms of H}}{\text{Mass of 1 molecule of Gas}}$$

$$\therefore \quad \frac{\text{M.M}}{\text{V.D}} = \frac{2}{1}$$

M.M. = 2xV.D.

3. Relation between mass and volume of a gas

Molecular Mass = 2 x vapour density

$$= \frac{2 \times \text{Mass certain volume of a gas at S.T.P.}}{\text{Mass of same volume of } H_2 \text{ at S.T.P.}}$$

$$= \frac{2 \times \text{Mass of 1L of a gas at S.T.P.}}{\text{Mass of 1L of } H_2 \text{ at S.T.P.}}$$

$$= \frac{2 \times \text{Mass of 1L of a gas at S.T.P.}}{0.089}$$

$$= \frac{2}{0.089} \times \text{Mass of 1L of a gas at S.T.P.}$$

= 22.4 x mass of 1L of gas at S.T.P.

Molecular Mass = Mass of 22.4 L of a gas at S.T.P.

Atomic Mass

Average relative mass of one atom of an element as compared with one atom of C^{12} taken as 12.

No. of times one atom of an element is heavier than $1/12^{th}$ atom of C^{12}.

Atomic mass unit (a.m.u.) or u → unified mass = It is equal to the mass of $1/12^{th}$ atom of C^{12}.

eg. Average atomic mass = $\frac{\sum \text{Mass of isotopes} \times \%}{100}$

eg.	Isotope	%
	^{20}Ne	90.51
	^{21}Ne	00.27
	^{22}Ne	9.22

Average atomic mass = $\frac{20 \times 90.51x + 21 \times 0.27 + 22 \times 9.22}{100}$

Gram atomic mass:- Atomic mass expressed in grams is known as gram atomic mass

1 g atomic mass of oxygen = 16 g

1 g atomic mass of nitrogen = 14 g

Number of gram atomic mass = $\frac{\text{Mass}}{\text{Atomic Mass}}$

Molecular Mass :- Average relative mass of one molecule of a substance (element or compound) as compared with one atom of C^{12} taken as 12.

Number of times one molecule of a substance (element or compound) is heavier than $1/12^{th}$ atom of C^{12}.

Gram Molecular Mass:- The molecular mass of a substance expressed in grams is known as its gram molecular mass.

One gram molecular mass of O_2 = 32 g

One gram molecular mass of CO_2 = 44 g

No. of gram molecular mass = $\frac{\text{Mass}}{\text{Molecular Mass}}$

Mole That amount of substance which contains 6.023×10^{23} atoms of substance in atomic state or 6.023×10^{23} molecules of substance in molecular state **or**

Mole is a collection of 6.023×10^{23} particles. Amount of the gas which has a volume of 22.4 litres at S.T.P.

No. of Moles of atoms = $\frac{\text{Mass}}{\text{Atomic Mass}}$

No. of moles = $\frac{\text{Mass}}{\text{Molecular Mass}}$

$$\text{No. of moles} = \frac{\text{No. of Molecules}}{6.023 \times 10^{23}}$$

$$\text{No. of moles} = \frac{\text{volume of Gas S.T.P.}}{22.4}$$

Avogadrow's Number

No. of atoms present in 1g atom of element or number of molecules present in 1 gm molecule of the substance. It is equal to 6.023×10^{23}.

Questions

1. Calculate the mass of (I) an atom of silver (ii) a molecule of carbon dioxide.
2. How many atoms and molecules of sulphur are present in 64.0 g of sulphur (S_8)?
3. Calculate the number of molecules present

(i) in 34.20 grams of cane sugar ($C_{12}H_{22}O_{11}$)

(ii) in one litre of water assuming that the density of water is $1 g/cm^3$.

(iii) In one drop of water having mass 0.05g.

4. Calculate the number of atoms of the constituent elements in 53 g of Na_2CO_3.
5. Calculate the number of moles in each of the following

(i) 392 grams of sulphuric acid (ii)44.8 litres of carbon dioxide at STP

(iii)6.022×10^{23} molecules of oxygen (iv)9.0 grams of aluminium

(v)1 metric ton of iron (1 metric ton = 10^3 kg) (vi)7.9 mg of Ca

(vii)65 μg of carbon.

6. Calculate the mass of (i) 0.1 mole of KNO_3 (ii) 1×10^{23} molecules of methane and (iii) 112 cm^3 of hydrogen at STP
7. Arrange the following in order of their increasing masses in grams?

(i) One atom of silver, (ii) One gram-atom of nitrogen,

(iii) One mole of calcium, (iv) One mole of oxygen molecules

(v) 10^{23} atoms of carbon and (vi) One gram of iron

8. Calculate the volume at STP occupied (I) 14 g of nitrogen, (ii) 1.5 moles of carbon dioxide and (iii) 10^{21} molecules of oxygen.
9. Which of the following weighs most?

(i) 50 g of iron (ii) 5 g atoms of nitrogen

(iii) 0.1 g atom of silver (iv) 1×10^{23} atoms of carbon

10. Calculate the mass of CO_2 which contains the same number of molecules as are contained in 40 g of oxygen.
11. Calculate the mass of Na_2CO_3 which will have the same number of molecules as contained in 12.3 g of $MgSO_4.7H_2O$.
12. What is the mass of carbon present in 0.5 mole of $K_4[Fe(CN)_6]$?
13. Chlorophyll, the green colouring matter of plants responsible for photosynthesis, contains 2.68% of magnesium by weight. Calculate the number of magnesium atoms in 2.0 g of chlorophyll.

Problems involving mole concept in chemistry

Solution:- Homogenous mixture of two or more non reacting substances is known as solution.

Binary Solution:- Solution having two components

Solvent:- Component in larger amount

Methods to express concentration of a solution

1. Mass percent or weight percent (w/w%)

Mass of the solute in grams present in 100 g of solution

$$\text{Mass Percent :-} \quad \frac{\text{Mass of Solute}}{\text{Mass of Solution}} \times 100$$

2. **Strength:-** Amount of solute present in 1L of solution.

$$\text{Strength} = \frac{\text{Amount of Solute in g}}{\text{Volume of Solution in L}}$$

Molarity (M)
Number of moles of solute present in one litre of solution.

$$\text{Molarity} = \frac{\text{No. of Moles of Solute}}{\text{Volume of solution in litres}}$$

Molality (m) no. of moles of solute dissolved in 1 kg (1000g) of solvent

$$m = \frac{\text{No. of Moles of Solute}}{\text{Mass of Solvent in kg}}$$

Normality (N) :- No. of gram equivalents of solute dissolved in 1L of solution.

$$\text{Normality} = \frac{\text{No. of g equivalent of Solute}}{\text{Volume of the solution in 'L'}}$$

$$\text{No. of gram equivalents} = \frac{\text{Mass}}{\text{Equivalent Mass}}$$

Equivalent Mass:- Parts by mass of a substance that displaces or combines directly or indirectly with 8 parts by mass of oxygen or 35.5 parts by mass of chlorine or 1.008 parts by mass of hydrogen.

$$\text{Equivalent Mass of acid} = \frac{\text{Molecular Mass}}{\text{Basicity of Acid}}$$

Basicity of Acids:- No. of H^+ ions given by one molecule of acid in aqueous solution
eg. Basicity of HCl = 1
Basicity of H_2SO_4 = 2

$$\text{Equivalent mass of Base} = \frac{\text{Molar Mass}}{\text{Acidity of base}}$$

Acidity of Base:- No. of OH^- ions given by one molecule of Base in aqueous solution

KOH, NaOH	Acidity = 1
$Ca(OH)_2$	Acidity = 2

Mole fraction

$$\text{Mole fraction of solute} = \frac{n_1}{n_1 + n_2}$$

n_1 = No. of Moles of Solute
n_2= No. of moles of solvent

Normality equation
$N_1 V_1 = N_2 V_2$

No. of g equivalents of solute in a solution
= Normality x volume in litres
Molarity equation
$M_1V_1=M_2V_2$
$X_1M_1V_1=X_2M_2V_2$
No. of moles of solute in a solution = Molarity x volume in litres

Parts per Million (p.p.m.) :- No. of parts by mass of solute in 10^6 parts by mass of solution

Questions
1. What is the mass percent of the solute in the solution obtained by dissolving 5 g of the solute in 50 g of water?

Nishant Gupta, D-122, Prashant vihar, Rohini, Delhi-85
Contact: 9953168795, 9268789880

2. A solution of oxalic acid, $(COOH)_2.2H_2O$ is prepared by dissolving 0.63 g of the acid in 250 cm^3 of the solution. Calculate (a) molarity (b) normality of the solution
3. Commercially available concentrated hydrochloric acid contains 38% HCl by mass
   (a) What is the molarity of this solution? The density is 1.19 g cm^{-3}.
   (b) What volume of concentrated hydrochloric acid is required to make 1.00 L of 0.10 M HCl?
4. A solution is prepared by adding 2 g of a substance A to 18 g of water. Calculate the mass percent of the solute.
5. Calculate the molarity of NaOH in the solution prepared by dissolving its 4 g in enough water to form 250 mL of the solution.
6. Concentrated aqueous sulphuric acid is 98% H_2SO_4 by mass and has a density of 1.84 g cm^{-3}. What volume of the concentrated acid is required to make 5.0 litre of 0.500 M H_2SO_4 solution?
7. A sample of $NaNO_3$ weighing 0.38 g is placed in a 50.0 mL measuring flask. The flask is then filled with water upto the mark on the neck. What is the molarity of the solution.

Percentage composition and Molecular formula

Amount of various constituent elements present per 100 g

% (mass) of an element =

$$\frac{\text{No. of atoms of the element x Atomic Mass}}{\text{Formula Mass}} \text{x} 100$$

Example 1.%Composition of Fe_3O_4

Formula Mass = Fe = 3x56 =168
O = 4x16 =64
232

$$\% \text{ of Fe} = \frac{3x56}{232} \text{x} 100 = \frac{16800}{232} = 72.4\%$$

% of O = 100-72.4 = 27.6%

2. % composition of Na_2CO_3

Formula Mass = Na = 2x23=46
C = 1x12=12
O = 3x16=48
106

$$\% \text{ of Na} = \frac{46}{106} \text{x} 100 = 43.40\%$$

$$\% \text{ of C} = \frac{12}{106} \text{x} 100 = 11.32\%$$

$$\% \text{ of O} = \frac{48}{106} \text{x} 100 = 45.28\%$$

1. Ferric sulphate is a crystalline compound of iron. It is used in water and sewage treatment to help the removal of suspended impurities. Calculate the mass percentage of iron, sulphur and oxygen in the compound.
2. Calculate the mass percent of different elements present in ethyl alcohol (C_2H_5OH).
3. Calculate the percentage of hydration, total oxygen in copper (II) sulphate pentahydrate.

Empirical formula and Molecular formula

Empirical Formula:- It is the simplest formula. It shows the simplest whole no ratio between the no. of atoms of different elements present in the compound.

Example – Empirical formula of hydrogen peroxide is H_1O_1

Empirical formula of Glucose is $C_1H_2O_1$

Molecular Formula

It is the exact formula. It shows the actual number of atoms present in one molecule of compound.

Example 1. Molecular formula of hydrogen peroxide is H_2O_2.

Example 2: Molecular formula of Glucose is $C_6H_{12}O_6$

Relation between empirical formula and Molecular formula

Molecular formula = nxEmpirical formula

$$n = \frac{\text{Molecular Mass}}{\text{Empirical Formula Mass}}$$

Molecular Mass = 2xVapour Density

To find out Empirical Formula

1. Divide the percentage of each element by its atomic mass. This will give the relative number of moles of various elements present in the compound.
2. Divide the quotients obtained in the above step by the smallest of them so as to get a simple ratio of moles of various elements.
3. Multiply the figures, so obtained by a suitable integer if necessary in order to obtain whole number ratio.
4. Finally write down the symbols of the various elements side by side and put the above numbers as the subscripts to the lower right hand corner of each symbol. This will represent the **empirical formula** of the compound.
5. Calculate the empirical formula as described above.

6. Find out the empirical formula mass by adding the atomic masses of all the atoms present in the empirical formula of the compound.
7. Divide the molecular mass (determined experimentally be some suitable method) by the empirical formula mass and find out the value of n.
8. Multiply the empirical formula of the compound with n so as to find out the **molecular formula** of the compound.

Questions

1. A substance on analysis, gave the following percentage composition : Na=43.4%, C=11.3%, O=45.3%. Calculate its empirical formula. [Na=23, C=12, O=16]
2. A compound has the following composition: Mg = 9.76, S=13.01%, O=26.01%, H_2O=51.22%. What is its empirical formula? [Mg = 24, S=32, O=16, H=1]
3. What is the simplest formula of the compound which has the following percentage composition : Carbon 80%, Hydrogen 20%? If the molecular mass is 30, calculate its molecular formula.
4. Butyric acid contains C, H, O elements. A 4.24 mg sample of butyric acid is completely burnt in oxygen. It gives 8.45 mg of carbon dioxide and 3.46 mg of water. What is the mass percentage of each element? Determine the empirical and molecular formula of butyric acid if molecular mass of butyric acid is determined to be 88 u.
5. A compound on analysis gave the following percentage composition : Na=14.31%, S=9.97%, H=6.22%, , O = 69.5%
   Calculate the molecular formula of the compound on the assumption that all the hydrogen in the compound is present in combination with oxygen as water of crystallisation. Molecular mass of the compound is 322. [Na=23, S=32, H=1 and O=16]
6. A chemical compound is found to have the following composition : K=42.5%; Fe=15.2%; C=19.5% and N=22.8%. Calculate the simplest formula of the compound.
7. An organic compound on analysis is found to have the following composition : C = 41.1%; H=5.75%; N=16.08% and rest is oxygen. Find out the empirical formula of the compound.
8. What is the empirical formula of the mineral having the composition CaO = 48%, P_2O_5 = 41.3% ; $CaCl_2$ = 10.7%?
9. An organic compound has the following percentage composition C=12.36%; H=2.13%; Br=85%. Its V.D. is 94. Find its molecular formula.
10. A compound contains 32% carbon, 4% hydrogen and rest oxygen. Its vapour density is 75. Calculate the empirical and molecular formula.
11. An organic compound contains C, H and O elements. 1.8 g of the compound on combustion gave 2.64 g of carbon dioxide and 1.08 g of water. Obtain the empirical formula of the compound.
12. A crystalline salt on being rendered anhydrous loses 46% of its mass. The percentage composition of anhydrous salt is:
    Al = 10.50%, K=15.1%, S=24.8%, O=49.6%.
    Find the simplest formula of the anhydrous and crystalline salt.
13. Calculate the empirical formula of a compound whose percentage composition is :
    C=21.9%, H=4.6%, Br=73.4%
14. An acid of molecular mass 104 contains 34.6% carbon, 3.85% hydrogen and the rest is oxygen. Calculate the molecular formula of the acid.

Chemical Equation

It is symbolic representation of a true chemical change

Skelton Equation

Equation in which no attempt has been made to equalize the number of atoms of various elements on both the sides are called **skeleton equations.**

Essentials of a chemical equation

(a) it should represent a true chemical change, i.e. if a reaction is not possible between certain substances, it cannot be represented by chemical equation.
(b) It should be balanced.
(c) It should be molecular, i.e., all the species should be represented in their molecular form.

Information conveyed by chemical equation

1. Names of reactants and products
   $N_2+3H_2\rightarrow 2NH_3$
   Reactants are nitrogen and hydrgoen and the product is ammonia.
2. Relative number of atoms and molecules of reactants and products
   $N_2+3H_2\rightarrow 2NH_3$
   One molecule of nitrogen combines with 3 molecules of hydrogen to give 2 molecules of ammonia.
3. Relative number of moles of reactants and products
   $N_2 + 3H_2\rightarrow 2NH_3$
   1 mole 3 moles 2 moles
   One mole of nitrogen combines with three moles of hydrogen to given two moles of ammonia
4. Relative masses of Reactants and products
   $N_2 + 3H_2 \rightarrow 2NH_3$
   28 g 6g 34 g
   28 g nitrogen combines with 6 g of hydrog5en to give 34 g of Ammonia

5. If a reactant or product is a gas then chemical equation conveys volume of it at S.T.P.

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

At S.T.P. 22.4 *l* 67.2 *l* 44.8 *l*

At S.T.P. 22.4 L of nitrogen combines with 67.2 L of hydrogen to given 44.8 L of ammonia

The chemical equation can be made more informative by incorporating the following changes:

(a) The physical states of reactants and products can be indicated by using the abbreviations; (s) for solid, (*l*) for liquid, (g) for gas and (aq) for aqueous solution. For example,
$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq)+H_2(g)$

(b) In order to indicate the strength of acid or base dil. for dilute ; conc. for concentrated is written before the formula of acid or base
$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq)+H_2(g)$

(c) the reaction condition such as presence of catalyst ,temperature , pressure etc. may be written above the arrow between the reactant and products
$2SO_2(g) + O_2(g) \xrightarrow{V_2O_5,775K} 2SO_3(g)$

(d) Heat change taking place during the reaction may be expressed in any of the following two ways
$N_2 + 3H_2 \rightarrow 2NH_3(g) + 93.6\ kJ$
$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)\ ;\ \Delta H = -93.6kJ$

Balancing of chemical equation

Exercise:*Balance the following equations by Hit and Trial method.*

1. $Fe_3O_4+H_2\rightarrow Fe+H_2O$
2. $C_2H_2+O_2\rightarrow CO_2+H_2O$
3. $Mg_3N_2+H_2O\rightarrow Mg(OH)_2+NH_3$
4. $SO_2+H_2S\rightarrow S+H_2O$
5. $Al_4C_3+H_2O\rightarrow Al(OH)_3+CH_4$
6. $KMnO_4+HCl\rightarrow KCl+MnCl_2+H_2O+Cl_2$
7. $KMnO_4+KOH\rightarrow K_2MnO_4+MnO_2+Cl_2$
8. $FeS_2+O_2\rightarrow Fe_2O_3+SO_2$
9. $Zn+NaOH\rightarrow Na_2ZnO_2+H_2$
10. $Na_2S_2O_3+I_2\rightarrow Na_2S_4O_6+NaI$
11. $C_2H_6+O_2\rightarrow CO_2+H_2O$
12. $Ca_3P_2+H_2O\rightarrow Ca(OH)_2+PH_3$

Balance the following equations by partial equation method.

13. $NaOH+Cl_2\rightarrow NaCl+NaClO_3+H_2O$
14. $P+HNO_3\rightarrow H_3PO_4+H_2O+NO_2$
15. $KMnO_4+H_2SO_4+(COOH)_2\rightarrow K_2SO_4+MnSO_4+CO_2+H_2O$
16. $PbS+O_3\rightarrow PbSO_4+O_2$
17. $KMnO_4+FeSO_4+H_2SO_4\rightarrow K_2SO_4+MnSO_4+Fe_2(SO_4)_3+H_2O$
18. $Mg+HNO_3\rightarrow Mg(NO_3)_2+NH_4NO_3+H_2O$
19. $Cu+HNO_3\rightarrow Cu(NO_3)_2+NO+H_2O$
20. $I_2+HNO_3\rightarrow HIO_3+NO_2+H_2O$

Problems Based of Stoichiometry

1. Calculate the mass of iron which will be converted into oxide (Fe_3O_4) by the action of 18 g of steam on it.
2. What mass of slaked lime would be required to decompose completely 4 grams of ammonium chloride and what would be the mass of each product?
3. 1.5 g of an impure sample of sodium sulphate dissolved in water was treated with excess of barium chloride solution when 1.74 g of $BaSO_4$ were obtained as dry precipitate. Calculate the percentage purity of the sample.
4. Current market prices of Al, Zn and Fe scraps per kg are Rs. 20, Rs. 16 and Rs.3 respectively. If H_2 is to be prepared by the reaction of one of these metals with H_2SO_4, which would be the cheapest metal to use? Which would be most expensive?
5. In order to find the strength of a sample of sulphuric acid, 10 g were diluted with water and a piece of marble weighting 7 g placed in it. When all action has ceased, the marble was removed, washed, dried and was found to weigh 2.2 g. What was the percentage strength of sulphuric acid?
6. 1.84 g of a mixture of $CaCO_3$ and $MgCO_3$ is strongly heated till no further loss of mass takes place. The residue weighs 0.96 g. Calculate the percentage composition of the mixture.
7. Calculate the amount of lime, $Ca(OH)_2$, required to remove hardness of 50,000 litres of well water which has been found to contain 1.62 g of calcium bicarbonate per 10 litres.
8. What volume of carbon dioxide measured at 27^0C and 746.7 mm pressure would be obtained by treating 10.0 g of pure marble with dilute hydrochloric acid? (Aq. Tension at 27^0C is 26.7 mm).
9. 1.0 g of a mixture of carbonate of calcium and magnesium gave 240 cm^3 of CO_2 at STP. Calculate the percentage composition of the mixture.
10. What volume of oxygen at STP is required to effect complete combustion of 200 cm^3 of acetylene and what would be the volume of carbon dioxide formed?
11. In the commercial manufacture of nitric acid, how many moles of NO_2 produces 7.33 mol of HNO_3 in the reaction:
$3NO_2(g)+H_2O(l) \rightarrow 2HNO_3(aq)+No(g)$?
12. How much of Fe can be theoretically obtained by the reduction of 1 kg of Fe_2O_3?
13. Calculate the mass of 60% H_2SO_4 required to decompose 50 g of chalk (calcium carbonate).

Nishant Gupta, D-122, Prashant vihar, Rohini, Delhi-85
Contact: 9953168795, 9268789880

14. Excess of $AgNO_3$ solution was added to 2.2 g of commercial sample of common salt dissolved in water. The mss of dried precipitate or silver chloride was 2.11 g . Calculate the present purity or common salt.
15. Calculate the mass of graphite that must be burnt to produce 13.2 g CO_2?
16. One gram of a mixture of potassium and sodium chlorides on treatment with excess of silver nitrate gave 2g of AgCl . What was the composition of the two salts in the original mixture?
17. Calculate the amount of water(g) produced by the combustion of 16 g of methane.
18. How many moles of methane are required to produce 22 g CO_2 (g) after combustion?
19. How much marble of 96.5% purity would be required to prepare 10 litres of carbon dioxide at STP when the marble is acted upon by dilute hydrochloric acid?
20. Calculate the volume of SO_2 at STP obtained by burning 500 g of S containing 4% sand by weight.
21. The drain cleaner, Drainex contains small bits of aluminium which reacts with caustic soda to produces dihydrogen. What volume of dihydrogen at 20^0C and one bar will be released when 0.15 g of aluminium reacts?
22. Find out the volume of Cl_2 at STP produced by the action of $100cm^2$ of 0.2 N HCl on excess of MnO_2.
23. 5.0 g of marble was added to 7.5 g dilute hydrochloric acid. After the reaction was over, it was found that 0.5 g of marble was left unused . Calculate the percentage strength of hydrochloric acid. What volume of CO_2 measured at STP will be evolved in the above reaction?
24. Calculate the volume of 1.00 mol L^{-1} aqueous sodium hydroxide that is neutralized by 200 mL of 2.00 mol L^{-1} aqueous hydrochloric acid and mass of sodium chloride produced.
25. 3.0 g of H_2 react with 29.0 g of O_2 to form H_2O.
    (i) Which is the limiting reactant
    (ii) Calculate the maximum amount of H_2O that can be formed
    (iii) Calculate the amount of the reactant left unreached. Molecular mass of H_2=2.016.
26. One litre of oxygen of STP is made to react with three litres of carbon monoxide at STP . Calculate the mass of each substance found after the reaction. Which one is the limiting reactant?
27. 50.0 kg of N_2(g) and 10.0 kg of H_2(g) are mixed to produce NH_3(g). Calculate the NH_3(g) formed. Identify the limiting reagent in the production of NH_3 in this situation.
28. Igniting MnO_2 coverts it quantitatively to Mn_3O_4. A sample of pyrolusite is the following composition: MnO_2=80%, SiO_2 and other inert contents = 15%, rest being water. The sample is ignited in air to constant weight. What is the percentage of Mn in the ignited sample? (Atomic weight of Mn=55)
29. A 2.0 g sample of a mixture containing sodium carbonate, sodium bicarbonate and sodium sulphate is gently heated till the evolution of CO_2 ceases. The volume of CO_2 at 750 mm Hg pressure and at 298 K is measured to be 123.9 ml. A 1.5 g of the same sample requires 150 ml of M/10 HCl for complete neutralization. Calculate the percentage composition of the components of the mixture.
30. A mixture of ethane (C_2H_6) and ethene (C_2H_4) occupies 40 litres at 1.00 atm and at 400 K The mixture reacts completely with 130 g of O_2 to produce CO_2 and H_2O. Assuming ideal gas behaviour, calculate the mole fractions of C_2H_4 andC_2H_6 in the mixtujre.
31. A mixture of 20 mL of CO, CH_4 and N_2 was burnt in excess of O_2 resulting in the reduction of 13 mL of volume. The residual gas was then treated with KOH solution to show a contraction of 14 mL in volume. Calculate volume of CO, CH_4 and N_2 in the mixture. All measurements are made at constant pressure and temperature.
32. 1.2 g mixture of Na_2CO_3 and K_2CO_3 was dissolved in water to form 100 cm^3 of a solution. 20 cm^3 of this solution required 40 cm^3 of 0.1 N HCl for neutralisation. Calculate the weight of Na_2CO_3 and K_2CO_3 in the mixture.
33. The formula weight of an acid is 82.0 in a titration. 100 cm^3 of a solution of this acid containing 39.0 g of the acid per litre were completely neutralised by 95.0 cm^3 of aqueous NaOH containing 40.0 g of NaOH per litre. What is the basicity of the acid?
34. A mixture of FeO and Fe_3O_4 was heated in air to a constant mass. It was found to gain 10% in its mass. Calculate the percentage composition of the original mixture.
35. How many grams of CaO are required to neutralize 852 g of P_4O_{10}.

Nishant Gupta, D-122, Prashant vihar, Rohini, Delhi-85
Contact: 9953168795, 9268789880

MCQ with only one answer

1. 0.50 g of a metal on oxidation gives 0.70 g of its oxide. Eq. wt. of metal is
   (a) 10 (b) 25
   (c) 20 (d) 40
2. Maximum no. of molecules is present in
   (a) 15 *l* of H_2 at STP (b) 5 *l* of N_2 at STP
   (c) 0.5 *g* of H_2 gas (d) 10 *g* of O_2 gas
3. How many moles of $Mg_3(PO_4)_2$ will contain 0.25 mole of oxygen atoms
   (a) 1.25×10^{-2} (b) 2.5×10^{-2}
   (c) 0.02 (d) 3.125×10^{-2}
4. In the reaction
   $2Al_{(s)} + 6HCl_{(aq)} \rightarrow 2Al^{3+}_{(aq)} + 6Cl^{-}_{(aq)} + 3H_{2(g)}$
   (a) 33.6 *l* $H_{2(g)}$ is produced regardless of temp and pressure for every mole of Al
   (b) 67.2 *l* H_2 (g) at STP is produced for every mole of HCl
   (c) 11.2 *l* of H_2 (g) is produced for every mole of HCl (aq)
   (d) 6l $HCl_{(aq)}$ is consumed for every 3*l* $H_{2(g)}$ is produced
5. V. D. of a metal chloride is 66. It's oxide contains 53% metal . At. Wt. of metal is?
   (a) 21 (b) 54
   (c) 27.06 (d) 2.086
6. Sulphate of a metal M contains 9.87 % of M . It is isomorphous with $ZnSO_4.7H_2O$. At wt. of N is
   (a) 40.3 (b) 36.1
   (c) 24.3 (d) 11.3
7. 6.02×10^{20} molecules of urea are present in 100 ml of its solution. Conc. of Urea solution is
   (a) 0.001 m (b) 0.1 m
   (c) 0.02 m (d) 0.01 m
8. 1 mole of magnesium nitride on reaction with an excess of water gives
   (a) 1 mole of NH_3 (b) 2 moles of HNO_3
   (c) 2 moles of NH_3 (d) 1 mole of HNO_3
9. What volume of H_2 gas at 273 K and 1 atm will be consumed in obtaining 21.6 g of elementary boron by hydrogen (At. mass = 10.8) from the reduction of Boron Trichloride
   (a) 89.6 *l* (b) 67.2 *l*
   (c) 44.8 *l* (d) 22.4 *l*
10. How many moles of electron weigh 1 Kg (mass of e = 9.1×10^{-31} Kg)
    (a) 6.023×10^{23} (b) $\frac{1}{9.10} \times 10^{31}$
    (c) $\frac{6.023}{9.10} \times 10^{54}$ (d) $\frac{1}{9.1 \times 6.023} \times 10^{8}$
11. An aq. solution of 6.3 g of oxalic acid dehydrate is made up to 250 ml. the volume of 0.1 m NaOH required to completely neutralize 10 ml of This solution is:-
    (a) 40 ml (b) 20 ml
    (c) 10 ml (d) 4 ml
12. A compound possesses 8% sulphur by mass. The least molecular mass is:-
    (a)200 (b) 400
    (c)155 (d) 355
13. Molarity of liq HCl with density equal to 1.17 g/cc is :-
    (a) 36.5 (b) 18.25
    (c) 32.05 (d) 4.65
14. Which law directly explains law of conservation of mass
    (a) Daltons Law (b) Avogadrow,s Law
    (c) Berzelius Law (d) Hund's Rule
15. A mixture of Naphthalene and benzene acid can be separated by
    (a) Extraction with cold water
    (b) Sublimation
    (c) Extraction by hot water
    (d) Steam distillation

Nishant Gupta, D-122, Prashant vihar, Rohini, Delhi-85
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