Chapter: 01

Basic Concepts

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Objective

1. The atomicity of $C_6H_{12}O_6$ is:

Ⓐ 12 Ⓑ 24 Ⓒ 6 Ⓓ 3

2. The size of an atom is in the range of:

Ⓐ 2×10^{-10} m Ⓑ 2×10^{-9} m Ⓒ 2×10^{-11} m Ⓓ 2×10^{-12} m

3. During combustion analysis CO_2 produced is absorbed by:

Ⓐ $CaCl_2$ Ⓑ KOH (50%) Ⓒ P_2O_5 Ⓓ $Mg(ClO_3)_2$

4. Fractional atomic mass is mainly due to:

Ⓐ Atomic mass is average mass of isobars.
Ⓑ Elements mostly consist of isotopes having different fractional abundances.
Ⓒ Mass of atom is in fraction.
Ⓓ Atomic mass is average masses of isotopes.

5. The volume occupied by 16 g of CH_4 at S.T.P is:

Ⓐ 1.3 dm^3 Ⓑ 1.8 dm^3 Ⓒ 2.2414 dm^3 Ⓓ 22.414 dm^3

6. The pressure of vapours maintain in ionization chamber of mass spectrometer during isotopic analysis is:

Ⓐ 10^{-11} torr Ⓑ 10^{-5} torr Ⓒ 10^{-7} torr Ⓓ 10^{-9} torr

7. Volume occupied by one mole of gas at standard temperature and pressure is:

Ⓐ 22.414 dm^3 Ⓑ 2.2414 dm^3 Ⓒ 2.4 dm^3 Ⓓ 1.3 dm^3

8. 1 gram formula of NaCl is equal to:

Ⓐ 12 g Ⓑ 23 g Ⓒ 35.5 g Ⓓ 58.5 g

9. The branch of chemistry which tells us the quantitative relationship between reactant and products is called:

Ⓐ Thermometry Ⓑ Physical chemistry Ⓒ Stoichiometry Ⓓ Organic chemistry

10. Average atomic mass of Neon is:

Ⓐ 20.18 Ⓑ 22.18 Ⓒ 21.81 Ⓓ 20.81

11. The number of atoms in 1.97 g of gold and ..... g of sodium arc equal.

Ⓐ 2300 Ⓑ 230 Ⓒ 23 Ⓓ 0.23

12. Oxygen molecule is heavier than hydrogen by:

Ⓐ 32 times Ⓑ 8 times Ⓒ 16 times Ⓓ 1 time

13. A pair of elements having single isotope are:

Ⓐ $_9F^{19}$, $_{79}Au^{197}$ Ⓑ $_{33}As^{75}$, $_7N^{14}$ Ⓒ $_8O^{16}$, $_7N^{14}$ Ⓓ $_{53}I^{127}$, $_{35}Br^{81}$

14. 1 mole of CH_3OH and C_2H_5OH have:

Ⓐ Equal number of atoms Ⓑ Equal number of molecules
Ⓒ Equal number of protons Ⓓ Equal number of ions

15. In 98g of sulphuric acid H_2SO_4 number of O atoms:

Ⓐ 1.2×10^{24} Ⓑ 6.02×10^{24} Ⓒ 6.02×10^{23} **Ⓓ 2.408×10^{24}**

16. Mass of one mole of chlorine(Cl_2) gas is:

Ⓐ 44g Ⓑ 23g **Ⓒ 71g** Ⓓ 35.g

17. Many elements have fractional atomic masses. This is because:

Ⓐ Atomic masses are average masses of isobars.
Ⓑ Atomic masses are average masses of isotopes.
Ⓒ The mass of the atom is itself fractional.
Ⓓ Atomic masses are average masses of isotopes proportional to their relative abundance

18. 22g of CO_2 sample has:

Ⓐ 1.5 moles of O atoms **Ⓑ 1 mole of O atoms**
Ⓒ 6.02×102 molecules of $C0_2$ Ⓓ ½ mole of O atoms

19. How many times Na atom is heavier than H-atom?

Ⓐ 23 times Ⓑ 32 times Ⓒ 46 times Ⓓ 11 times

20. The phenomenon of isotropy was first discovered by:

Ⓐ Rutherford Ⓑ Dalton Ⓒ Berzelius **Ⓓ Soddy**

21. Mass in grams of 2.74 moles of $KMnO_4$:

Ⓐ 294 g Ⓑ 0.715g **Ⓒ 432.92 g** Ⓓ 1416.2 g

22. Which is a molecular ion?

Ⓐ CH_4^+ Ⓑ Ca^{2+} Ⓒ Na^+ Ⓓ Al^{3+}

23. Nickel has isotopes:

Ⓐ 4 Ⓑ 6 Ⓒ 8 **Ⓓ 5**

24. Which of the following is a mono-isotopic element?

Ⓐ Calcium Ⓑ Chlorine **Ⓒ Fluorine** Ⓓ Silver

25. The mass of 10 moles of electrons is:

Ⓐ 5.5 mg Ⓑ 1.84 mg Ⓒ 16.73 mg Ⓓ 10.08 mg

26. 1 amu is equal to:

Ⓐ 1.661×10^{-27} kg Ⓑ 1.661×10^{27} kg Ⓒ 1.661×10^{-30} kg Ⓓ 1.661×10^{-24} kg

27. Atoms of which one of the following element have independent existence:

Ⓐ Oxygen **Ⓑ Krypton** Ⓒ Nitrogen Ⓓ Flourine

28. Which of the following element can exist in monoatomic form?

Ⓐ Helium Ⓑ Nitrogen Ⓒ Chlorine Ⓓ Oxygen

29. The atomic masses of elements were determined by:

Ⓐ Soddy **Ⓑ J. Berzelius** Ⓒ Moseley Ⓓ John Dalton

30. Haemoglobin is a macromolecule and consists of approximately atoms:

Ⓐ 10,000 Ⓑ 15000 Ⓒ 5000 Ⓓ 28000

31. Which is not a molecular ion?

Ⓐ CO^+ Ⓑ NH_4^+ Ⓒ CH_4^+ Ⓓ He^+

32. Mass of one mole of electrons is:

Ⓐ 1.673 mg Ⓑ 1.008 mg Ⓒ 0.184 mg Ⓓ 0.55 mg

33. The mass of two moles of electrons is:

Ⓐ 1.10 mg Ⓑ 1.008 mg Ⓒ 1.673 mg Ⓓ 0.184 mg

34. Tin has isotopes:

Ⓐ 5 Ⓑ 7 Ⓒ 9 Ⓓ 11

35. Isotopes differ in number of:

Ⓐ Proton Number Ⓑ Neutron Ⓒ Electron Ⓓ Proton

36. How many isotopes are present in palladium:

Ⓐ 6 Ⓑ 5 Ⓒ 4 Ⓓ 6

37. The element nickel has isotopes:

Ⓐ 2 Ⓑ 3 Ⓒ 7 Ⓓ 5

38. Isotopes differ in:

Ⓐ Arrangement of electrons in orbitals.
Ⓑ Chemical properties.
Ⓒ Extent to which they may by affected in electromagnetic field.
Ⓓ Properties which depend upon mass.

39. Bromine has isotopes:

Ⓐ 4 Ⓑ 6 Ⓒ 8 Ⓓ 2

40. Cadmium has isotopes:

Ⓐ 3 Ⓑ 6 Ⓒ 9 Ⓓ 5

41. The number of neutron present in $^{39}_{19}K$ is:

Ⓐ 20 Ⓑ 19 Ⓒ 18 Ⓓ 29

42. The height of the peaks in a mass spectrum shows:

Ⓐ mass number of isotopes Ⓑ relative abundance of isotopes
Ⓒ number of protons Ⓓ number of isotopes

43. Ascorbic acid is vitamin:

Ⓐ C Ⓑ E Ⓒ B Ⓓ A

44. Empirical formula of glucose is:

Ⓐ CH_2O Ⓑ $C_2H_{12}O_6$ Ⓒ $C_2H_2O_2$ Ⓓ CHO

45. Empirical Formula of Benzene is:

Ⓐ C_3H_3 Ⓑ CH Ⓒ CHO Ⓓ CH_2O

46. During combustion analysis $Mg(ClO_4)_2$is employed to absorb:

Ⓐ CO Ⓑ H_2O vapours Ⓒ CH_4 Ⓓ CO_2

47. In combustion analysis H_20 vapours are absorbed by:

Ⓐ 50% KOH Ⓑ $Mg(ClO_4)_2$ Ⓒ $Mg(ClO_3)_2$ Ⓓ $Mg(ClO_2)_2$

48. The number of moles of $C0_2$ which contains 8.0g of oxygen is:

Ⓐ 0.25 Ⓑ 1.50 Ⓒ 1.0 Ⓓ 0.50

49. The number of moles of CO_2 which contains 16g of Oxygen is:

Ⓐ 0.50 Ⓑ 1.50 Ⓒ 1.0 Ⓓ 0.25

50. The mass of CO_2 containing 8 grams of oxygen (O_2) in gram is:

Ⓐ 22 Ⓑ 11 Ⓒ 32 Ⓓ 16

51. Amount of NaOH required to produce $250cm^3$ of 1M solution in grams is:

Ⓐ 25 Ⓑ 20 Ⓒ 15 Ⓓ 10

52. One mole of SO_2 contain:

Ⓐ 18.01×10^{23} molecules of $S0_2$ Ⓑ 6.02×10^{23} atoms of Sulphur
Ⓒ 6.02×10^{23} atoms of Oxygen Ⓓ 4 gram atoms of SO_2

53. The largest number of molecules are present in:

Ⓐ 5.4g of N_2O_5 Ⓑ 3.6g of H_20 Ⓒ 2.8g of CO Ⓓ 4.6g of C_2H_50H

54. 27g of Al will react completely with how much mass of O_2 to produce Al_20_3:

Ⓐ 32g of Oxygen Ⓑ 8g of Oxygen Ⓒ 16g of Oxygen Ⓓ 24g of Oxygen

55. The number of atoms present in 0.1 mole of oxygen gas are:

Ⓐ 9.03×10^{22} Ⓑ $2\times6.02\times10^{22}$ Ⓒ 3.01×10^{23} Ⓓ 6.02×10^{22}

56. The volume occupied by 1.4g of N_2 at S.T.P in dm^3 is:

Ⓐ 22.4 dm^3 Ⓑ 11.2 cm^3 Ⓒ 2.24 dm^3 Ⓓ 1.12 dm^3

57. The volume occupied by 32g of O_2 at S.T.P is:

Ⓐ 0.224 dm^3 Ⓑ 22.414 dm^3 Ⓒ 2.2414 dm^3 Ⓓ 224.414 dm^3

58. The limiting reactant is the one which:

Ⓐ Is taken in lesser quantity in volume as compared to other reactants.
Ⓑ Gives the maximum amount of the product under consideration.
Ⓒ Is taken in lesser quantity in grams as compared to other reactants.
Ⓓ Gives the minimum amount of the product which is required.

59. Percentage of nitrogen in NH_3 is:

Ⓐ $\frac{14}{17}\times100$ Ⓑ $\frac{3}{34}\times100$ Ⓒ $\frac{14}{34}\times100$ Ⓓ $\frac{3}{17}\times100$

60. The percentage of oxygen in organic compound is:

Ⓐ calculated Ⓑ experimentally Ⓒ estimated Ⓓ calorimetrically

Fill in the blanks

Q1: The unit of relative atomic mass is …… .

Q2: The exact masses of isotopes can be determined by …….. spectrograph.

Q3: The phenomenon of isotropy was first discovered by ……. .

Q4: A limiting reagent is that which controls the quantities of …….. .

Q5: 4g of CH_4 at 0°C and 1 atm pressure has …….. molecules of CH_4.

Q6: Stoichiometric calculations can be performed only when …….. is obeyed.

Answers

1.	**amu**	2.	**Mass**	3.	**Soddy**
4.	**Products**	5.	**1.505×10^{23}**	6.	**Conservation**

➢ Write "T" for a true statement and "F" for a false statement.

1. Neon has three isotopes and the fourth one with atomic mass 20.18 amu.
2. The number of atoms in 1.79 g of gold and 0.023 g of sodium are equal.
3. The number of electrons in the molecules of CO and N_2 are 14 each, so 1 g of each gas will have same number of electrons.
4. Actual yield of a chemical reaction may be greater than the theoretical yield.

Answers

1.	**False**	2.	**False**	3.	**True**	4.	**False**

Chapter : 01	***Basic Concepts***

Subjective

Q1: ***Differentiate between homo-atomic and hetero-atomic molecule with examples?***

Ans: *The difference between homo-atomic and hetero-atomic molecule is:*

Homo-atomic Molecules	***Hetero-atomic Molecules***
The molecules which consist of atoms of same element are called homo-atomic molecules. e.g. N_2, Cl_2, etc.	*The molecules which consist of atoms of different elements are called hetero-atomic molecules. e.g.* HCl, H_2SO_4 etc.

Q2: ***Calculate mass in grams of 10^{-3} moles of H_20. Molecular mass of H_20 = 18?***

Ans: No. of moles of H_20 (n) = 10^{-3} moles

Molar mass of H_20 (M) = 18g/mole

Formula:

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

$$10^{-3} = \frac{m}{18}$$

$$m = 10^{-3} \times 18g$$

$$m = 0.018g$$

Q3: ***Calculate the percentage of Nitrogen in NH_3.***

Ans: %age of Nitrogen = ?

Formula:

$$\%\text{age of Nitrogen} = \frac{\text{Mass of Nitrngen in fertilizer}}{\text{Molar mass}} \times 100$$

$$\text{Molar mass of } NH_3 = 14 + 3$$

$$\text{Molar mass of } NH_3 = 17 \text{ g mol}^{-1}$$

$$\%\text{age of Nitrogen} = \frac{\text{Mass of Nitrngen in fertilizer}}{\text{Molar mass}} \times 100$$

$$\%\text{age of Nitrogen} = \frac{14}{17} \times 100$$

$$\%\text{age of Nitrogen} = 82.35\%$$

Q4: ***Why Atomic Masses mentioned in literature are in fraction? OR Why elements have fractional atomic masses?***

Ans: *The atomic masses of most elements are fractional because they exist as a mixture of isotopes of different masses. Since the atomic mass of isotopes are different so average atomic mass is calculated for a single element.*
So, average atomic mass is in fractions as mentioned in literature.

Q5: ***Define atomic mass unit. Give its value in grams?***

Ans: *Atomic mass unit:*

The 1/12th *of the mass of one atom of carbon is called* 1 *a.m.u.*
$1 \text{ amu} = 1.661 \times 10^{-24} \text{ g}$
$1 \text{ amu} = 1.661 \times 10^{-27} \text{ kg}$

Q6: ***Why atoms cannot be observed by an ordinary optical microscope?***

Ans: *An ordinary optical microscope can measure the size of an object up to or above 500 nm. But the size of an atom is smaller. So a clear and accurate image of an object that is smaller than wavelength of visible light cannot be obtained using ordinary microscope.*

Q7: ***Differentiate between atom and molecule?***

Ans: *The difference between atom and molecule is:*

Atom	Molecule
The smallest particle of an element which may or may not have an independent existence. e.g. H, He, Ne etc.	*The smallest particle of a pure substance which can exist independently. e.g.* He, N_2, H_2 etc.

Q8: ***Define ion and give its two examples?***

Ans: *Those species which carry either positive or negative charge is called ion.*

Examples:

$$Na^+, K^+, Cl^-, Br^-$$

Q9: ***N_2 and CO have the same number of electrons, protons and neutrons. Explain with reason?***

Ans: *Number of electrons protons and neutrons in:*

	N_2	CO
	N + N	C + O
Electrons:	7 + 7 = 14	6 + 8 = 14
Protons:	7 + 7 = 14	6 + 8 = 14
Neutrons:	7 + 7 = 14	6 + 8 = 14

So N_2 *and* CO *have same number of electrons, protons and neutrons.*

Q10: ***Define atomicity? Give two examples.***

Ans: ***Atomicity:***

The number of atoms present in a molecule is called atomicity.

For example:

Atomicity of H_20 *is three and* CH_4 *is five etc.*

Q11: ***Differentiate between ion and molecular ion?***

Ans: *The difference between ion and molecular ion is:*

lons are those species which carry either positive or negative charge. Whenever an atom of an element loses one or more electrons positive ions are formed, these are called cations.

For example:

Na^+, Ca^{2+}, Mg^{2+} etc.

But when neutral atom picks up one or more electrons, a negative ion is produced which is called anions.

For example:

F^-, S^{2-}, Cl^-, Br^- etc.

When an atom loses or gains an electron, it forms an ion.

Similarly a molecule may. Also lose or gain an electron to form a molecular ion e.g., CH_4^+, CO^+, N_2^+ *etc.*

Q12: ***Define cation and anion?***

Ans: ***Cation:***

Whenever an atom of an element loses one or more electrons positive ions are formed these are called cations.

For example:

Na^+, Ca^{2+}, Mg^{2+} etc.

Anion:

When neutral atom picks up one or more electrons a negative ion is produced which is called anion.

For example:

F^-, S^{2-}, Cl^-, Br^- etc.

Q13: ***What are isotopes? Why they have same chemical but different physical properties?***

Ans: ***Isotope:***

Atoms of the same elements can possess different atomic masses but same atomic numbers, such atoms of an element are called isotopes.

For example:

Carbon has three isotopes written as ${}^{12}_{6}C$, ${}^{13}_{6}C$ *and* ${}^{14}_{6}C$ *expressed as* C-12, C-13 *and* C-14.

Isotopes of elements have same chemical properties and same position in the periodic table, but due to different atomic masses have different physical properties.

Q14: ***Explain relative abundance of isotopes with suitable example.***

Ans: *The percentage of isotope of an element in comparison to other isotopes of the same element is called its relative abundance. It is measured by mass spectrometry.*

For example:

Neon

- ${}^{20}_{10}Ne = 90.92\%$
- ${}^{21}_{10}Ne = 0.26\%$
- ${}^{22}_{10}Ne = 8.82\%$

Q15: ***How do no individual Ne atom in the sample of the element has mass of 20.18 a. m.u?***

Ans: *The overall atomic mass of neon, which is an ordinary isotopic mixture, is the average of the determined atomic masses of individual isotopes. Hence*

$$\text{Average atomic mass} = \frac{20 \times 90.92 + 21 \times 0.26 + 22 \times 8.82}{100}$$

$$\text{Average atomic mass} = 20.18$$

Hence the average atomic mass of neon is 20.18 *amu.*

It is important to realize that no individual neon atom in the sample has a mass of 20.18 amu.

Q16: ***Write only names of any four methods employed for the separation of Isotopes.***

Ans: The separation of isotopes can be done by the methods based on their properties. Some important methods are gaseous diffusion, thermal diffusion, distillation, ultracentrifuge, electromagnetic separation and laser separation.

Q17: ***Explain mathematical relationship for m/e of an ion in mass 'spectrometry?***

Ans: The mathematical relationship for (m/e) *is:*

$$\frac{m}{e} = \frac{H^2\, r^2}{2E}$$

Where H = strength of magnetic field.
E = strength of electrical field.
r = radius of circular path.

Q18: ***What is the Function of ionization Chamber in Mass Spectrometer?***

Ans: Vapours are allowed to enter the ionization chamber where fast moving electrons are thrown upon them. The atoms of isotopic elements present in the form of vapours are ionized. These positively charged ions of isotopes of an element have different masses depending upon the nature of the isotopes present in them.

Q19: ***What is mass spectrum?***

Ans: ***Mass spectrum:***

It is plot of data in such a way that m/e ratios are taken on abscissa (x axis), and relative number of ions as ordinate (y-axis).

For example:

Mass spectrum of Neon

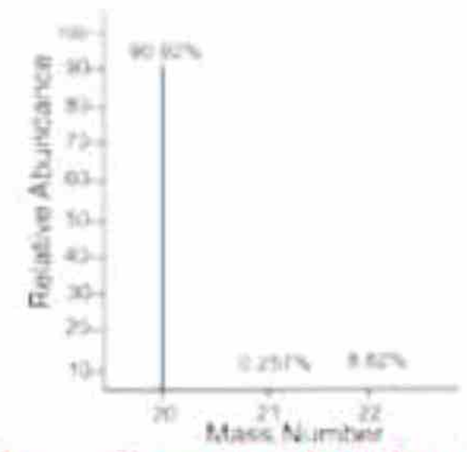

Fig Computer plotted graph for the isotopes of neon

Q20: ***What is the function of electric field in mass spectrometer?***

Ans: When a potential difference (E) of 500-2000 *volts is applied between perforated accelerating plates, then the positive ions produced in ionization chamber are strongly accelerated towards the negative plates. It is called acceleration chamber.*

Q21: ***Define empirical formula with two examples.***

Ans: ***Empirical Formula:***

It is the simplest formula that gives the small whole number ratio between the atoms of different elements present in a compound. The empirical formula of glucose ($C_6H_{12}O_6$) *is* CH_20 *and that of benzene* (C_6H_6) *is* CH.

Q22: ***Define Molar volume.***

Ans: ***Molar volume:***

The volume occupied by one mole of an ideal gas at standard temperature and pressure is called molar volume.
Its value is $22.414 dm^3$

For example:

$1 \text{ mole of } H_2 = 6.02 \times 10^{23}$

Molecules $= 22.4146 dm^3$ at S.T.P.

Q23: ***Define Molecular formula.***

Ans: ***Molecular formula:***

The formula which shows the exact number of atoms of each element present in one molecule of a compound is called molecular formula.

For example:

- Benzene (C_6H_6)
- Glucose ($C_6H_{12}O_6$)

Q24: ***Give the main points for the determination of molecular formula?***

Ans: *There are three steps involved in:*

- *Determination of percentage composition by combustion analysis.*
- *Determination of empirical formula-by using percentage composition.*
- *Calculation of molecular formula using empirical formula mass by following relation:*

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

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

Q25: ***Give two examples of compounds having same Empirical and Molecular formula?***

Ans: *The examples of compounds having same Empirical and Molecular formula are:*

- H_20
- $C_{12}H_{22}O_{11}$
- NH_3
- CO_2

Q26: ***What is the function of 50% KOH and Mg $(ClO_4)_2$ in combustion analysis?***

Ans: Mg $(ClO_4)_2$ and KOH *are used as absorber in combustion analysis. The water vapours are absorbed by* Mg $(ClO_4)_2$ *and* CO_2 *is absorbed by* 50% KOH.

Q27: ***Write down the names of water absorber and CO_2 absorber in combustion analysis experiment?***

Ans: *Name of the water absorber and* CO_2 *absorber in combustion analysis experiment is:*

- H_20 absorber = Mg $(CIO_4)_2$
- CO_2 Absorber = 50% KOH

Q28: ***Write down four steps for determination of empirical formula?***

Ans: *Empirical formula can be determined by using following steps:*

- *Determination of percentage composition.*
- *Finding the number of grams of each element.*
- *Determine atomic ratio of each element.*
- *If the atomic ratio is simple whole number it gives the empirical formula, otherwise multiply with a suitable digit to get the whole number atomic ratio.*

Q29: ***Differentiate between empirical and molecular formula.***

Ans: *The difference between empirical and molecular formula is:*

Empirical formula	***Molecular formula***
➢ *It shows simplest whole number ratio between atoms of a compound.* ➢ *Both covalent and ionic compounds have empirical formula.* ➢ *Example: Benzene* (C_6H_6) *has empirical formula as* (CH) *and of glucose is* (CH_20).	➢ *It shows exact number of atoms of each element present in a compound.* ➢ *Ionic compound do not have molecular formula.* ➢ *Example: Molecular formula of benzene is: Benzene* (C_6H_6), *Glucose* $(C_6H_{12}O_6)$.

Q30: ***Molecular formula is a nth multiple of empirical formula. Explain with examples.***

Ans:

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

Where "n" is a simple integer. Those compounds who's molecular and empirical formulas are same have "nth" value of unity:

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

For example:

Molecular formula of glucose is $(C_6H_{12}O_6)$ *and empirical formula is* CH_20.

Where $n = \frac{180}{30} = 6$

Q31: ***180g of glucose and 342g of sucrose have the same number of molecules but different number of atoms present in them. Why?***

Ans: Mass of glucose $= 180$ g

Formula:

$$\text{Number of moles of glucose} = \frac{\text{Mass of glucrose}}{\text{Molar mass of glucose}}$$

$$\text{Number of moles of glucose} = \frac{180}{180}$$

$$\text{Number of moles of glucose} = 1$$

$$\text{Mass of sucrose} = 342\text{g}$$

$$\text{Number of moles of sucrose} = \frac{\text{Mass of sucrose}}{\text{Molar mass of glucose}}$$

$$\text{Number of moles of sucrose} = \frac{342}{342}$$

$$\text{Number of moles of sucrose} = 1$$

And one mole of each compound contains 6.02×10^{23} *numbers of particles. So, both have same number of particles. Glucose* ($C_6H_{12}O_6$) *molecules have* 24 *atoms, while sucrose* ($C_{12}H_{22}O_{11}$) *molecules have* 45 *atoms. So, equal number of molecules of glucose and sucrose will have different number of atoms.*

Q32: ***Calculate the number of CO_2 molecules in 20.0 grams of it?***

Ans: No. of CO_2 molecules (N) =?

Given mass (m) $= 20$g

Molar mass of $CO_2 = 44$g/mole

Avogadro's number $N_A = 6.02 \times 10^{23}$

Molar mass of $CO_2 = 44$g/mole

Formula:

$$N = \frac{m}{M} \times N_A$$

$$N = \frac{20}{44} \times 6.02 \times 10^{23}$$

$$N = 2.73 \times 10^{23}$$

No. of CO_2 molecules (N) 2.73×10^{23}

Q33: ***Mg atom is twice heavier than that of carbon atom. Comment?***

Ans: ***As,***

12g of C = 1 mole of C = 6.02×10^{23} atoms of carbon

24 g of Mg =1 mole of Mg = 6.02×10^{23} atoms of magnesium

Equal numbers of magnesium atoms have twice mass as compared to equal number of carbon atoms. Therefore each magnesium atom is twice heavier than carbon atom.

Q34: ***One Mole of water has two moles of bonds, three moles of atoms. Explain?***

Ans: In one mole of water, moles of oxygen = 1 mole

In one mole of water, moles of hydrogen = 2 moles

In one mole of water, number of bonds = 2 moles

In one mole of water, number of electrons (8 electron of 0 and two electron of H) = 10 moles

In one mole of water, number of particle of oxygen (8P + 8N + 8e) = 24 moles

In one mole of water, number of particle hydrogen = 2P + 2e = 4 moles

So,

Total number of particles in one mole of water = 24 + 4 = 28 moles

1 mole of H_2O = 2 moles H atom +1 mole of oxygen atom = 3 moles

Q35: ***What do you know about gram atom? How we can calculate gram atom of an element? Give its relationship.***

Ans: When the substance at our disposal is an element then the atomic mass of that element expressed in grams is called one gram atom. It is also called one gram mole or simply a mole of that element.

$$\text{Number of gram atoms or moles of an element} = \frac{\text{Mass of an element in grams}}{\text{Molar mass of the substance}}$$

For example:

1 gram atom of hydrogen = 1.008g

1 gram atom of carbon = 12.000g

1 gram atom of uranium = 238.0g

Q36: ***How many oxygen atoms are present in 4.8g of Ozone (O=16 amu)?***

Ans: Molecular mass of $O_3 = 16 \times 3 = 48\ gmol^{-1}$

Formula:

$$\text{Moles of } CO_3 = \frac{\text{Mass in grams}}{\text{Formula mass}}$$

$$\text{Moles of } CO_3 = \frac{4.8g}{48gmol^{-1}}$$

$$\text{Moles of } CO_3 = 0.1 \text{ mol}$$

1 mole of O_3 has 0 moles = 3 moles of oxygen atoms

0.1 moles of O_3 has = 0.1×3

0.1 moles of O_3 has = 0.3 moles.

Q37: ***23g of Na and 238g of U-have equal number of atoms in them. Justify?***

Ans: Mass of Na = 23g

Formula:

$$\text{Number of moles of Na} = \frac{\text{Mass of Na}}{\text{Molar mass of Na}}$$

$$\text{Number of moles of Na} = \frac{23}{23}$$

Number of moles of Na = 1 mole

Mass of U = 238 g

$$\text{Number of moles of U} = \frac{\text{Mass of U}}{\text{Molar mass of U}}$$

$$\text{Number of moles of U} = \frac{238}{238}$$

Number of moles of U = 1 mole

Since one mole of each element contains 6.02×10^{23} particles.

Q38: ***Calculate mass in Kg of 2.6×10^{20} molecules of SO_2?***

Ans: No. of molecules of $S0_2$ (N)= 2.6×10^{20}

Molar mass of $S0_2$ (M) = 32 + 32

Molar mass of $S0_2$ (M) = 64 g/mole

Mass in Kg of $S0_2$ (m) = ?

Formula:

$$N = \frac{m}{M} \times N_A$$

$$m = \frac{N \times M}{N_A}$$

$$m = \frac{2.6 \times 10^{20} \times 64}{2.6 \times 10^{23}}$$

$$m = 27.64 \times 10^{-3} g$$

$$m = 27.64 \times 10^{-6} Kg$$

Q39: ***Calculate the number of moles of O atoms in*** **9.00** ***gram of*** **$Mg(NO_3)_2$?**

Ans: Formula mass of $Mg(NO_3)_2 = 24 + 14 \times 2 + 16 \times 6$

Formula mass of $Mg(NO_3)_2 = 148\ gmol^{-1}$

Formula:

$$\text{Molar of } Mg(NO_3)_2 = \frac{\text{Mass in grams}}{\text{Formula mass}}$$

$$\text{Molar of } Mg(NO_3)_2 = \frac{9g}{148\ gmol^{-1}}$$

Molar of $Mg(NO_3)_2 = 0.06$ mol

1 mole of $Mg(NO_3)_2$ has O moles = 6 moles of oxygen atoms.

0.06 moles of $Mg(NO_3)_2$ has $= 0.06 \times 6$

0.06 moles of $Mg(NO_3)_2$ has $= 0.36$ moles

Q40: ***One mole of*** **H_2SO_4** ***should completely react with two moles of*** **NaOH.** ***How does Avogadro's number help to explain it?***

Ans: One mole of H_2SO_4 *completely neutralize with two moles of* NaOH

$$NaOH \longrightarrow Na^+ + OH^-$$

$$H_2SO_4 \longrightarrow 2H^+ + SO_4^{2-}$$

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

$$2OH^- + 2H^+ \longrightarrow 2H_2O$$

I mole of H_2SO_4 *when completely ionizes produces two moles of* H^+ *ions or* $2 \times 6.02 \times 10^{23}$ *ions. Two mole of* NaOH *also produces two moles of* OH^- *ions or* $2 \times 6.02 \times 10^{23}$ *ions. When these solutions are mixed together, they have equal number of* H^+ *and* OH^- *ions and solution becomes neutral.*

Q41: ***Calculate the moles of*** **100 g** ***of silicon?*** **(Atomic Mass of Si is 28).**

Ans: Given mass of silicon (m) = 100 g

Molar mass of silicon (M) $= 28\ g\ mol^{-1}$

No. of moles (n) = ?

Formula:

$$n = \frac{m}{M} = \frac{100}{28}$$

n = 3.56 moles

Q42: ***Calculate the moles of chlorine atoms in*** **0.882g $C_2H_4Cl_2$?**

Ans: Given mass (m) = 0.822g

Number of mole of Cl atom (n) = ?

Molar mass of $C_2H_4Cl_2$ (M) $= 99\ gmol^{-1}$

By using Formula:

$$\text{No. of moles of } C_2H_4Cl_2 = \frac{0.822}{99}$$

No. of moles of $C_2H_4Cl_2 = 0.00830$

No. of moles of $C_2H_4Cl_2 = 8.30 \times 10^{-3}$

One mole of $C_2H_4Cl_2 = 2$ moles of Cl atoms

8.30×10^{-3} Moles of $C_2H_4Cl_2 = 2 \times 8.30 \times 10^{-3}$

8.30×10^{-3} Moles of $C_2H_4Cl_2 = 0.017$ moles of Cl atoms.

Q43: ***Calculate number of gram atoms of Na when its mass is*** **0.1kg*****? Atomic mass of*** **Na = 23g mol^{-1}.**

Ans: Gram atoms of Na = ?

Given mass of Na = 0.1 kg = 100 g

Atomic mass of Na = 23 g mol^{-1}

Formula:

$$n = \frac{m}{M} = \frac{100\ g}{23\ g\ mol^{-1}}$$

$n = 4.34$ moles

The gram atoms (moles) of Na are 4.34 moles.

Q44: ***How many molecules of*** **H_2O** ***are present when its amount is*** **0.25 moles*****?***

Ans: No. of H_2O molecules (N) = ?

No. of moles of H_2O (n) = 0.25 mole

Avogadro's (N_A) = 6.02×10^{23}

Number of water molecule = $n \times N_A$

Number of water molecule = $0.25 \times 6.02 \times 10^{23}$

No. of H_2O molecules (N) = 1.505×10^{23} molecules

Q45: ***One mole of*** **$K_2Cr_2O_7$** ***has thrice the number of ions than the number of formula units when ionized?***

Ans: $K_2Cr_2O_7$ ionize as:

$$K_2Cr_2O_7 \rightleftharpoons 2K^+ + Cr_2O_7^{-2}$$

This equation show that 1 mole of $K_2Cr_2O_7$ produces 2 moles of K^+ and 1 mole of $Cr_2O_7^{-2}$.

Hence total three moles of ions are produced by the ionization of 1 mole of $K_2Cr_2O_7$.

Q46: **100cm^3 of NH_3** ***gas and*** **100cm^3** ***of*** **H_2** ***gas at STP contain equal number of molecules. Justify it?***

Ans: Different gases have different molecular masses. According to Avogadro's law "Equal volumes of different gases contain equal number of molecules." As there are large spaces between gas molecule so molecular gases and molecular sizes do not affect the overall occupied volume.

Q47: ***Calculate the mass in grams of*** **2.74 moles of $KMnO_4$*****. Formula mass of*** **$KMnO_4$** ***is*** **158g mol^{-1}.**

Ans: Moles of $KMnO_4$(n) = 2.74 mole

Molar mass (M) = $39 + 55 + 16 \times 4$

Molar mass (M) = 158 g mol^{-1}

1 mole of $KMnO_4$ = 158 g

2.74 Moles of $KMnO_4 = 2.74 \times 158$

2.74 Moles of $KMnO_4$ = 432.92 g

Q48: ***Calculate number of molecules in*** **9 g** ***of ice?***

Ans: Given mass of H_2O (m) = 9 g

Molar mass of H_20 (M) = 18 g mol^{-1}

No. of molecules of H_20 in 9 g of ice (n) = ?

By using Formula:

$$N = \frac{m}{M} \times N_A$$

$$N = \frac{9}{18} \times 6.02 \times 10^{23}$$

$$N = 3.01 \times 10^{23} \text{molecules.}$$

Q49: ***Why do 2g of H_2 16g of CH_4 , 44g of CO_2 occupy separately the volume of 22.414 dm^3 although the sizes and masses of molecules of three gases are very different from each other?***

Ans: 2 g of H_2 = 1 mole = 6.02×10^{23} molecules = 22.414 dm^3 volume at STP.

16 g of CH_4 = 1 mole = 6.02×10^{23} molecules = 22.414 dm^3 volume at STP.

44 g of CO_2 = 1 mole = 6.02×10^{23} molecules = 22.414 dm^3 volume at STP.

According to Avogadro's law, equal numbers of molecules of all gases occupy same volumes at same temperature and pressure.

Since, H_2, CH_4 and CO_2 *have same number of molecules that is why these occupy same volume.*

Q50: ***Calculate mass of 10^{-3} moles of $MgSO_4$.***

Ans: $MgSO_4$ is an ionic compound. We will consider its formula mass in place of molecular mass.

$$\text{Number of gram formula or moles of a substance} = \frac{\text{Mass of the ionic substance}}{\text{Formula mass of the ionic substance}}$$

Formula mass of $MgSO_4 = 24 + 96$

Formula mass of $MgSO_4 = 120 \text{ g mol}^{-1}$

Number of moles of $MgSO_4 = 10^{-3}$ moles

By using Formula:

$$10^{-3} = \frac{\text{Mass of } MgSO_4}{120 \text{ g mol}^{-1}}$$

$$\text{Mass of } MgSO_4 = 10^{-3} \text{ moles} \times 120 \text{ g mol}^{-1}$$

$$\text{Mass of } MgSO_4 = 120 \times 10^{-3}$$

$$\text{Mass of } MgSO_4 = 0.12\text{g}$$

Q51: ***Write down limitations of a chemical equation.***

Ans: Chemical equations have certain limitations. They do not tell about the conditions and the rate of reaction. Chemical equation can even be written to describe a chemical change that does not occur.

Q52: ***Law of conservation of mass has to be obeyed during stoichiometric calculations. Justify it.***

Ans: Stoichiometric calculations depend upon balanced chemical equation, while the balanced chemical equation is written according to law of conservation of mass. So therefore law of conservation of mass is to be obeyed during stoichiometric calculations.

Q53: ***List steps involved to identify a limiting reactant?***

Ans: To identify a limiting reactant the following steps are performed:

- *Identify the reactant which produces the least amount of product as limiting reactant.*
- *Find out the number of moles of product with the help of a balanced chemical equation.*
- *Calculate the number of moles from given amount of reactant.*

Q54: ***Write two assumptions of stoichiometry?***

Ans: They are following assumptions of stoichiometry:

- *No side reaction occurs.*
- *While doing calculations, the law of conservation of mass and law of definite proportions are obeyed.*
- *All the reactants are completely converted into products.*

Q55: ***By using balanced chemical equation. What type of relationships can be studied?***

Ans: The following types of relationships can be studied with the help of balanced chemical equation.

- *Mass - Mole relationship.*
- *Mass -Volume relationship.*
- *Mass - Mass relationship.*

Q56: ***Define limiting reactant, give an example?***

Ans: The limiting reactant is a reactant that control the amount of product formed in a chemical reaction due to its smaller amount.

Example:

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

In this reaction H_2 *is a limiting reactant.*

Q57: ***How do many chemical reactions take place in our surroundings involve the limiting reactant?***

Ans: Many chemical reactions taking place in our surrounding involve the limiting reactants.

Example:

- *Combustion is a daily life reaction in which fuel is the limiting reactant and oxygen is non-limiting reactant. Once the fuel like* CH_4 *will consume the combustion will stop.*
- *Rusting of iron is another example. In this reaction iron is the limiting reactant once it will completely convert in to rust the further reaction will stop.*

Q58: ***Many chemical reactions taking place in our surroundings involve the limiting reactants.***

Ans: ***Example:***

Burning of coal.

Explanation:

In burning of coal oxygen (O_2) *and coal* (C) *are reactants. As oxygen is present in excess amount so coal* (C) *is our limiting reactants.*

$$C + O_2 \xrightarrow[(Excess)]{} CO_2$$

Q59: ***What is difference between actual yield and theoretical yield?***

Ans: The difference between actual yield and theoretical yield is:

Actual yield	***Theoretical yield***
The amount of the products obtained in chemical reaction is called the actual yield of that reaction.	*The amount of the products calculated from the balanced chemical equation represents the theoretical yield. The theoretical yield is the maximum amount of the product that can be produced by a given amount of a reactant according to the balanced chemical equation.*

Q60: ***Actual yield is usually less than theoretical yield. Why?***

Ans: Actual yield of a chemical reaction is always less than its theoretical yield due to following reaction:

- *Due to reversibility of reactions.*
- *Loss of product due to inexperienced worker, the product is lost in separation techniques e.g. crystallization, distillation, filtration etc.*
- *Some of reactants might take part in competing side reactions.*

Q61: ***Calculate %age of Nitrogen in*** $NH_2\ CO\ NH_2$ **(at mass of N=14, C=12 H=1, O=16).**

Ans:

$$\text{Molecular mass of } NH_2\ CO\ NH_2 = 14 + 2(1) + 12 + 16 + 14 + 2(1)$$

$$\text{Molecular mass of } NH_2\ CO\ NH_2 = 60 \text{ amu}$$

$$\%\text{age of Nitrogen in } NH_2\ CO\ NH_2 = \frac{28}{60} \times 100$$

$$\%\text{age of Nitrogen in } NH_2\ CO\ NH_2 = 46.66\%$$

Q62: ***Define Gram formula with one example.***

Ans: *Formula unit mass of an ionic compound expressed in grams is called one gram – formula or one mole of that ionic compound.*

For Example:

$$1 \text{ gram - formula (1mole) of NaCl} = 58.5 \text{ g mol}^{-1}$$

Q63: ***Define Avogadro's number. Give its numerical value.***

Ans: ***Avogadro's number:***

Avogadro's number is the number of atoms, molecules and ions in one-gram atom of an element, one-gram molecule of a compound and one-gram ion of a substance, respectively. To understand Avogadro's number let us consider the following quantities of substances.

1.008 g of hydrogen = 1 mole of hydrogen = 6.02×10^{23} atoms of H18 g of H_20 =1 mole of water = 6.02×10^{23} molecules of water.

Q64: ***Define mole and molar volume? OR***

Define mole with two examples.

Ans: *The atomic mass of an element, molecular mass of a molecule and formula mass of a formula unit is expressed in grams is called gram atomic mass, gram molecular mass and gram formula unit respectively. These are also called gram mole or simply mole.*

For Example:

1 gram atom of hydrogen = 1.008 g

1 gram formula of NaCl = 58.50 g

1 gram molecule of water = 18.0 g

Q65: ***Why we use the term relative atomic mass? OR***

Define Relative atomic mass. Give two examples.

Ans: *The masses of atoms are extremely small* (10^{-24} g to 10^{-22} g) *so we cannot weigh such an extremely small masses by using any sensitive balance. That's why we use relative atomic mass unit scale.*

The relative atomic masses of some elements are:

H = 1.008 amu.

O = 15.999 amu.

Q66: ***What is the molecular ion? Give an example and its generation?***

Ans: ***Molecular ion:***

When a molecule loses or gains an electron an ion is formed called molecular ions.

These ions can be generated by passing:

- *X-rays through a gas.*
- *Alpha (α) particles through a gas.*
- *High energy electron beam through a gas.*

Q67: ***What is the principle of mass spectrometry?***

Ans: *It is a technique in which gas molecules are converted to gaseous molecular ions, which are then separated on the basis of their* m/e *ratio (mass to charge ratio).*

Q68: ***Calculate the number of water molecules in 10 g of ice.***

Ans: *The number of water molecules in 10g of ice is:*

By using Formula:

$$\text{No. of } H_2O \text{ molecules} = \frac{\text{Mass of } H_2O \text{ in grams}}{\text{Molar mass of } H_2O} \times N_A$$

$$\text{No. of } H_2O \text{ molecules} = \frac{10 \text{ g}}{18 \text{ g}} \times 6.02 \times 10^{23}$$

$$\text{No. of } H_2O \text{ molecules} = 3.34 \times 10^{23} \text{ molecules of } H_2O$$

Q69: ***What is significance of John Dalton's work about atom?***

Ans: In 1808, an English school teacher, John Dalton developed an atomic theory and recognized the law of conservation of matter and the law of definite proportion could be explained by the existence of atom.

The main postulate of this theory is all matter is composed of atoms of different elements', which differ in their properties.

Q70: ***How many moles are present in 18g of H_2O?***

Ans: The moles are present in 18g of H_2O is:

$$\text{No. of moles} = \frac{\text{Mass of substance}}{\text{Molar mass of substance}}$$

$$\text{No. of moles of } H_2O = \frac{18\text{ g}}{18\text{ g}}$$

$$\text{No. of moles of } H_2O = 1 \text{ mole of water}$$

Q71: ***What is electrometer? Give its function in mass spectrometer.***

Ans: ***Electrometer:***

Electrometer is also called as ion collector, it develops the electrical current which is proportional to the number of ions falling on it and thus the current strength in each case gives the relative abundance of each of the isotope.

The relative abundance of these isotopes is recorded in the form of peaks on the graph of the relative abundance of the isotopes versus the mass number.

Q72: ***How many moles of $C0_2$ can be produced from burning one mole of octane, Mass of octane is 114?***

Ans: $$2C_8H_{18} + 25O_2 \longrightarrow 16CO_2 + 18H_2O$$

According to balanced equation, stoichiometric calculation can be done as;

2 moles of Octane producing moles of $CO_2 = 16$ moles

1 moles of Octane producing moles of $CO_2 = \frac{16}{2}$

1 moles of Octane producing moles of $CO_2 = 8$ moles of CO_2

Q73: ***What is difference between gram ion and gram atom?***

Ans: The difference between gram ion and gram atom is:

Gram ion	Gram atom
➢ *The ionic mass of ionic specie expressed in grams is called one gram ion or one mole ions.* ➢ *For example:* 1 gram ion of $CO_3^{-2} = 60$ g 1 gram ion of $OH^- = 17$ g	➢ *The atomic mass of an element expressed in grams is called one gram atom or one gram mole or simply mole of that element.* ➢ *For example:* 1 gram atom of carbon = 12.00 g 1 gram atom of hydrogen = 1.008g

Q74: ***Define Macromolecule. Give example.***

Ans: Molecules of high molecular mass usually greater than 10,000 *are called Macromolecules.*

For Example:

- *Hemoglobin*
- *Proteins*
- *Polyester etc.*

Q75: ***Write down role of magnetic separator in mass spectrometer. What is function of magnetic field in mass spectrometer?***

Ans: The magnetic separator, (also called analyzer) separates the positive ions on the basis of different mass to charge (m/e) ratios. It also makes the ions to move in circular path.

Q76: ***How many atoms are present in 0.1 g of Na-23?***

Ans: The atoms are present in 0.1 g of Na-23 is:

$$\text{No. of atoms} = \frac{\text{Mass of element in g}}{\text{Molar mass}} \times N_A$$

$$\text{No. of atoms} = \frac{0.1\ \text{g}}{23\ \text{g mol}^{-1}} \times 6.02 \times 10^{23}$$

$$\text{No. of atoms} = 2.6 \times 10^{21}\ \text{atoms}$$

Q77: ***Why mass of Ne gas is in fractions?***

Ans: *Atomic mass Neon is* 20.18 amu. *Actually, atomic mass of an element is taken as average of atomic masses of all of its isotopes proportional to their relative abundances.*

Neon has three isotopes, Ne-20, Ne - 21, Ne - 22 in %ages 90.92%, 0.26%, 8.82% *respectively.*

Q78: ***What is Justification of two strong peaks of almost equal heights in mass spectrum of Bromine?***

Ans: *Two strong peaks of almost equal length in mass spectrum of Bromine means, it has two isotopes of nearly equal %age abundance.*
E.g. Br-79 *with* 50.54% *abundance and* Br-81 *with* 49.49% *abundance white iodine has one isotope.*

Q79: ***What are mono-isotopic elements?***

Ans: *Element having single isotope is called mono-isotopic element.*

For Example:

- *Iodine (I)*
- *Gold (Au) etc.*

Q80: ***Define stoichiometry. Give two assumptions for stoichiometric calculations. Justify it.***

Ans: ***Stoichiometry:***

Stoichiometry is a branch of chemistry which tells us the quantitative relationship between reactants and products in a balanced chemical equation.
We have to assume the following conditions:

- *No side reaction occurs.*
- *All the reactants are completely converted into the products.*

Q81: ***How can the efficiency of a chemical reaction be expressed? How percentage yield is calculated?***

Ans: *A chemist is usually interested in the efficiency of a reaction. The efficiency of a reaction is expressed by comparing the actual and theoretical yields in the form of percentage* (%) *yield.*

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

Chapter : 01 Basic Concepts

Long Questions

Q1: What is combustion analysis? How can you calculate the percentage of the carbon, hydrogen and oxygen in an organic compound by this method?

Q2: Define the following and give one example of each. (i). Avogadro's number. (ii). Molecular ion. (iii). Mole. (iv). Isotopes.

Q3: Write a note on limiting reactant and explain by giving two examples. OR Define Limiting reactant and how it is identified by different steps.

Q4: Define the following terms and give two examples of each. (i) Percentage yield. (ii) Gram ion. (iii) Gram atom. (i) Gram formula.

Q5: Define the following terms with examples. (i) Molecular ion. (ii) Molar volume. (iii) Isotope. (iv) Relative atomic mass.

Q6: Explain isotopes with their relative abundance.

Q7: Ascorbic acid (Vitamin C) contains 40.92% carbon, 4.58% hydrogen and 54.5% oxygen by mass. What is the empirical formula of ascorbic acid? At, masses: C =12.01, H = 1.008, O = 16.0.

Q8: Differentiate between empirical and molecular formula. OR Write down the steps to calculate empirical formula.

Q9: The combustion analysis of an organic compound shows it to contain 65.44% carbon, 5.50% hydrogen and 29.06% oxygen. What is the empirical formula of the compound? If the molecular mass of this compound is 110.15 $gmol^{-1}$. Calculate the molecular formula of compound.

Q10: Serotonin (Molar Mass = 176 $gmol^{-1}$) is a compound that conducts nerve impulses in brain and muscles. It contains 68.2%C, 6.86%H, 15.9%N and 9.08%O. What is its Molecular formula?

Q11: Write a note on Avogadro's number.

Q12: A well-known ideal gas is enclosed in a container having volume 500 cm^3 at S.T.P its mass comes out to be 0.72g. What is the molar mass of this gas?

Q13: Calculate the number of grams of Al_2S_3 which can be prepared by the reaction of 20g of Al and 30g of sulphur. How much non-limiting reactant is in excess?

Q14: A mixture of NH_3 gas can be prepared by heating together two solids NH_4Cl and $Ca(OH)_2$. If a mixture containing 100g of each solid is heated than calculate the number of grams of NH_3 produced.

$$2NH_4Cl_{(S)} + Ca(OH)_{2(S)} \longrightarrow CaCl_{2(S)} + 2NH_{3(g)} + 2H_2O_{(l)}$$

Molar mass of $NH_4Cl = 53.5\ g\ mol^{-1}$ and $Ca(OH)_2 = 74\ g\ mol^{-1}$

Q15: Mg metal reacts with HCl to give hydrogen gas. What is the minimum volume of HCl solution (27% by weight required to produce 12.1 g H_2)? The density of HCl solution is 1.14 g cm^{-3}.

$$4Mg_{(S)} + 2HCl_{(aq)} \longrightarrow MgCl_{2(aq)} + H_{2(g)}$$

At, Mass of $Mg = 24\ g\ mol^{-1}$

At, Mass of $Cl = 35.5\ g\ mol^{-1}$

Q16: Calculate the number of grams of K_2SO_4 and water produced when 14g of KOH are reacted with excess of H_2SO_4. Also calculate the number of molecules of water produced:

$$2KOH_{(aq)} + H_2SO_{4(aq)} \longrightarrow K_2SO_{4(aq)} + 2H_2O_{(l)}$$

At, Mass of K $= 39, S = 32, O = 16$

Q17: *Define Stoichiometry. Give assumptions; Mention any two important laws which help to perform stoichiometric calculation.*

Q18: *What is difference between actual yield and theoretical yield? Why actual yield is less than theoretical yield?*

Q19: *When lime stone* ($CaCO_3$) *is roasted quick line* (CaO) *is produced according to the equation. The actual yield of* CaO is 2.5 Kg *when* 4.5 Kg *quick line of lime stone is roasted. What is the percentage yield of this reaction?*

$$CaCO_3 \longrightarrow CaO + CO_2$$

Atomic mass of $Ca = 40, C = 12, O = 16$

Q20: *Silicon Carbide (Sic)is an important ceramic material. It is produced by allowing sand* (SiO_2) *to react with Carbon at high temperature.*

$$SiO_2 + 3C \longrightarrow SiC + 2CO$$

When 100 kg *sand is reacted with excess of Carbon,* 51.4 kg *of* SiC *is produced.*

Q21: *What is the percentage yield of* SiC?

Q22: *Ethylene glycol is used as automobile antifreeze. It has* 38.7% carbon, 9.7% hydrogen *and* 51.6% oxygen. *Its molar mass is* 62.1 grams mol^{-1}. *Determine its empirical and molecular formula.*