CHEMISTRY (XI)

Chapter 1

Basic Concepts

Short Questions

1. Define atom. Give examples.

Ans: Atom is the smallest particle of an element, which can take part in a chemical reaction. For example, He and Ne, etc. have atoms, which have independent existence while atoms of hydrogen, nitrogen and oxygen do not exist independently.

2. What are the main points of Dalton's atomic theory?

Ans: In 1808, an English school teacher, John Dalton, recognized that the law of conservation of matter and the law of definite proportions could be explained by the existence of atoms. He developed an atomic theory the main postulate of which is that all matter is composed of atoms of different elements which differ in their properties.

3. Define molecule. Give examples.

Ans: A molecule is the smallest particle of a pure substance which can exist independently. For example, He, Cl_2, O_3, P_4, S_8, HCl, NH_3, H_2SO_4, $C_6H_{12}O_6$.

4. Define macromolecules. Give an example.

Ans: Some molecules are so big that they are called macromolecules. Haemoglobin is such a macromolecule found in blood. It helps to carry oxygen from our lungs to all parts of our body. Each molecule of haemoglobin is made up of nearly 10,000 atoms and it is 68,000 times heavier than a hydrogen atom.

5. Define ion. Give examples.

Ans: Ions are those species which carry either positive or negative charge. For example, Cl^-, Na^+, F^-, K^+.

6. Define cation. Give an example.

Ans: The positively charged species are called cation. For example, the most common positive ions are formed by the metal atoms such as Na^+, K^+, Ca^{2+}, Mg^{2+}, Al^{3+}.

7. Define anion. Give an example.

Ans: The negatively charged species are called anion. For example, there are many examples of negative ions which consist of group of atoms like OH^-, CO_3^{2-}, SO_4^{2-}, PO_4^{3-}, MnO_4^{1-}, $Cr_2O_7^{2-}$ etc

8. What are molecular ions? How are they formed? Give their use.

Ans: A molecule may lose or gain an electron to form a molecular ion, e.g., CH_4^+, CO^+, N_2^+ Cationic molecular ions are more abundant than anionic ones. These ions can be generated by passing high energy electron beam or α-particles or X-rays through a gas. The breakdown of molecular ions obtained from the natural products can give important information about their structure.

9. Define relative atomic mass.

Ans: Relative atomic mass is the mass of an atom of an element as compared to the mass of an atom of carbon taken as 12. Examples are:

Element	Relative Atomic Mass (amu)	Element	Relative Atomic Mass (amu)
H	1.008	Cl	35.453
O	15.9994	Cu	63.546
Ne	20.1797	U	238.0289

10. Define isotopes. Give examples.

Ans: Atoms of the same element having same atomic number but different mass number. Such atoms of an element are called isotopes.

Examples: H-1, H-2, H-3 called as protium, deuterium and tritium, C-12, C-13, C-14 and O-16, O-17, O-18

11. The atomic masses may be in fraction. Why?

Ans: The atomic masses are in fraction when elements have isotopes as the average of atomic masses of isotopes will definitely be in fraction. For example,

$$\text{Average atomic mass} = \frac{20x90.92 + 21x0.26 + 22x8.82}{100} = 20.18\,Answer$$

The atomic mass of some of the monoisotopic elements is also in fraction.

12. Why the isotopes have the same chemical properties but different physical properties?

Ans: The isotopes of the same element have same chemical properties as they have same electronic configuration and chemical properties depend upon the number of electrons. They have different physical properties as they have different number of neutrons in the nucleus and physical properties depend upon the nucleus.

13. No individual neon atom in the sample of the element has a mass of 20.18 amu. Why?

Ans: Neon has three isotopes of atomic masses 20, 21 and 22 with relative abundances as 90.92%, 0.26% and 8.82%

The relative atomic mass of neon comes out to be 20.18 amu. So, 20.18 amu is the average atomic mass of all the three isotopes of neon.

$$\text{Average atomic mass} = \frac{20 \times 90.92 + 21 \times 0.26 + 22 \times 8.82}{100} = 20.18\ \textit{Answer}$$

14. Comment on how atomic number and atomic mass define stability of isotopes.

Ans: The elements of odd atomic number almost never possess more than two stable isotopes. The elements of even atomic number usually have larger number of isotopes and isotopes whose mass numbers are multiples of four are particularly abundant. For example, ^{16}O, ^{24}Mg, ^{28}Si, ^{40}Ca and ^{56}Fe form nearly 50% of the earth's crust. Out of 280 isotopes that occur in nature, 154 have even mass number and even atomic number.

15. What are mono-isotopic elements?

Ans: The elements with only one isotope are called mono-isotopic elements. Examples, arsenic, fluorine, iodine and gold.

16. Define mole. Give examples.

Ans: When the atomic mass of an element, molecular mass of a compound, formula mass of an ionic compound and ionic mass of ionic specie is expressed in grams then it is called mole.

OR

The atomic mass, molecular mass, formula mass or ionic mass of the substance expressed in gram is called molar mass of the substance.

Examples: 1 gram atom of hydrogen = 1.008 g

1 gram molecule of water = 18 g

1 gram formula of NaCl = 58.5 g

1 gram ion of OH^- = 17 g

17. Define mass spectrometry.

Ans: Mass spectrometer is an instrument which is used to measure the exact masses of different isotopes of an element.

18. Mention the types of mass spectrometer.

Ans: Following are the types of mass spectrometer:

Aston's mass spectrograph

Designed to identify the isotopes of an element on the basis of their atomic masses.

Dempster's mass spectrometer

Designed for the identification of elements which were available in the solid state.

19. Name the steps involved in the working of mass spectrometer.

Ans: Following are the steps involved:

1. The substance whose analysis for the separation of isotopes is required is converted into the vapour state. The pressure of these vapours is kept very low, that is, 10^{-6} to 10^{-7} torr.

2. These vapours are allowed to enter the ionization chamber where fast moving electrons are thrown upon them. The atoms of isotopic element present in the form of vapours are ionized.

3. When a potential difference (E) of 500-2000 volts is applied between perforated accelerating plates, then these positive ions are strongly attracted towards the negative plate. In this way, the ions are accelerated.

4. These ions are then allowed to pass through a strong magnetic field of strength (H), which will separate them on the basis of their (m/e) values. Actually, the magnetic field makes the ions to move in a circular path. The ions of definite m/e value will move in the form of groups one after the other and fall on the electrometer.

20. Write functions of $Mg(ClO_4)_2$ and KOH in combustion analysis.

Ans: **Function of $Mg(ClO_4)_2$:-**Magnesium perchlorate acts as a dehydrating agents so it absorbs water during combustion analysis.

Function of KOH:-Potassium hydroxide has the ability to absorb carbon dioxide so it is used to absorb CO_2 produced during combustion analysis.

21. Give the mathematical relationship used to calculate mass to charge ratio.

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

$$m/e = H^2r^2/2E$$

Where H is the strength of magnetic field, E is the strength of electrical field, r is the radius of circular path.

22. What does the mathematical relationship express in mass spectrometry?

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

$$m/e = H^2r^2/2E$$

Where H is the strength of magnetic field, E is the strength of electrical field, r is the radius of circular path. If E is increased, by keeping H constant then radius will increase and positive ion of a particular m/ewill fall at a different place as compared to the first place.

23. What happens if strength of the electric field is increased in mass spectrometry?

Ans: If E is increased, by keeping H constant then radius will increase and positive ion of a particular m/e will fall at a different place as compared to the first place. This can also be done by changing the magnetic field.

24. What is the function of an electrometer?

Ans: Electrometer is also called an ion collector and develops the electrical current. The strength of the current thus measured gives the relative abundance of ions of a definite m/e value. Similarly, the ions of other isotopes having different masses are made to fall on the collector and the current strength is measured. The current strength in each case gives the relative abundance of each of the isotopes. The same experiment is performed with C-12 isotope and the current strength is compared.

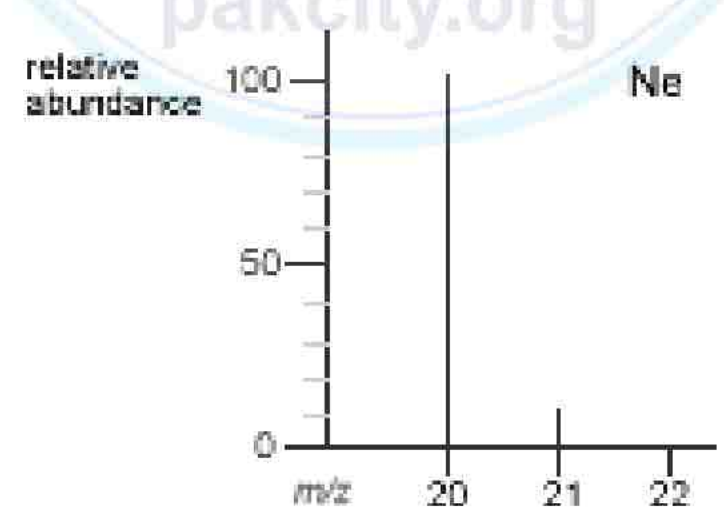
25. Mention the methods which can be used for the separation of isotopes.

Ans: Following are the methods used for the separation of isotopes:

1. Gaseous diffusion
2. Thermal diffusion
3. Distillation
4. Ultracentrifuge
5. Electromagnetic separation
6. Laser separation

26. What is mass spectrum?

Ans: The mass spectrum is a graph showing the relative abundance of isotopes plotted against the mass number.



27. Mention the formula to find percentage of element.

Ans: Following is the formula to calculate percentage of an element:

$$\text{Percentage of an element} = \frac{\text{Mass of the element in the compound}}{\text{mass of the compound}} \times 100$$

28. How can the percentage of an element be determined by formula mass?

Ans: Percentage composition of a compound can also be determined theoretically if we know the formula mass of the compound. The following equation can be used for this purpose:

$$\text{Percentage of an element} = \frac{\text{Mass of the element in one mole of the compound}}{\text{Formula mass of the compound}} \times 100$$

29. Define empirical formula.

Ans: It is the simplest formula that gives the small whole number ratio between the atoms of different elements present in a compound. In an empirical formula of a compound, AxBy, there are x atoms of an element A and y atoms of an element B. The empirical formula of glucose ($C_6H_{12}O_6$) is CH_2O and that of benzene (C_6H_6) is CH.

30. Write the steps involved in the determination of empirical formula?

Ans: Empirical formula of a compound can be calculated following the steps mentioned below:

1. Determination of the percentage composition.

2. Finding the number of gram atoms of each element. For this purpose divide the mass of each element (% of an element) by its atomic mass.

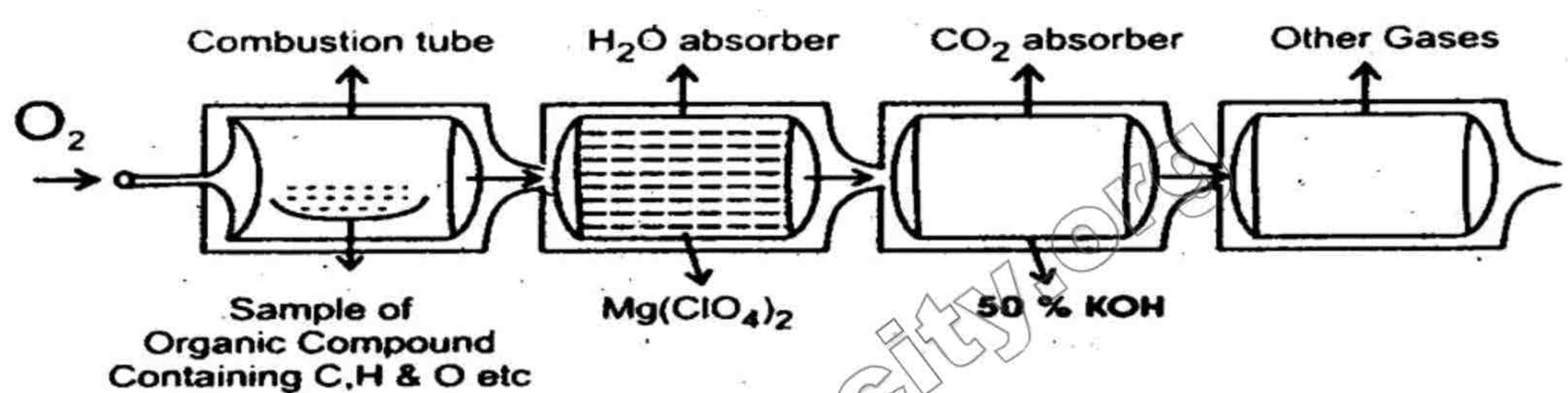
3. Determination of the atomic ratio of each element. To get this, divide the number of moles of each element (gram atoms) by the smallest number of moles.

4. 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.

31. Mention the steps involved in combustion analysis.

Ans: Following are the steps involved in combustion analysis:

1. A weighed sample of the organic compound is placed in the combustion tube. This combustion tube is fitted in a furnace. Oxygen is supplied to burn the compound.
2. Hydrogen is converted to H_2O and carbon is converted to CO_2.
3. These gases are absorbed in Mg $(ClO_4)_2$ and 50% KOH, respectively.
4. The difference in the masses of these absorbers gives us the amounts of H_2O and CO_2 produced.
5. The amount of oxygen is determined by the method of difference.



32. Mention the formulae used in combustion analysis.

Ans: Following formulae are used in combustion analysis:

$$\%\text{ of carbon} = \frac{\text{Mass of } CO_2}{\text{Mass of organic compound}} \times \frac{12.00}{44.00} \times 100$$

$$\%\text{ of hydrogen} = \frac{\text{Mass of } H_2O}{\text{Mass of organic compound}} \times \frac{2.016}{18} \times 100$$

The percentage of oxygen is obtained by the method of difference.

% of oxygen = 100 - (% of carbon + % of hydrogen) .

33. Define molecular formula.

Ans: That formula of a substance which is based on the actual molecule is called molecular formula. It gives the total number of atoms of different elements present in the molecule of a compound. For example, molecular formula of benzene is C_6H_6 while that of glucose is $C_6H_{12}O_6$.

34. Give the formula relating empirical formula with molecular formula.

Ans: Following is the formula relating empirical and molecular formula:

Molecular formula = n (Empirical formula)

Where 'n' is a simple integer.

35. Mention the compounds having same empirical and molecular formula.

Ans: Those compounds whose empirical and molecular formulae are the same are numerous. For example, H_2O, CO_2, NH_3 and $C_{12}H_{22}O_{11}$ have the same empirical and molecular formulas. Their simple multiple 'n' is unity.

36. What the value of 'n' signifies or tells in a molecular and empirical formula?

Ans: Actually the value of 'n' is the ratio of molecular mass and empirical formula mass of a substance.

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

37. Define gram atom. Give examples.

Ans: The atomic mass of the element expressed in grams is called one gram atom. It is also called one gram mole or simply a mole of that element.

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

For example,

1 gram atom of hydrogen= 1.008 g

1 gram atom of carbon= 12.000 g

1 gram atom of uranium= 238.0 g

38. Define gram molecule. Give examples.

Ans: The molecular mass of a substance expressed in grams is called gram molecule or gram mole or simply the mole of a substance.

$$\text{Number of gram molecules or moles of a molecular substance} = \frac{\text{Mass of molecular substance in grams}}{\text{Molar mass of the substance}}$$

For example,

1 gram molecule of water= 18.0 g

1 gram molecule of H_2SO_4= 98.0 g

1 gram molecule of sucrose= 342.0 g

39. Define gram formula mass. Give examples.

Ans: The formula unit mass of an ionic compound expressed in grams is called gram formula of the substance.

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

For example,

1 gram formula of NaCl= 58.50 g

1 gram formula of Na_2CO_3= 106 g

1 gram formula of $AgNO_3$= 170g

40. Define gram ion. Give examples.

Ans: Ionic mass of an ionic specie expressed in grams is called one gram ion or one mole of ions.

$$\text{Number of gram ions or moles of an species} = \frac{\text{Mass of the ionic species in grams}}{\text{Formula mass of the ionic species}}$$

For Exampal,

1 g ion of OH^- = 17 g

1 g ion of SO_4^{2-} = 96 g

1 g ion of CO_3^{2-} = 60 g

41. Define Avogadro's number. Give examples.

Ans: 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.

Examples

1.008 g of hydrogen = 1 mole of hydrogen = 6.02×10^{23} atoms of H

18 g of H_2O =1 mole of water =6.02×10^{23} molecules of water

96 g of SO_4^{2-} = 1 mole of SO_4^{2-} = 6.02×10^{23} ions of SO_4^{2-}

42. What is the number of covalent bonds in 8 g of CH_4?

Ans:

1 mole of CH_4 = 16 g

So,

0.5 mole of CH_4 = 8 g

No. of molecules = $0.5 \times 6.02 \times 10^{23}$

$= 3.01 \times 10^{23}$ molecules

Each molecule has four bonds. So,

The total number of bonds = $4 \times 3.01 \times 10^{23}$

$= 12.04 \times 10^{23}$

$= 1.204 \times 10^{24}$

43. How Avogadro's number relates to the masses of chemical substances?

Ans: Avogadro's number relates to the atoms, molecules or ions in one gram mole of an element, compound or ion. One gram mole of the substance is the atomic mass, molecular mass or ionic mass expressed in grams. It means that the number of the species is related with the masses of the species. For example, 23 grams of sodium and 238 grams of uranium have equal number of atoms in them.

1 mole of Na= 23 g= 6.02×10^{23} atoms

1 mole of U= 238 g= 6.02×10^{23} atoms

44. 180 grams of glucose and 342 grams of sucrose have same number of molecules but different number of atoms present in them. Justify it.

Ans:

180 grams of glucose = 1 mole

342 grams of sucrose = 1 mole

One mole of every substance has equal number of molecules i.e. 6.02×10^{23}

One molecule of $C_6H_{12}O_6$ has number of atoms = 24

$= 24 \times 6.02 \times 10^{23}$

$= \mathbf{1.44 \times 10^{25}}$

One molecule of $C_{12}H_{22}O_{11}$ has number of atoms = 45

$= 45 \times 6.02 \times 10^{23}$

$= \mathbf{2.70 \times 10^{25}}$

45. Mg atom is twice heavier than that of carbon atom. How?

Ans: The atomic mass of magnesium is 24 g/mol and that of carbon is 12 g/mol. That is why, Mg atom is twice heavier than carbon atom.

46. Calculate the mass in grams of 10^{-3} moles of water.

Ans: The formula applied is:

Mass of water = No. of moles × molar mass

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

$= 1.8 \times 10^{-2}$

$=$ **0.018 g**

47. Two grams of H_2, 16 g of CH_4 and 44 g of CO_2 occupy separately the volumes of 22.414 dm^3 at STP although the sizes and masses of molecules of three gases are very different from each other. Give reason.

Ans: One mole of any ideal gas at STP occupies a volume of 22.414 dm^3. The distance between the gas molecules is 300 times greater than the diameter, therefore, the sizes and masses of different gases do not affect the volume. Hence,

2 g of H_2= 1 mole = 6.02×10^{23} molecules 22.4 dm^3 at STP

16g of CH_4=1mole=6.02×10^{23}molecules22.4dm^3at STP

44g of CO_2=1mole=6.02×10^{23}molecules22.4dm^3at STP

So, they occupy a volume of 22.414 dm^3 which is called molar volume (V_m).

48. What is the number of H^+ ions in 9.8 g of H_3PO_4?

Ans: 1 mole of H_3PO_4 = 98 g

0.1 mole of H_3PO_4 = 9.8 g

$$H_3PO_4 \rightleftharpoons 3H^+ + PO_4^{3-}$$

1 molecule of H_3PO_4 has $3H^+$. So, the number of H^+ ions in 0.1 moles of H_3PO_4 is 1/10th of Avogadro's number multiplied with 3 i.e. $3\times6.02\times10^{22} = 1.806 \times 10^{23}$

49. One mole of H_2SO_4 should completely react with 2 moles of NaOH. Justify it.

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

1 mole 2 moles 1 mole 2 moles

Apply Avogadro's number concept:

1 mole of H_2SO_4generate H+ ions $=2x6.02x10^{23}H^+$ ions

2 moles of NaOH generate OH^- ions$=2x6.02x10^{23}H^+$ions

It is clear from the above calculations that the number of H^+ and OH^- ions formed are same although the number of moles of H_2SO_4 and NaOH are different that is why 1 mole of H_2SO_4 reacts completely with 2 moles of NaOH.

50. One mg of K_2CrO_4 has thrice the number of ions than the number of formula units when ionized in water. Justify.

Ans: $K_2CrO_4 \rightleftharpoons 2K^+ + CrO_4^{2-}$

Mass of $K_2Cr_2O_7$ = 1 mg = 0.001 g

Molar mass of $K_2Cr_2O_7$=294gmol^{-1}

$$\text{Number of formula units of K2Cr2O7} = \frac{\text{Mass}}{\text{Molarmass}} x N_A$$

$$= \frac{0.001 x 06.02 x 10^3}{294} \quad = 2.04 x 10^{18}$$

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

$2.04 x 10^{18}$ $\quad$ $2 x 2.04 x 10^{18}$ $\quad$ $2.04 x 10^{18} = 6.12 \times 10^{18}$ ions

Hence, it is justified that total number of ions (6.12×10^{18} ions) is thrice the number of formula units ionized ($2.04 x 10^{18}$)

51. N_2 and CO have same number of electrons, protons and neutrons. Justify.

Ans: **According to calculations for N_2**

Number of electrons in N_2 = 7 + 7 = 14

Number of protons in N_2 = 7 + 7 = 14

Number of neutrons in N_2 = 14-7 = 7 + 7 = 14

According to calculations for C and O

Number of electrons in C = 6

Number of electrons in O = 8

Total number of electrons = 6+8 = 14

Number of protons in C = 6

Number of protons in O = 8

Total number of protons = 6+8 = 14

Number of neutrons in C = 6

Number of neutrons in O = 8

Total number of neutrons = 6 + 8 = 14

Hence, N_2 and CO have same number of electrons, protons and neutrons

52. Calculate the mass in kg of 2.6×10^{23} molecules of SO_2.

Ans:

6.02×10^{23} molecules of SO_2 have mass = 64 g

One molecule of SO_2 has mass = 64 g

6.02×10^{23} molecules of SO_2 have mass = $64/ 6.02 \times 10^{23}$

2.6×10^{23} molecules of SO_2 have mass = $64 \times 2.6 \times 10^{23}/6.02 \times 10^{23}$

$= 64 \times 2.6/6.02 = 27.641$ g

53. One mole of H_2O has two moles of bonds, three moles of atoms, ten moles of electrons and twenty-eight moles of the total fundamental particles present in it. Justify.

Ans: The formula of water is:

H-O-H

i. Two moles of covalent bonds
ii. Total three moles of atoms
iii. 10 moles of electrons because 8moles of electrons are contributed by one mole of oxygen and 2 moles of electrons are contributed by 2moles of hydrogen.
iv. Number of particles in one mole of oxygen=8P+8n+8e=24 mole particles
No.of particles in two moles of Hatoms =4 mole particles
Total no.of particles=24+4=28 moles

54. Define stoichiometry.

Ans: Stoichiometry is a branch of chemistry which tells us the quantitative relationship between reactants and products in a balanced chemical equation.

55. What are the basic assumptions of stoichiometry?

Ans: Following are the basic assumptions of stoichiometry:

1. All the reactants are completely converted into products
2. No side reaction occurs
3. Law of conservation of mass and law of definite proportions are obeyed

56. Which relationships can be studied in stoichiometry?

Ans: Following relationships can be studied in stoichiometry:

1) Mass-mass Relationship

If we are given the mass of one substance, we can calculate the mass of the other substances involved in the chemical reaction.

2) Mass-mole Relationship or Mole-mass Relationship

If we are given the mass of one substance, we can calculate the moles of other substance and vice versa.

3) Mass-volume Relationship

If we are given the mass of one substance we can calculate the volume of the other substances and vice-versa.

4) Mole-mole Relationship

If we are given the moles of one substance we can calculate the moles of the other substances and vice versa.

57. What is a limiting reactant? How it helps to control the reaction?

Ans: The limiting reactant is a reactant that controls the amount of the product formed in a chemical reaction due to its smaller amount. Whenever it is consumed then the further formation of the product stops although excess amount of the other reagent is still in the reaction flask. The unavailability of limiting reactant on consumption stops product formation. This is how it controls the reaction.

For example, consider the reaction between hydrogen and oxygen to form water.

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O\ (l)$$

When we take 2 moles of hydrogen (4g) and allow it to react with 2 moles of oxygen (64g), then we will get only 2 moles (36 g) of water. Actually, we will get 2 moles (36g) of water because 2 moles (4g) of hydrogen react with 1 mole (32 g) of oxygen according to the balanced equation. Since less hydrogen is present as compared to oxygen, so hydrogen is a limiting reactant. If we would have reacted 4 moles (8g) of hydrogen with 2 moles (64 g) of oxygen, we would have obtained 4 moles (72 g) of water.

58. How can the efficiency of a chemical reaction be expressed? OR Why a chemist is interested in calculating efficiency of a reaction?

Ans: The efficiency of a chemical reaction can be expressed by the percentage yield of the chemical reaction. Percentage yield depends upon the ratio of actual yield and theoretical yield.

$$\%\,Yield = \frac{Actual\,Yield}{Theoretical\,Yield} \text{ x } 100$$

Efficiency of a reaction is also called percentage yield.

A chemist is interested in calculating efficiency of a reaction so that the results come in front of him in terms of percentage which tells about the success of a reaction.

59. Why the experimental yield is mostly less than the theoretical yield?

Ans: Following are the reasons for actual yield to be less than theoretical yield:

1. A practically inexperienced worker has many shortcomings and cannot get the expected yield.
2. The processes like filtration, separation by distillation, separation by a separating funnel, washing, drying and crystallization if not properly carried out, decrease the actual yield.
3. Some of the reactants might take part in a competing side reaction and reduce the amount of the desired product.

So in most of the reactions the actual yield is less than the theoretical yield.

60. Distinguish between actual yield and theoretical yield.

Ans: The amount of the product that is obtained in real in a chemical reaction is called actual yield. The amount of the product calculated from balanced chemical equation is called theoretical yield. For example,

$$CaCO_3\ (s) \longrightarrow CaO\ (s) + CO_2\ (g)$$

56 g of CaO is produced according to the balanced chemical equation so it is the theoretical yield. In real suppose 40 g of CaO is produced so it is the actual yield.

61. Concept of limiting reactant is not applicable to the reversible reactions. Explain it.

Ans: In case of reversible reactions we cannot understand the concept of limiting reactant as certain amounts of reactants are left unreacted at equilibrium stage. The idea of limiting reactant becomes clear when out of the two reactants one reactant gets consumed completely.

62. The reaction of combustion in atmosphere consumes O_2 which is in excess. What is limiting reactant?

Ans: In the process of combustion oxygen (O_2) is always in excess and is left in the atmosphere after complete burning of material. Hence, burning material is limiting reactant.

63. 11 g of carbon is reacted with 32 g of oxygen to give CO_2. Which is the limiting reactant?

Ans: $C + O_2 \longrightarrow CO_2$

According to the balanced equation,

12 g of C reacts with 32 g O_2 to give 44 g CO_2. Therefore, oxygen is in excess and carbon is the limiting reactant.

64. Give examples of limiting reactant in daily life.

Ans: Following are the examples of yield in daily life:

1. Burning of wood in which oxygen is in excess and wood is the limiting reactant.
2. The concept of limiting reactant is analogous to the relationship between the number of "kababs" and the "slices" to prepare "sandwiches". If we have 30 "kababs" and five breads "having 58 slices", then we can only prepare 29 "sandwiches". One "kabab" will be extra (excess reactant) and "slices" will be the limiting reactant. It is a practical problem that we cannot purchase exactly sixty "slices" for 30 "kababs" to prepare 30 "sandwiches".
3. Rusting of iron in which iron is the limiting reactant and moisture and air are in excess.

65. Mention steps to identify limiting reactant.

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

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

66. Why isotopes have same chemical but different physical properties?

Ans: Chemical properties depend upon atomic numbers i.e. arrangement of electrons. Isotopes have the same electronic configuration and as chemical properties depend upon the number of electrons present in outermost shell so isotopes have same chemical properties. Physical properties depend upon atomic masses i.e. number of protons and neutrons so isotopes have different physical properties.

67. Why oxygen cannot be determined directly in combustion analysis?

Ans: During combustion analysis, an excess of oxygen is provided to make sure that the organiccompound must be burnt to produce CO_2 and H_2O. At the end some of the oxygen is obtained as extra amount and is not surely given out by the organic compound only. So oxygen cannot be determined directly in combustion analysis. We have to subtract total percentage of C and H from100.

%age of O=100–(%age of C+% age of H)

68. How 4.9 g of $H_2SO_4$when completely ionized in water have equal number of +ve and –ve charges but the number of positively charged ions is twice the number of negatively charged ions?

Ans:

Number of moles of H_2SO_4 = 4.9/98 = 0.5 mole

0.5 mole of H_2SO_4 yields 1 mole of H^+ ions and 0.5 mole of SO_4^{2-} ions as shown below:

H_2SO_4	$\longrightarrow$	$2H^+$ +	SO_4^{2-}
0.5 mole		1 mole	0.5 mole
$0.5 \times 6.02 \times 10^{23}$		$1 \times 6.02 \times 10^{23}$	$0.5 \times 6.02 \times 10^{23}$
Molecules		ions	ions

The relationship shows that total positive charges are equal to total negative charges because each SO_4^{2-} ion has -2 charge and two H^+ have also +2 charge. However, the above relationship shows that number of positive ions are twice as compared to negative ions.

69. 23g of sodium and 238g of uranium have equal number of atoms in them. Justify.

Ans:

It is justified as follows:

	Given mass of Na	=23 g
	Atomic mass of Na	=23g/mol
	No. of moles of Na	=mass of Na/atomic mass of Na
		=23/23 =1mole
So,	1 mole of Na contains	$=6.02 \times 10^{23}$ atoms
	Given mass of U	=238g
	Atomic mass of U	=238g/mol
	No. of moles of U	=mass of Na/atomic mass of Na
		=238/238 = 1mole
So,	1 mole of U contains	$=6.02 \times 10^{23}$ atoms

70. Differentiate between empirical and molecular formula.

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. For example, the empirical formula of glucose ($C_6H_{12}O_6$) is CH_2O and that of benzene (C_6H_6) is CH.

Molecular Formula: - The formula of a substance which is based on the actual molecule is called molecular formula. It gives the total number of atoms of different elements present in the molecule of a compound. For example, the molecular formula of benzene is C_6H_6while the molecular formula of glucose is $C_6H_{12}O_6$.

71. A compound may have same molecular and empirical formula. Justify.

Ans: Some compounds have the molecules in which elements are already present in the simplest whole number ratio. So such compounds have the same empirical and molecular formula. For example, empirical and molecular formula for water is H_2O and for carbon dioxide is CO_2.

Molecular formula is related with empirical formula by the following relationship:

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

Where 'n' is a simple integer.

The value of the 'n'is the ratio of the molecular mass and empirical formula mass of a substance.

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

72. Law of conservation of mass must be considered during stoichiometric calculations. How? OR How is law of conservation of mass obeyed during stoichiometric calculations?

Ans: Law of conservation of mass must be obeyed while doing stoichiometric calculations. For example, in the following reaction of hydrogen with oxygen to form water, the mass of reactants must be equal to the mass of the products; for this purpose, the balanced chemical equation is used for stoichiometric calculations.

$$\begin{array}{ccccc} 2H2 & + & O_2 & \longrightarrow & 2H_2O \\ 4g & & 32g & & 36g \end{array}$$

Here 4 gram of hydrogen reacts with 32 g of oxygen to form 36 gram of water, so law of conservation of mass is obeyed.

73. Calculate the number of water molecules in 10 g of ice.

Ans:

Given data:

Mass of ice (water)=10g
Molar mass of ice=18g/mol
No.of molecules of water=?

Solution:

$$n= \frac{\text{Mass of ice}}{\text{molar mass of ice}} \text{xNA}$$

$$n= \frac{10}{18} \text{x}6.02\text{x}10^{23}$$

$$= 3.31\text{x}10^{23}\text{molecules}$$

74. What is Avogadro's number? Give equation to relate the Avogadro's number andmass of element.

Ans: 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.

For example,

1.008g of hydrogen = 1moleofhydrogen =6.02x 10^{23} atoms of H

18gofwater = 1mole of H_2O =6.02x 10^{23} atoms of H

96g of SO_4^{2-} =1moleof SO_4^{2-} =6.02x10^{23} ions of SO_4^{2-}

Equation:

$$\text{Number of atoms of an element} = \frac{\text{Mass of the element}}{\text{Atomic mass}} \text{x } N_A$$

75. Silver has atomic number 47 and has 16 known isotopes but two occur naturally i.e, Ag–107and Ag –109. Given the following mass spectrometric data, calculate the average atomic mass of silver.

Isotopes	Mass(amu)	Percentage abundance
^{107}Ag	106.90509	51.84
^{109}Ag	108.90476	48.16

Solution

$$\text{AverageAtomicMassofAg} = \frac{(\text{Mass of }^{107}\text{Agx \% abundance}) + (\text{Mass of }^{109}\text{Agx \% abundance})}{100}$$

$$\text{AverageAtomicMassofAg} = \frac{(106.90509\text{x}51.84) + (108.90476\text{ x}48.16)}{100}$$

$$\text{AverageAtomicMassofAg} = 107.83\ amu$$

76. Calculate the following quantities:

(a) Mass in grams of 2.74 moles of $KMnO_4$

Solution:

$KMnO_4$ (M=39 + 55 + 64 = 158 g mol$^-$)

Given Information:

Moles of $KMnO_4$ = 2.74 mol

Requirement:

Mass in grams of $KMnO_4$= ?

Formula Applied:

$$n = \frac{\text{Given mass}}{\text{Molarmass}}$$

Calculation and Result:

Mass of $KMnO_4$ = No.of moles × Molar mass

= 2.74×158

= **432.9 g**

(b) **Moles of O atoms in 9.00g of Mg $(NO_3)_2$**

Solution: **$Mg(NO_3)_2$(M=24 + 28 + 96 = 148 g mol$^-$**

Given information:

Mass of Mg $(NO_3)_2$ = 9g

Requirement:

Moles of moles of oxygen atoms = ?

Formula Applied:

$$n = \frac{\text{Given mass}}{\text{Molarmass}}$$

Calculation and Result:

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

=0.061 mol

1 mole of $Mg(NO_3)_2$contains oxygen = 6 mol

0.061 moles of $Mg(NO_3)_2$contains oxygen= 6×0.061

= **0.36mol**

(c) **Number of O atoms in 10.037 g of $CuSO_4$. $5H_2O$**

Solution: **$CuSO_4$ $5H_2O$ (M=63.5 + 32 + 64 + 90=249.5 g mol^{-1})**

Given Information:

Mass of $CuSO_4.5H_2O$ = 10.037g

Requirement:

No. of Oxygen atoms = ?

Formula Applied:

$$\text{Number of formula units} = \frac{\text{mass} \times \text{Avogadro Number}}{\text{Molar mass}}$$

Calculation and Result:

$$\text{Number of formula units of } CuSO_4.5H_2O = \frac{10.037}{249.5} \times 6.02\times10^{23}$$

$$= 2.42 \times 10^{22}$$

1 formula unit of $CuSO_4.5H_2O$contains = 9 oxygen atoms

2.41×10^{22} formula units of $CuSO_4$ $5H_2O$ contains = $9\times2.42\times10^{22}$ oxygen atoms

= 2.17×10^{23} oxygen atoms

(d) **Mass in grams of 5.136 moles of silver carbonate**

Solution:

Ag_2CO_3 M=2(107.87)+ 12+ 48 = 275.74g mol^{-1}

Given Information:

Moles of Silver Carbonate = 5.136 mol

Requirement:

Mass in grams of Silver Carbonate = ?

Formula Applied:

$$n = \frac{\text{Given mass}}{\text{Molarmass}}$$

Mass = moles x molar mass

Calculation and Result:

Mass of Ag_2CO_3 = No. of moles $\times$ molar mass

= 5.136×275.74

= **1416.2 g**

***(e)* Mass in grams of 2.78×10^{21} molecules of CrO_2Cl_2**

CrO_2Cl_2 (M)=52+32+71=155gmol⁻

Given Information:

Molecules of $CrCO_2Cl_2$ = 2.78×10^{21}

Formula Applied:

$$\text{Numberofparticles} = \frac{\text{Mass}}{\text{Molarmass}} \text{x} N_A$$

$$\text{Mass} = \frac{\text{Numberofparticles}}{N_A} \text{xMolarmass}$$

Calculation and Result:

$$\text{Mass of } CrO_2Cl_2 = \frac{2.78\times10^{21}\text{x}155}{6.02\ \text{x}10^{23}}$$

= **0.715 g**

(f) **Number of moles and formula units in 100g of $KClO_3$**

Solution: **$KClO_3$**

$M=39+35.5+48=122.5gmol^{-1}$

Given Information:

Mass of $KClO_3$ = 100g

Requirement:

Number of moles of KClO3 =?

Number of Formula units of KClO3=?

Formula Applied:

$$n = \frac{\text{Given mass}}{\text{Molarmass}}$$

Number of particles = number of moles x N_A

Calculation and Result:

i. **Calculation of Number of moles:**

Moles of $KClO_3$ = $\frac{100}{122.5}$ = 0.816 mol

ii. **Calculation of Number of formula units:**

Number of formula units of KClO3 = $0.816 \times 6.02 \times 10^{23}$

$= 4.91 \times 10^{23}$

***(g)* Number of K^+ ions, ClO^{3-} ion, Cl atoms and O atoms in it**

$KClO_3 : K^+$ ions

1 : 1

$4.91 \times 10^{23} : 4.91 \times 10^{23}$

Number of K^+ ions $= 4.91 \times 10^{23}$

KClO3: ClO^{-3}

1 1

$4.91 \times 10^{23} : 4.91 \times 10^{23}$

Number of ClO^{3-} ions $= 4.91 \times 10^{23}$

$KClO_3$: Chlorine atoms

1 1

$4.91 \times 10^{23} : 4.91 \times 10^{23}$

Number of Chlorine atoms $= 4.91 \times 10^{23}$

$KClO_3$:Oxygen atoms

1 3

$4.9 \times 10^{23} : 3 \times 4.9 \times 10^{23}$

$: 1.47 \times 10^{24}$

Number of Oxygen atoms$=1.47\text{x}10^{24}$

(h) Moles of Cl atoms in 0.822 g $C_2H_4Cl_2$

Solution: **$C_2H_4Cl_2$ (M=24 + 4+71 = 99 g mol^{-1})**

Given Information:

Mass of $C_2H_4Cl_2$ = 0.822g

Requirement:

Moles of Cl atoms = ?

Calculation and Result:

99g of $C_2H_4Cl_2$ contains moles of Cl = 2

1g of $C_2H_4Cl_2$ contains moles of Cl = $\frac{2}{99}$

0.822 g $C_2H_4Cl_2$ contains moles of Cl = $\frac{2}{99} \times 0.822$

= **0.017 mol**

(i) Mass in kilogram of 2.6×10^{20} molecules of SO_2

Solution: **SO_2 (M=32 + 32 = 64 g mol^{-1})**

Given Information:

Molecules of SO_2 = 2.6×10^{20}

Requirement:

Mass in kilogram of SO_2 = ?

Formula Applied:

$$\text{Number of particles} = \frac{\text{mass}}{\text{molar mass}} \text{ x } N_A$$

$$\text{Mass} = \frac{\text{Number of particles x molar mass}}{N_A}$$

Calculation and Result:

$$\text{Mass} = \frac{2.6\times10^{20} \text{ x } 64}{6.02\text{x}10^{23}}$$

$$= 27.64\text{x}10^{-3} \text{ g}$$

$$= 27.64\text{x}10^{-3}\text{x}10^{-3} \text{ g}$$

$$= 27.64 \times 10^{-6}\ g$$
$$= \mathbf{2.764 \times 10^{-5}\ g}$$

77. In each pair, choose the larger of the indicated quantity, or state if the samples are equal

(a) Individual particles: 0.4 mole of oxygen molecules or 0.4moles of oxygen atoms.

Solution:

Given Information:

Moles of oxygen molecules=0.4 moles Moles of oxygen atoms=0.4 moles

Requirement:

Individual particles (larger in number) =?

Calculation and Result:

0.4moles of O_2 or 0.4mol of oxygen atoms are equimolar, so **both have equal number of individual particles.**

$$\text{The number of particles} = \text{Moles} \times N_A$$
$$= 0.4 \times 6.02 \times 10^{23}$$
$$= \mathbf{2.4 \times 10^{23}}$$

(b) Mass:0.4moles of ozone molecules or 0.4moles of oxygen atoms

Solution:

Given Information:

Moles of O_3(ozone) molecules =0.4

Moles of oxygen atoms =0.4

Requirement:

Larger mass = ?

Calculation and Result:

Mass of Ozone(O_3)molecules = moles× mol.Mass

$= 0.4 \times 48 =$ **19.2 g**

Mass of oxygen atoms = moles x molar

= mass0.4 × 16

= 6.4 g

Ozone has larger mass

(c) Mass:0.6 moles of C_2H_4 or 0.6 mole of I_2

Solution:

Given Information:

Moles of C_2H_4 = 0.6

Moles of I_2 = 0.6

Requirement:

Larger mass = ?

Calculation and Result:

Mass of C_2H_4 = moles×mol.Mass

= 0.6×28=**16.8g**

Mass of I_2 = Moles x mol. mass

= 0.6×254 =**152.4 g**

Iodine has larger mass

(d) Individual particles:4.0g N_2O_4 and 3.3g SO_2

Solution:

Given Information:

Mass of N_2O_4 = 4g

Mass of SO_2 = 3.3g

Requirement:

Individual particles(in larger number) = ?

Calculation and Result:

Mass of N_2O_4 = 4g

Moles of N_2O_4 = $\frac{4}{92}$

= 0.043 mol

1mole N_2O_4 contains molecules = 6.02×10^{23}

0.043moleN_2O_4contains molecules = $6.02 \times 10^{23}\times 0.043$

= 2.6×10^{22}molecules

Mass of SO_2 = 3.3g

= $\frac{3.3}{64}$ =0.051 (moles)

1moles SO_2contains molecules = 6.02×10^{23}

0.051moleSO_2contains molecules = $6.02 \times 10^{23}\times 0.051$

= 3.1×10^{22}

SO_2 contains larger number of particles

(e) Total ions: 2.3 mol of $NaClO_3$ or 2.0 mole of $MgCl_2$

Given information

Moles of $NaClO_3$ = 2.3

Moles of $MgCl_2$ = 2.0

Requirement

Number of total ions (in larger number) = ?

Calculation and Result

Number of ions in 2.3 moles of $NaClO_3$

$NaClO_3$	$\rightarrow$	$Na^+ + ClO_3^-$	
1 mol		1 mol	1mol
Total	=	2 mol	
1 mole $NaClO_3$ contains total ions	=	2 mol	
2.3 mole $NaClO_3$ contains ions	=	2.3 x 2	
	=	4.6 mol	
	=	$4.6 \times 6.02 \times 10^{23}$	
	=	$\mathbf{2.76 \times 10^{24}}$	

$MgCl_2$	$\rightarrow$	Mg^{+2} +	$2Cl^-$
1 mol		1 mol	2mol
Total	=	3 mol	
1 moles $MgCl_2$ contains ions	=	3 mol	
2 moles $MgCl_2$ contains ions	=	$3 \times 2 = 6$ mol	
	=	$6 \times 6.02 \times 10^{23}$ ions	
	=	$\mathbf{3.61 \times 10^{24}}$	

$MgCl_2$ contains larger number of ions

(f) Molecules: 11.0 g of H_2O or 11.0 g of H_2O_2

Given Information:

Mass of H_2O = 11.0 g

Mass of H_2O_2 = 11.0 g

Requirement:

No. of Molecules (in larger number) = ?

Calculation and Result:

Number of molecules in 11g H_2O = ?

Mass of H_2O = 11g

Moles of H_2O = $\frac{11}{18}$

= 0.61 mol

Number of molecules of H_2O = No. of moles × N_A

= $0.61 \times 6.02 \times 10^{23}$ molecules

= **3.67×10^{23} molecules**

Number of molecules in 11g H_2O_2

Mass of H_2O_2 = 11g

Moles of H_2O_2 = $\frac{11}{34} = 0.32$ mol

Number of molecules of H_2O_2 = No. of moles × N_A

= $0.32 \times 6.02 \times 10^{23}$

= **1.92×10^{23} molecules**

H_2O contains larger number of molecules

(g) Na+ ion: 0.500 moles of NaBr or 0.0145 kg NaCl

Given Information:

Moles of NaBr = 0.500 moles

Mass of NaCl = 0.0145 kg

= 14.5 g

Requirement:

Number of Na^+ ions = ?

Calculation and Result:

1 moles NaBr contains Na^+ ions = 6.02×10^{23}

0.5 mole NaBr contains Na^+ ions = $6.02 \times 10^{23} \times 0.5$

= 3.01×10^{23} Na^+ ions

Mole of NaCl = $\frac{14.5}{58.5} = 0.248$ moles

1 mole NaCl contains Na^+ ions = 6.02×10^{23}

0.248 mole NaCl contains Na^+ ions = $0.248 \times 6.02 \times 10^{23}$

= 1.49×10^{23} Na^+ ions

0.5 moles NaBr contains larger number of Na^+ ions

(h) Mass: 6.02×10^{23} atoms of U^{235} or 6.02×10^{23} atoms of U^{238}

Given Information:

Atoms of U^{235} = 6.02×10^{23}

Atoms of U^{238} = 6.02×10^{23}

Requirement:

Mass (larger) = ?

Calculation and Result:

Mass of 6.02×10^{23} atoms of U^{235}

6.02×10^{23} atoms of U^{235} = 1 mole = 235 g (molar mass)

6.02×10^{23} atoms of U^{238} = 1 mole = 238 g (molar mass)

U^{238} has larger mass

78. Calculate the percentage of nitrogen in the four important fertilizers

(i) *NH_3* *(ii)NH_2CONH_2* *(iii)$(NH_4)_2SO_4$* *(iv)* *NH_4NO_3*

Solution:

Required:

%age of Nitrogen = ?

Formula Applied:

$$\%\text{age of Nitrogen} = \frac{\text{Mass of Nitrogen in fertilizer}}{\text{Molar mass}}\times100$$

(i) NH_3 (Ammonia)

Molar mass of NH_3 = $14 + 3 = 17 gmol^{-1}$

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

= $\frac{14}{17} \times 100 =$ **82.35 %**

(ii) Urea (NH_2CONH_2)

Molar mass of NH_2CONH_2 = $60 gmol^{-1}$

%age of N = 28/60 × 100= 46.67%

(iii) $(NH_4)_2SO_4$ (Ammonium Sulphate)

Molar mass of $(NH_4)_2SO_4$ = $132 gmol^{-1}$

%age of N= 28/132 ×100= **21.21%**

(iv) NH_4NO_3 (AmmoniumNitrate)

Molar mass of NH_4NO_3 = $80 gmol^{-1}$

%age of N = $\frac{28}{80} \times 100 =$ **35 %**

79. Calculate the percentage of nitrogen and phosphorus in each of the following:

(i) $NH_4H_2PO_4$ **(ii)** $(NH_4)_2HPO_4$ **(iii)** $(NH_4)_3PO_4$

Solution:

Required:

%age of Nitrogen & Phosphorous = ?

(i) $NH_4H_2PO_4$ Ammonium hydrogen phosphate

Molar mass of $NH_4H_2PO_4$ = $115\ gmol^{-1}$

% age of N = $\frac{14}{115} \times 100 = \mathbf{12.17\%}$

% age of P = $\frac{31}{115} \times 100 = \mathbf{26.96\%}$

(ii) $(NH_4)_2HPO_4$ (Diammonium Hydrogen Phosphate)

Molar mass of $(NH_4)_2HPO_4$ = $132\ g\ mol^{-1}$

% age of N = $\frac{28}{132} \times 100 = \mathbf{21.21\%}$

% age of P = $\frac{31}{132} \times 100 = \mathbf{23.48\%}$

(iii) $(NH_4)_3PO_4$ (Ammonium Phosphate)

Molar mass of $(NH_4)_3PO_4$ = $3(14 + 4) + 31 + 4(16)$

= $3(18) + 31 + 64$

= $54 + 31 + 64 = 149\ g\ mol^{-1}$

% age of N = $\frac{42}{149} \times 100$

= **28.19%**

% age of P = $\frac{\text{Mass of P}}{\text{Molar mass}} \times 100$

= $\frac{31}{149} \times 100$

= **20.8 %**