TEACHERS FORUM™

QUESTION BANK

(solved)

Class XI

CHEMISTRY

SUBJECT EXPERTS

CONTENTS

1 SOME BASIC CONCEPT OF CHEMISTRY

Every experimental measurement has some amount of uncertainty associated with it. The uncertainty in the experimental or the calculated values is indicated by mentioning the number of significant figures.

Physical properties are those properties which can be measured or observed without changing the identity or the composition of the substance. E.g. colour, odour, melting point, boiling point, density etc.

Chemical properties are characteristic reactions of different substances; these include acidity or basicity, combustibility etc.Many properties of matter such as length, area, volume, etc., are quantitative in nature.

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

Temperature-:There are three common scales to measure temperature

°C (degree celsius), °F (degree Fahrenheit) and K (kelvin). Here, K is the SI unit.

Significant Figures : The reliability of a measurement is indicated by the number of digits used to represent it. To express it more accuratelywe express it with digits that are known with certainty. These are called as Significant figures. They contain all the certain digits plus one doubtful digit in a number.

Elements : An element is the simplest form of matter that cannot be split into simpler substances or built from simpler substances by any ordinary chemical or physical method. There are 114 elements known to us, out of which 92 are naturally occurring while the rest have been prepared artificially.

Compounds : A compound is a pure substance made up of two or more elements combined in a definite proportion by mass, which could be split by suitable chemical methods.

Mixtures : A mixture is a combination of two or more elements or compounds in any proportion so that the components do not lose their identity. Air is an example of a mixture Mixtures are of two types, homogeneous and heterogeneous.

Homogeneous mixtures have the same composition throughout the sample. The components of such mixtures cannot be seen under a powerful microscope. They are also called solutions. Examples of homogeneous mixtures are air, seawater, gasoline, brass etc.

Heterogeneous mixtures consist of two or more parts (phases), which have different compositions. These mixtures have visible boundaries of separation between the different constituents and can be seen with the naked eye e.g., sand and salt, chalk powder in water etc.

LAWS OF CHEMICAL COMBINATIONS

Law of Conservation of Mass : It was given by Antoine Lavoisier in 1789. It states that matter (mass) can neither be created nor destroyed.

Law of Definite Proportions or Law of Constant Composition: This law was proposed by Louis Proust in 1799. It states that 'A chemical compound always consists of the same elements combined together in the same ratio, irrespective of the method of preparation or the source from where it is taken'.

Law of Multiple Proportions Proposed by Dalton in 1803. It states that: 'When two elements combine to form two or more compounds, then the different masses of one element, which combine with a fixed mass of the other, bear a simple ratio to one another'.

For example, hydrogen combines with oxygen to form two compounds, namely, water and hydrogen peroxide.

Hydrogen + Oxygen → Water

2g 16g 18g

Hydrogen + Oxygen → Hydrogen Peroxide

2g 2g 34g

Here, the masses of oxygen (i.e. 16 g and 32 g) which combine with a fixed mass of hydrogen (2g) bear a simple ratio, i.e. 16:32 or 1: 2.

Gay Lussac's Law of Gaseous Volumes : According to this law when gases combine or are produced in a chemical reaction they do so in a simple ratio by volume provided all gases are at same temperature and pressure.

Avogadro Law : According to this law equal volumes of gases at the same temperature and pressure should contain equal number of molecules.

Mole Concept : Mole is defined as the amount of a substance, which contains the same number of chemical units (atoms, molecules, ions or electrons) as there are atoms in exactly 12 grams of pure carbon-12.

An empirical formula represents the simplest whole number ratio of various atoms present in a compound. E.g. CH is the empirical formula of benzene.

The molecular formula shows the exact number of different types of atoms present in a molecule of a compound. E.g. C6H6 is the molecular formula of benzene.

Relationship between empirical and molecular formulae :

Molecular formula = n x empirical formula

Limiting Reagent : The reactant which gets consumed first or limits the amount of product formed is known as limiting reagent.

For e.g. in the reaction $2SO_2(g) + O_2(g) \longrightarrow 2\,SO_3(g)$, 2 moles of SO_2 reacts completely with 1 mole of O_2 to form 2 moles of SO_3. If we take 10 moles each of SO_2 and O_2, we get only 10 moles of SO_3 because 10 moles of SO_2 requires only 5 moles of O_2 for the complete reaction. So here SO_2 is the limiting reagent and 5 moles of O_2 remains unreacted.

Density : Density of a substance is its amount of mass per unit volume.

$$\text{SI unit of density} = \frac{\text{SI unit of mass}}{\text{SI unit of volume}}$$

$$= \frac{\text{kg}}{\text{m}^2} \text{ or kg m}^{-3}$$

The temperatures on two scales (°F abd °C are related to each other by the following relationship: $°F = \frac{9}{5}(°C) - 32$

The kelvin scale is related to celsius scale as : K = °C + 273.15

Molecular Mass : Molecular mass is the sum of atomic masses of the elements present in a molecule. It is obtained by multiplying the atomic mass of each element by the number of its atoms and adding them together.

One mole is the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 g (or 0.012 kg) of the ^{12}C isotope.

1 mol of hydrogen atoms = 6.022×10^{23} atoms

1 mol of water molecules = 6.022×10^{23} water molecules

1 mol of sodium chloride = 6.022×10^{23} formula units of sodium chloride

The mass of one mole of a substance in grams is called its molar mass.

$$\text{Mass \% of an element} = \frac{\text{mass of that element in the compound x 100}}{\text{molar mass of the compound}}$$

Mole Fraction : It is the ratio of number of moles of a particular component to the total number of moles of the solution. If a substance 'A' dissolves in substance 'B' and their number of moles are n_A and n_B respectively; then the mole fractions of A and B are given as

$$\text{Mole fraction of A} = \frac{\text{No. of moles of A}}{\text{No. of moles of solution}} = \frac{n_A}{n_A + n_B}$$

$$\text{Mole fraction of B} = \frac{\text{No. of moles of B}}{\text{No. of moles of solution}} = \frac{n_B}{n_A + n_B}$$

Molarity : It is defined as the number of moles of the solute in 1 litre of the solution. Thus,

$$\text{Molarity (M)} = \frac{\text{No. of moles of solute}}{\text{Volume of solution in litres}}$$

NCERT SOLUTIONS

1. Calculate the molecular mass of the following:

(i) H_2O (ii) CO_2 (iii) CH_4

Ans. **(i)** The molecular mass of H_2O

= (2 × Atomic mass of hydrogen) + (1 × Atomic mass of oxygen)

= [2(1.0084) + 1(16.00 u)]

= 2.016 u + 16.00 u = 18.02 u

(ii) The molecular mass of CO_2

= (1 × Atomic mass of carbon) + (2 × Atomic mass of oxygen)

= [1(12.011 u) + 2 (16.00 u)]

= 12.011 u + 32.00 u = 44.01 u

(iii) The molecular mass of CH_4

= (1 × Atomic mass of carbon) + (4 × Atomic mass of hydrogen)

= [1(12.011 u) + 4 (1.008 u)]

= 12.011 u + 4.032 u = 16.043 u

2. Calculate the mass percent of different elements present in sodium sulphate (Na_2SO_4).

Ans. Molar mass of Na_2SO_4 = [(2 × 23.0) + (32.0) + 4 (16.0)] = 142 g

Mass percent of sodium = $\frac{46.0\text{ g}}{142\text{ g}}$ x 100 = 32.39%

Mass percent of sulphur = $\frac{32\text{ g}}{142\text{ g}}$ x 100 = 22.54%

Mass percent of oxygen = $\frac{64.0\text{ g}}{142.0\text{ g}}$ x 100 = 45.07%

3. Determine the empirical formula of an oxide of iron which has 69.9% iron and 30.1% dioxygen by mass.

Ans. % of iron by mass = 69.9 %, % of oxygen by mass = 30.1 %

Relative moles of iron in iron oxide = $\frac{\text{\% of iron by mass}}{\text{Atomic mass of iron}} = \frac{69.9}{55.85} = 1.25$

Relative moles of oxygen in iron oxide = $\frac{\text{\% of oxygen by mass}}{\text{Atomic mass or oxygen}} = \frac{30.1}{16.00} = 1.88$

Simplest molar ratio of iron to oxygen = 1.25: 1.88 = 1: 1.5 = 2: 3

$\therefore$ The empirical formula of the iron oxide is Fe_2O_3.

4. Calculate the amount of carbon dioxide that could be produced when

(i) 1 mole of carbon is burnt in air.

(ii) 1 mole of carbon is burnt in 16 g of dioxygen.

(iii) 2 moles of carbon are burnt in 16 g of dioxygen.

Ans. The balanced reaction of combustion of carbon can be written as:

$C + O_2 \rightarrow CO_2$

(i) As per the balanced equation, 1 mole of carbon burns in1 mole of dioxy gen (air) to produce1 mole of carbon dioxide.

So CO_2 produced = 12g + 32g = 44 g.

(ii) According to the question, only 16 g of dioxygen is available. Hence, it will react with 0.5 mole of carbon to give 22 g of carbon dioxide.

(iii) Here, only 16 g of dioxygen is available. It is a limiting reactant. Thus, 16 g of dioxygen can combine with only 0.5 mole of carbon to give 22 g of carbon dioxide.

5. Calculate the mass of sodium acetate (CH_3COONa) required to make 500 mL of 0.375 molar aqueous solution. Molar mass of sodium acetate is 82.0245 g mol^{-1}

Ans. 0.375 M aqueous solution of sodium acetate means that 1000 mL of solution contain

= 0.375 M of sodium acetate

$\therefore$ Number of moles of sodium acetate in 500 mL = $\frac{0.375}{1000}$ x 500 = 0.1875 mole

Given, Molar mass of sodium acetate = 82.0245 g $mole^{-1}$

$\therefore$ Required mass of sodium acetate = 82.0245 g mol^{-1} × 0.1875 mole = 15.38 g

6. Calculate the concentration of nitric acid in moles per litre in a sample which has a density 1.41 g mL^{-1} and the mass percent of nitric acid in it being 69%.

Ans. Molar mass of HNO_3 = 1 + 14 + 3(16) = 63 g mol^{-1}

$\therefore$ Number of moles in 69 g of HNO_3 = $\frac{69\text{ g}}{63\text{ g mol}^{-1}}$ = 1.095 mol

Volume of 100g of nitric acid solution = $\frac{\text{Mass of solution}}{\text{density of solution}}$

$= \frac{100\text{ g}}{1.41\text{ g mL}^{-1}} = 70.92\text{ mL} = 0.07092\text{ L}$

$\therefore$ Concentration of nitric acid = $\frac{1.095\text{ mole}}{0.07092\text{ L}}$ = 15.44 mol/L

7. How much copper can be obtained from 100 g of copper sulphate ($CuSO_4$)?

Ans. 1 mole of $CuSO_4$ contains 1 mole of copper.

Molar mass of $CuSO_4$ = 63.5 + 32.00 + 4(16.00) = 159.5 g

159.5 g of $CuSO_4$ contains 63.5 g of copper.

$\therefore$ Amount of copper that can be obtained from 100 g $CuSO_4$ = $\frac{63.5 \text{ x } 100\text{ g}}{159.5}$ = 39.81 g

8. Determine the molecular formula of an oxide of iron in which the mass per cent of iron and oxygen are 69.9 and 30.1 respectively. Given that the molar mass of the oxide is 159.69 g mol^{-1}.

Ans. The empirical formula of the oxide is Fe_2O_3. [From Qn No:3]

Empirical formula mass of Fe_2O_3 = 2(55.85) + 3(16.00) = 159.79 g

$$\therefore\ n = \frac{\text{Molar mass}}{\text{Emprical formula mass}} = \frac{159.8\text{ g}}{159.7\text{ g}} = 0.999 = 1 \text{ (apporx)}$$

So, the molecular formula is same as empirical formula (Fe_2O_3).

9. Calculate the atomic mass (average) of chlorine using the following data:

	% Natural Abundance	Molar Mass
^{35}Cl	75.77	34.9689
^{37}Cl	24.23	36.9659

Ans. The average atomic mass of chlorine = 0.7577 × 34.9689 + 0.2423 × 36.9659

= 26.4959 + 8.9568 = 35.4527 u

10. In three moles of ethane (C_2H_6), calculate the following:

(i) Number of moles of carbon atoms.

(ii) Number of moles of hydrogen atoms.

(iii) Number of molecules of ethane.

Ans. (i) 1 mole of C_2H_6 contains 2 moles of carbon atoms.

$\therefore$ Number of moles of carbon atoms in 3 moles of C_2H_6 = 2 × 3 = 6 moles

(ii) 1 mole of C_2H_6 contains 6 moles of hydrogen atoms.

$\therefore$ Number of moles of carbon atoms in 3 moles of C_2H_6 = 3 × 6 = 18 moles

(iii) 1 mole of C_2H_6 contains 6.023 × 10^{23} molecules of ethane.

$\therefore$ Number of molecules in 3 moles of C_2H_6 = 3 × 6.02 × 10^{23} = 18.06 × 10^{23} molecules.

11. What is the concentration of sugar ($C_{12}H_{22}O_{11}$) in mol L^{-1} if its 20 g are dissolved in enough water to make a final volume up to 2 L?

Ans. Molar mass of ($C_{12}H_{22}O_{11}$) = (12 × 12) + (1 × 22) + (11 × 16)g = 342 g

$$\text{No. of moles} = \frac{\text{Given mass}}{\text{Molar mass}} = \frac{20}{342} = 0.0585 \text{ mole.}$$

$$\therefore\ \text{Molar concentration} = \frac{\text{No. of moles of solute}}{\text{Volume of solution in L}} = \frac{0.0585}{2\text{ L}} = 0.0293 \text{ mol/L}$$

12. If the density of methanol is 0.793 kg L^{-1}, what is its volume needed for making 2.5 L of its 0.25 M solution?

Ans. Molar mass of methanol (CH_3OH) = (1 × 12) + (4 × 1) + (1 × 16)

$= 32\ g\ mol^{-1} = 0.032\ kg\ mol^{-1}$

Molarity of methanol solution $= \dfrac{0.793\ kg\ L^{-1}}{0.032\ kg\ mol^{-1}} = 24.78\ mol\ L^{-1}$

Now, $M_1V_1 = M_2V_2$

$(24.78\ mol\ L^{-1})\ V_1 = (2.5\ L)\ (0.25\ mol\ L^{-1})$

$\Rightarrow\ V_1 = 0.0252\ L = 25.22\ mL$

13. Pressure is determined as force per unit area of the surface. The SI unit of pressure, Pascal is as shown below:

$1Pa = 1N\ m^{-2}$

If mass of air at sea level is $1034\ g\ cm^{-2}$, calculate the pressure in Pascal.

Ans. Pressure is defined as force acting per unit area of the surface.

$$P = \frac{F}{A} = \frac{mg}{A} = \frac{1034\ g \times 9.8\ ms^{-2}}{cm^2} \times \frac{1\ kg}{1000\ g} \times \frac{(100)^2\ cm^2}{1\ m^2}$$

$= 1.01332 \times 10^5\ kg\ m^{-1}s^{-2} = 1.01332 \times 10^5\ Pa$

14. What is the SI unit of mass? How is it defined?

Ans. The SI unit of mass is kilogram (kg).

15 Match the following prefixes with their multiples:

	Prefixes	Multiples
(i)	micro	10^6
(ii)	deca	10^9
(iii)	mega	10^{-6}
(iv)	giga	10^{-15}
(v)	femto	10

Ans.

	Prefixes	Multiples
(i)	micro	10^{-6}
(ii)	deca	10
(iii)	mega	10^6
(iv)	giga	10^9
(v)	femto	10^{-15}

16. What do you mean by significant figures?

Ans. Every experimental measurement has some amount of uncertainty associated with it.

The uncertainty in the experimental or the calculated values is indicated by mentioning the number of significant figures. Significant figures are meaningful digits which are known with certainty.

17. A sample of drinking water was found to be severely contaminated with chloroform, $CHCl_3$, supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass).

(i) Express this in percent by mass.

(ii) Determine the molality of chloroform in the water sample.

Ans. (i) 1 ppm is equivalent to 1 part out of 1 million (10^6) parts.

$\therefore$ Mass percent of 15 ppm chloroform in water $= \frac{15}{10^6} \times 100$

$= 1.5 \times 10^{-3}$ %

Molar mass of $CHCl_3 = 12 + 1 + 3(35.5) = 119.5\ g\ mol^{-1}$

(ii) 100 g of the sample contains 1.5×10^{-3} g of $CHCl_3$.

$\Rightarrow$ 1000 g of the sample contains 1.5×10^{-2} g of $CHCl_3$.

$\therefore$ Molality of chloroform in water $= \frac{1.5 \times 10^{-2}\ g}{\text{Molar mass of } CHCl_3} = \frac{1.5 \times 10^{-2}}{119.5} = 1.25 \times 10^{-4}$ m

18. Express the following in the scientific notation:

(i) 0.0048 (ii) 234,000 (iii) 8008 (iv) 500.0 (v) 6.0012

Ans. (i) $0.0048 = 4.8 \times 10^{-3}$ (ii) $234,000 = 2.34 \times 10^5$

(iii) $8008 = 8.008 \times 10^3$ (iv) $500.0 = 5.000 \times 10^2$

(v) $6.0012 = 6.0012 \times 10^0$

19. How many significant figures are present in the following?

(i) 0.0025 (ii) 208 (iii) 5005 (iv) 126,000 (v) 500.0 (vi) 2.0034

Ans. (i) 2 (ii) 3 (iii) 4 (iv) 3 (v) 4 (vi) 5

20. Round up the following upto three significant figures:

(i) 34.216 (ii) 10.4107 (iii) 0.04597 (iv) 2808

Ans. (i) 34.2 (ii) 10.4 (iii) 0.0460 (iv) 2810

21. The following data are obtained when dinitrogen and dioxygen react together to form different compounds:

Mass of dinitrogen Mass of dioxygen

(i) 14 g	16 g
(ii) 14 g	32 g
(iii) 28 g	32 g
(iv) 28 g	80 g

(a) Which law of chemical combination is obeyed by the above experimental data? Give its statement.

(b) Fill in the blanks in the following conversions:

(i) 1 km = ... mm = ... pm

(ii) 1 mg = ... kg = ... ng

(iii) 1 mL =..................... L = ... dm^3

Ans. (a) The given experimental data obeys the law of multiple proportions.

The law states that if two elements combine to form more than one compound, then the masses of one element that combines with the fixed mass of another element are in the ratio of small whole numbers.

(b) (i) $1\text{ km} = 1\text{ km} \times \frac{1000\text{ m}}{1\text{ km}} \times \frac{100\text{ cm}}{1\text{ m}} \times \frac{10\text{ mm}}{1\text{ cm}} = 10^6\text{ mm}$

$1\text{ km} = 1\text{ km} \times \frac{1000\text{ m}}{1\text{ km}} \times \frac{1\text{ pm}}{10^{-12}\text{ m}} = 10^{15}\text{ pm}$

(ii) $1\text{ mg} = 1\text{ mg} \times \frac{1\text{ g}}{1000\text{ mg}} \times \frac{1\text{ kg}}{1000\text{ g}} = 10^{-6}\text{ kg}$

$1\text{ mg} = 1\text{ mg} \times \frac{1\text{ g}}{1000\text{ mg}} \times \frac{1\text{ ng}}{10^{-9}\text{ g}} = 10^6\text{ ng}$

(iii) $1\text{ mL} = 1\text{ mL} \times \frac{1\text{ L}}{1000\text{ mL}} = 10^{-3}\text{ L}$

$1\text{ mL} = 1\text{ cm}^3 = 1\text{ cm}^3 \times \frac{1\text{ dm} \times 1\text{ dm} \times 1\text{ dm}}{10\text{ cm} \times 10\text{ cm} \times 10\text{ cm}}\text{ cm}^3 = 10^{-3}\text{ dm}^3$

22. If the speed of light is $3.0 \times 10^8\text{ m s}^{-1}$, calculate the distance covered by light in 2.00 ns.

Ans. Distance travelled = Speed of light × Time taken

$= (3.0 \times 10^8\text{ ms}^{-1})(2.00 \times 10^{-9}\text{ s})$

$= 6.00 \times 10^{-1}\text{ m} = 0.600\text{ m}$

23. In a reaction : $A + B_2 \rightarrow AB_2$

Identify the limiting reagent, if any, in the following reaction mixtures.

(i) 300 atoms of A + 200 molecules of B (ii) 2 mol A + 3 mol B

(iii) 100 atoms of A + 100 molecules of B (iv) 5 mol A + 2.5 mol B

(v) 2.5 mol A + 5 mol B

Ans. (i) According to the given reaction, 1 atom of A reacts with 1 molecule of B. Thus,

200 molecules of B will react with 200 atoms of A, and 100 atoms of A will be left unused. So, B is the limiting reagent.

(ii) According to the reaction, 1 mole of A reacts with 1 mole of B. Thus, 2 mole of A will react with only 2 mole of B. So, A is the limiting reagent.

(iii) According to the given reaction, 1 atom of A combines with 1 molecule of B.

Thus, all 100 atoms of A will combine with all 100 molecules of B. Hence, the mixture is stoichiometric where no limiting reagent is present.

(iv) 1 mole of atom A combines with 1 mole of molecule B. So, 2.5 mole of B will combine with only 2.5 mole of A. So, B is the limiting reagent.

(v) Here, 1 mole of atom A combines with 1 mole of molecule B. So, 2.5 mole of A will combine with only 2.5 mole of B and the remaining 2.5 mole of B will be left as such. So, A is the limiting reagent.

24. Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation:

$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$

(i) Calculate the mass of ammonia produced if 2.00×10^3 g dinitrogen reacts with

1.00×10^3 g of dihydrogen.

(ii) Will any of the two reactants remain unreacted?

(iii) If yes, which one and what would be its mass?

Ans. (i) $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$

From the equation, 1 mole of N_2 ie., 28 g reacts with 3 mole of H_2 ie., 6 g of H_2.

$\Rightarrow$ 2000 g of N_2 will react with $\frac{6\text{ g}}{28\text{ g}}$ x 2000 g = 428.6 g of H_2.

Given, Amount of dihydrogen = 1.00×10^3 g

Hence, N_2 is the limiting reagent.

$\therefore$ 28 g of N_2 produces 34 g of NH_3.

So, mass of ammonia produced by 2000 g of N_2 = $\frac{34\text{ g}}{28\text{ g}}$ x 2000 g = 2428.57 g

(ii) Since H_2 is the excess reagent, it will remain unreacted.

(iii) Mass of H_2 left unreacted = 2000 g – 428.6 g = 571.4 g

25. How are 0.50 mol Na_2CO_3 and 0.50 M Na_2CO_3 different?

Ans. Molar mass of Na_2CO_3 = 2 × 23 + 12 + 3 × 16 = 106 g mo^{l-1}

Now, 1 mole of Na_2CO_3 means 106 g of Na_2CO_3.

$\therefore 0.5$ mol of $Na_2CO_3 = \dfrac{106\text{ g}}{1\text{ mole}}$ x 0.5 mol $Na_2\ CO_3 = 53$ g

ie., 53 g of Na_2CO_3 is present in 1 L of water.

26. If ten volumes of dihydrogen gas react with five volumes of dioxygen gas, how many volumes of water vapour would be produced?

Ans. Reaction of H_2 and O_2 can be written as: $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

Thus, two volumes of H_2 react with one volume of O_2 to produce two volumes of water vapour.

Hence, ten volumes of H_2 will react with five volumes of O_2 to produce ten volumes of water vapour.

27. Convert the following into basic units:

(i) 28.7 pm (ii) 15.15 pm (iii) 25365 mg

Ans. (i) 1 pm = 10^{-12} m

$\therefore 28.7$ pm = 28.7 × 10^{-12} m = 2.87 × 10^{-11} m

(ii) 1 pm = 10^{-12} m

$\therefore 15.15$ pm = 1.515 × 10^{-12} m

(iii) 1 mg = 10^{-3} g

25365 mg = 2.5365 × 10^4 × 10^{-3} g

Since, 1 g = 10^{-3} kg

2.5365 × 10^1 g = 2.5365 × 10^1 × 10^{-3} kg = 2.5365 × 10^{-2} kg

28. Which one of the following will have largest number of atoms?

(i) 1 g Au (s) (ii) 1 g Na (s)

(iii) 1 g Li (s) (iv) 1 g of Cl_2(g)

Ans. (i) 1 g of Au = $\dfrac{1}{197}$ mol of Au = $\dfrac{6.02 \times 10^{23}}{197}$ atoms of Au = 3.06 × 10^{21} atoms of Au (s)

(ii) 1 g of Na = $\dfrac{1}{23}$ mol of Na = $\dfrac{6.02 \times 10^{23}}{23}$ atoms

(iii) 1 g of Li $= \frac{1}{7}$ mol of Li $= \frac{6.02 \times 10^{23}}{7}$ atoms of Li

(iv) 1 g of $Cl_2 = \frac{1}{71}$ mol of $Cl_2 = \frac{6.022 \times 10^{23}}{71}$ atoms of Cl_2

So, 1 g of Li will have the largest number of atoms.

29. Calculate the molarity of a solution of ethanol in water in which the mole fraction of ethanol is 0.040 (assume the density of water to be one).

Ans. Mole fraction of $C_2H_5OH = \frac{\text{Number of moles of } C_2H_5OH}{\text{Number of moles of solution}}$

$$\text{ie., } 0.040 = \frac{n(C_2H_5OH)}{n(C_2H_5OH) + n(H_2O)} \qquad \text{.......... (1)}$$

Number of moles present in 1 L water $= \frac{1000 \text{ g}}{18 \text{ g mol}^{-1}} = 55.55$ mol

$$(1) \Rightarrow \frac{n(C_2H_5OH)}{n(C_2H_5OH) + 55.55} = 0.040$$

$$n\,C_2H_5OH = 0.040\,n\,(C_2H_5OH) + (0.040)\,(55.55)$$

$$0.96\,n\,C_2H_5OH = 0.040 \times 55.55$$

$$\therefore\ n\,C_2H_5OH = \frac{0.040 \times 55.55}{0.96} = 2.314 \text{ mol}$$

$$\therefore \text{ Molarity of solution} = \frac{2.314 \text{ mol}}{1\text{L}} = 2.314 \text{ M}$$

30. What will be the mass of one ^{12}C atom in g?

Ans. 1 mole of carbon atoms $= 6.023 \times 10^{23}$ atoms of carbon = 12 g of carbon

$\therefore$Mass of one ^{12}C atom $= \frac{12 \text{ g}}{6.022 \times 10^{23}} = 1.993 \times 10^{-23}$ g

31. How many significant figures should be present in the answer of the following calculations?

(i) $\frac{0.02856 \times 298.15 \times 0.112}{0.5785}$ (ii) 5×5.364 (iii) $0.0125 + 0.7864 + 0.0215$

Ans. (i) Least precise number (0.112) has 3 significant figures.

So, Number of significant figures in the answer should be 3.

(ii) Number of significant figures in the answer

= Number of significant figures in 5.364 = 4

(iii) Since the least number of decimal places in each term is four, the number of significant figures in the answer should be 4.

32. Use the data given in the following table to calculate the molar mass of naturally occurring argon isotopes:

Isotope	Isotopic molar mass	Abundance
36 Ar	35.96755 $gmol^{-1}$	0.337%
38 Ar	37.96272 $gmol^{-1}$	0.063%
40 Ar	39.9624 $gmol^{-1}$	99.600%

Ans. Molar mass of argon

$$= 35.96755 \times 0.00337 + 37.96272 \times 0.00063 + 39.9624 \times 0.99600$$

$$= 39.948 \text{ gmol}^{-1}$$

33. Calculate the number of atoms in each of the following

(i) 52 moles of Ar (ii) 52 u of He (iii) 52 g of He.

Ans. (i) 1 mole of Ar = 6.022×10^{23} atoms of Ar

$\therefore$ 52 mol of Ar = $52 \times 6.022 \times 10^{23} = 3.131 \times 10^{25}$ atoms

(ii) 4 u of He = 1 atom of He

$$1 \text{ u of He} = \frac{1}{4} \text{ atom of He}$$

$$\therefore 52\text{u of He} = \frac{52}{4} = 13 \text{ atoms}$$

(iii) 4 g of He = 6.022×10^{23} atoms of He

$$\therefore 52 \text{ g of He} = \frac{6.022 \text{ x } 10^{23} \text{ x } 52}{4} = 7.8286 \times 10^{24} \text{ atoms}$$

34. A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at STP) of this welding gas is found to weigh 11.6 g. Calculate

(i) empirical formula, (ii) molar mass of the gas, and (iii) molecular formula.

Ans. (i) 1 mole (44 g) of CO_2 contains 12 g of carbon.

$$\therefore 3.38 \text{ g of } CO_2 \text{ will contain carbon} = \frac{12 \text{ g}}{44 \text{ g}} \text{ x } 3.38 \text{ g} = 0.9217 \text{ g}$$

18 g of water contains 2 g of hydrogen.

$$\therefore 0.690 \text{ g of water will contain hydrogen} = \frac{2 \text{ g}}{18 \text{ g}} \text{ x } 0.690 = 0.0767 \text{ g}$$

Since no other products are formed, the total mass of the compound

$$= 0.9217 \text{ g} + 0.0767 \text{ g} = 0.9984 \text{ g}$$

$\therefore$ Percent of C in the compound $= \dfrac{0.9217 \text{ g}}{0.9984 \text{ g}} \times 100 = 92.32\%$

Percent of H in the compound $= \dfrac{0.0767 \text{ g}}{0.9984 \text{ g}} \times 100 = 7.68\%$

Moles of carbon in the compound $= \dfrac{92.32}{12.00} = 7.69$

Moles of hydrogen in the compound $= \dfrac{7.68}{1} = 7.68$

$\therefore$ Ratio of carbon to hydrogen in the compound = 7.69 : 7.68 = 1: 1

Hence, the empirical formula of the gas is CH.

(ii) Given, weight of 10.0L of the gas (at S.T.P) = 11.6 g

$\therefore$ Weight of 22.4 L of gas at STP $= \dfrac{11.6 \text{ g}}{10.0 \text{ L}} \times 22.4 \text{ L} = 25.984 \text{ g} \approx 26 \text{ g}$

(iii) Empirical formula mass of CH = 12 + 1 = 13 g

$$\therefore n = \frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{26 \text{ g}}{13 \text{ g}} = 2$$

$\therefore$ Molecular formula of gas $= (CH)_n = C_2H_2$

35. Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction, $CaCO_{3(s)} + 2\,HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

What mass of $CaCO_3$ is required to react completely with 25 mL of 0.75 M HCl?

Ans. 0.75 M of HCl = 0.75 mol of HCl are present in 1 L of water

$$= 0.75 \times 36.5 = 27.375 \text{ g}$$

So, 1000 mL of solution contains 27.375 g of HCl.

$\therefore$ Amount of HCl present in 25 mL of solution $= \dfrac{27.375 \text{ g}}{1000 \text{ mL}} \times 25 \text{ mL} = 0.6844 \text{ g}$

From the given chemical equation,

2 mol of HCl ie. $2 \times 36.5 = 73$ g react with 1 mol of $CaCO_3$ ie,100 g.

$\therefore$ Amount of $CaCO_3$ that will react with 0.6844 g $= \dfrac{100}{73} \times 0.6844 \text{ g} = 0.968 \text{ g}$

36. Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid according to the reaction

$4HCl_{(aq)} + MnO_{2(s)} \rightarrow 2H_2O_{(l)} + MnCl_{2(aq)} + Cl_{2(g)}$

How many grams of HCl react with 5.0 g of manganese dioxide?

Ans. 1 mol MnO_2 ie., 55 + 2 × 16 = 87 g reacts completely with 4 mol of HCl.

ie., 4 × 36.5 = 146 g

So 5.0 g of MnO_2 will react with = $\frac{146\ g}{87\ g}$ x 5.0 g = 8.4 g of HCl

ADDITIONAL QUESTIONS AND ANSWERS

1. (a) Determine the number of moles present in 0.55 mg of electrons.

(i) 1 mole (ii) 2 moles (iii) 1.5 moles (iv) 0.5 moles

(b) Give the empirical formula of the following.

$C_6H_{12}O_6$, C_6H_6, CH_3COOH, $C_6H_6Cl_6$

(c) Two elements, carbon and hydrogen combine to form C_2H_6, C_2H_4 and C_2H_2. Identify the law illustrated here.

Ans. (a) 1 mole

[1 mole = 6.022 x 10^{23} electrons

mass of electron = 9.1 x 10^{-31} kg

$\therefore$ 1 mole = 6.023 x 10^{23} x 9.1 x 10^{-31} kg

= 54.8 x 10^{-8} kg

= 54.8 x 10^{-8} x 10^{6} mg

= 54.8 x 10^{-2} mg

= 0.548 mg = 0.55 mg]

1 kg = 1000 g
1 g = 1000 mg

(b) CH_2O, CH, CH_2O, CHCl

(c) Law of multiple proportions

2. Empirical formula represents the simplest whole number ratio of various atoms present in a compound.

(a) Give the relation between empirical formula and molecular formula.

(b) An organic compound has the following percentage composition C = 12.36%, H = 2.13%, Br = 85%. Its vapour density is 94. Find its molecular formula.

(c) What is mole fraction?

Ans. (a) An empirical formula represents the simplest whole number ratio of various atoms present in a compound whereas the molecular formula shows the exact number of different types of atoms present in a molecule of a compound.

M.F = n × Empirical formula

(b) Empirical Formula : CH_2Br

Molecular Formula : $C_2H_4Br_2$

(c) **Mole Fraction :** It is the ratio of number of moles of a particular component to the total number of moles of the solution. If a substance 'A' dissolves in substance 'B' and their number of moles are n_A and n_B respectively; then the mole fractions of A and B are given as

$$\text{Mole fraction of A} = \frac{\text{No. of moles of A}}{\text{No. of moles of solution}} = \frac{n_A}{n_A + n_B}$$

$$\text{Mole fraction of B} = \frac{\text{No. of moles of B}}{\text{No. of moles of solution}} = \frac{n_B}{n_A + n_B}$$

3. (a) When nitrogen and hydrogen combines to form ammonia, the ratio between the volumes of gaseous reactants and products is 1: 3: 2. Name the law of chemical combination illustrated here.

(b) A compound is made up of two elements A and B, has A = 70% and B = 30%. The relative number of moles of A and B in the compound are 1.25 and 1.88 respectively. If the molar mass of the compound is 160, find the molecular formula of the compound.

Ans. (a) Gay Lussac's law of gaseous volumes.

(b) The empirical formula of the compound is A_2B_3

$$\text{Atomic mass} = \frac{\text{Mass percentage}}{\text{Relative no. of moles}}$$

$$\text{Atomic mass of A} = \frac{70}{1.25} = 56$$

$$\text{Atomic mass of B} = \frac{30}{1.88} = 15.96 \approx 16$$

Empirical formula mass = (56 × 2) + (16×3) = 112 + 48 = 160

Molecualr mass = 160 (given)

Now, $n = \frac{160}{160} = 1$

Molecular formula = (Empirical formula) × n = $(A_2B_3) \times 1 = A_2B_3$

4. 12 g of ^{12}C contains Avogadro's number of carbon atoms.

(a) Give the Avogadro's number.

(b) The mass of 2 moles of ammonia gas is ....

(i) 2 g (ii) 1.2×10^{22}g (iii) 17 g (iv) 34g

(c) Calculate the volume of ammonia gas produced at STP when 140 g of nitrogen gas reacts with 30 g of hydrogen gas. (Atomic mass: N = 14u, H = 1 u)

Ans. (a) 6.022×10^{23}

(b) (iv) 34 g

(c) N_2 reacts with H_2 as per the equation

$N_2(g) \quad + \quad 3H_2(g) \rightleftharpoons 2NH_3(g)$

1 mol	3 mol	2 mol
28 g	6 g	34 g
22.4 mL	67.2 mL	44.8 mL(at STP)

Volume of $NH_3(g) = \dfrac{44.8L \times 140g}{28g} = 224$ L

5. 'A given compound always contains exactly the same proportion of elements by weight.'

(a) (i) Name the above law.

(ii) Write the name of the Scientist who proposed this law.

(b) Calculate the number of molecules in each of the following:

(i) 1 g N_2 (ii) 1 g CO_2 (Given that N_A is 6.022 x 10^{23}, molecular mass of N_2 = 28 and CO_2 = 44)

Ans. (a) (i) Law of definite proportions/ Law of definite composition

(ii) Joseph Proust

(b) 1 mole = 6.02 x 10^{23} particles

1 mole = molar mass

(i) 1 mole of N_2 = 6.02 x 10^{23} molecules of N_2

28g of N_2 = 6.02 x 10^{23} molecules of N_2

1g of $N_2 = \dfrac{6.02 \times 10^{23}}{28}$ molecules of N_2

= 0.215 x 10^{23}

= 2.15 x 10^{23} molecules of N_2

(ii) 1 mole of CO_2 = 6.023 x 10^{23} molecules of CO_2

44g of CO_2 = 6.023 x 10^{23} molecules of CO_2

1g of $CO_2 = \dfrac{6.023 \times 10^{23}}{44}$ molecules of CO_2

= 0.1369 x 10^{23} molecules of CO_2

= 1.369 x 10^{22} molecules of CO_2

6. Hydrogen combines with oxygen to form two different compounds, namely water (H_2O) and hydrogenperoxide (H_2O_2).

(a) Which law is obeyed by this combination?

(b) State the law.

(c) How many significant figures are present in the following?

(i) 0.0025 (ii) 285

Ans. (a) Law of Multiple Proportions

(b) According to this law, if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers.

(c) (i) 0.0025 - 2 significant figures

(ii) 285 - 3 significant figures

7. (a) How many moles of dioxygen are present in 64g of dioxygen? (Molar mass of dioxygen is 32).

(b) The following data were obtained when dinitrogen (N_2) and dioxygen (O_2) react together to form different compounds.

Mass of N_2	Mass of O_2
14 g	16 g
14 g	32 g
28 g	32 g
28 g	80 g

Name the law of chemical combination obeyed by the above experimental data.

(c) Define empirical formula. How is it related to the molecular formula of a compound?

Ans. (a) Number of moles = $\frac{\text{Given mass}}{\text{Molar mass}} = \frac{64\text{ g}}{32\text{ g mol}^{-1}}$ = 2 mol

(b) Law of multiple proportions.

(c) An empirical formula represents the simplest whole number ratio of various atoms present in a compound whereas the molecular formula shows the exact number of different types of atoms present in a molecule of a compound.

M.F = n × Empirical formula

8. (a) Atoms have very small mass and so usually the mass of atoms are given relative to a standard called atomic mass unit. What is atomic mass unit (amu)?

(b) In a reaction $A + B_2 \rightarrow AB_2$, identify the limiting reagent in the reaction mixture containing 5mol A and 2.5mol B.

(c) Calculate the mass of NaOH required to make 500 ml of 0.5M aqueous solution. (Molar mass of NaOH = 40)

Ans. (a) One atomic mass unit is defined as a mass exactly equal to onetwelfth the mass of one carbon - 12 atom.

(b) Here 1 mole of A reacts completely with 1 mole of B_2 to form 1 mole of AB_2. Thus, 5 mole of A requires 5 mole of B_2. So B_2 is the limiting reagent.

(c) Mass of NaOH = $\frac{\text{Molarity x Molar mass of NaOH x Vol. of NaOH}}{1000}$

$$= \frac{0.5 \times 40 \times 500}{1000} = 10 \text{ g}$$

9. The mole concept helps in handling a large number of atoms and molecules in stoichiometric calculations.

(a) Define 1 mol.

(b) What is the number of hydrogen atoms in 1 mole of methane (CH_4)?

(c) Calculate the amount of carbon dioxide formed by the complete combustion of 80g of methane as per the reaction:

$CH_4\ (g) + 2O_2\ (g) \longrightarrow CO_2\ (g) + 2\ H_2O\ (g)$

(Atomic mass of C = 12.01u, H = 1.008u, O = 16u)

Ans. (a) One mole is the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 g (or 0.012 kg) of the ^{12}C isotope.

(b) No. of hydrogen atoms in 1 mole of CH_4 = No. of mole × N_A × n

$= 1 \times 6.023 \times 10^{23} \times 4$

$= 24.092 \times 10^{23}$ atoms

(c) $CH_{4(g)} + 2O_2 \rightarrow CO_2 + 2H_2O$

According to the equation, 16g of CH_4 gives 44 g of CO_2

$\Rightarrow$ 1 g of CH_4 gives $\frac{44}{16}$ g of CO_2

$\therefore$ 80 g of CH_4 gives $\frac{44}{16} \times 80 = 220$g of CO_2

10. (a) Mole is a very large number to indicate the number of atoms, molecules etc. Write another name for one mole.

(b) (i) How the molecular formula is different from that of empirical formula?

(ii) An organic compound on analysis gave the following composition.
Carbon = 40%, Hydrogen = 6.66% and oxygen = 53.34%.

Calculate its molecular formula if its molecular mass is 90.

Ans. (a) gram atomic mass

(b) (i) An empirical formula represents the simplest whole number ratio of various atoms present in a compound whereas the molecular formula shows the exact number of different types of atoms present in a molecule of a compound.

(ii)

Element	%	At. mass	%/At. mass	simplest at.ratio	simplest whole number ratio
C	40	12	$\frac{40}{12} = 3.33$	$\frac{3.33}{3.33} = 1$	1

H	6.66	1	$\frac{6.66}{1} = 6.66$	$\frac{6.66}{3.33} = 2$	2
O	53.34	16	$\frac{63.34}{16} = 3.33$	$\frac{3.33}{3.33} = 1$	1

$\therefore$ Empirical formula = CH_2O

E.F mass = 12 + 2 +16 = 30

$$\text{Now, } n = \frac{\text{molecular mass}}{\text{E.F mass}} = \frac{90}{30} = 3$$

$\therefore$ M.F = n × Empirical formula = 3 × (CH_2O) = $C_3H_6O_3$

11. The combination of elements to form compounds is governed by the laws of chemical combination.

a. Hydrogen combines with oxygen to form compounds, namely water and hydrogen peroxide. State and illustrate the related law of chemical combination.

b. What is mean by limiting reagent in a chemical reaction?

c. 28 g of nitrogen is mixed with 12 g of hydrogen to form ammonia as per the reaction, $N_2 + 3\,H_2 \longrightarrow 2NH_3$. Which is the limiting reagent in this reaction?

Ans. (a) **Law of Multiple Proportions :** This law was proposed by Dalton in 1803. According to this law, if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers.

For example, hydrogen combines with oxygen to form two compounds, namely, water and hydrogen peroxide.

Hydrogen + Oxygen → Water

2g 16g 18g

Hydrogen + Oxygen → Hydrogen Peroxide

2g 32g 34g

Here, the masses of oxygen (i.e. 16 g and 32 g) which combine with a fixed mass of hydrogen (2g) bear a simple ratio, i.e. 16:32 or 1: 2.

(b) The reagent which limits a reaction or the reagent which is completely consumed in a chemical reaction is called limiting reagent or limiting reactant.

(c) Since N_2 is completely used up in a reaction, N_2 is the limiting reagent.

12. The laws of chemical combination govern the formation of compounds from elements.

(a) State the law of conservation of mass. Who put forward this law? **(1½)**

(b) The following data are obtained when dinitrogen and dioxygen react together to form different compounds.

Sl. No.	Mass of dinitrogen (in g)	Mass of dioxygen (in g)
1	14	16
2	14	32
3	28	48
4	28	80

Which law of chemical combination is illustrated by the above experimental data? Explain?

Ans. (a) It states that matter can neither be created nor destroyed. This law was put forth by Antoine Lavoisier in 1789.

(b) Law of multiple proportions proposed by John Dalton.

The oxides of nitrogen :

NO	NO_2	N_2O_3	and	N_2O_5
14g:16g	14g:32g	28g:48g		28g:80g

The different masses of oxygen which combine with a fixed mass (28g) of nitrogen are in the ratio 32:64:48:80 = 2:4:3:5, which is a simple whole number ratio. Hence, the law is verified.

13. The laws of chemical combination are the basis of the atomic theory.

(a) Name the law of chemical combination illustrated by the pair of compounds, CO and CO_2.

(b) State and explain the law of conservation of mass.

(c) Calculate the molarity of a solution containing 8 g of NaOH in 500 mL of water.

Ans. (a) Law of multiple proportion.

(b) **Law of Conservation of Mass :** It states that matter can neither be created nor destroyed. This law was put forth by Antoine Lavoisier in 1789.

(c) $$\text{Molarity} = \frac{\text{weight of the solute in grams}}{\text{Molar wt of solute} \times \text{volume of solution in litre}}$$

$$= \frac{8}{40 \times 0.5} = \frac{1}{2.5} = 0.4\ \text{M}$$

14. One mole is the amount of substance that contains as many particles as 12 g of C^{12} isotope of carbon.

(a) What do you mean by molar mass of a compound?

(b) Calculate the number of moles in 1 L of water (Density of water 1 g/mL). Also calculate the number of water molecules in 1 L water.

Ans. (a) The mass of one mole of a substance in grams is called its molar mass.

(b) $$\text{Number of moles} = \frac{\text{Mass of the substance}}{\text{Gram molecular mass}}$$

$$\text{Number of moles in 1 litre of water} = \frac{1000\ \text{gL}^{-1}}{18\ \text{g mol}^{-1}} = 55.55\ \text{mol L}^{-1}$$

15. If the mass percent of various elements of a compound is known, its empirical formula can be calculated.
 (a) What is mass percent?

(b) A compound contains 4.07% hydrogen, 24.27% carbon and 71.65% chlorine. Its molecular mass is 98.96. What are the empirical and molecular formulae?

Ans. (a) Mass % of an element = $\frac{\text{mass of that element in the compound x 100}}{\text{molar mass of the compound}}$

(b)

Element	%	Relative no. of atoms	Dividing by smallest factor	Rate of no. atoms
C	24.27%	$\frac{24.27}{12} = 2.02$	$\frac{2.02}{2.01} = 1.00$	1
H	4.07%	$\frac{4.07}{1} = 4.07$	$\frac{4.07}{2.01} = 2.02$	2
Cl	71.65%	$\frac{71.65}{35.5} = 2.018$	$\frac{2.01}{2.01} = 1$	1

Empirical formula = CH_2Cl

Empirical formula mass = 12×1 + 1×2 + 1×35.5

= 12 + 2 + 35.5 = 49.5

Molecular mass = 98.96

$$n = \frac{98.96}{49.5} = 2$$

Molecular formula of the compound = $n(CH_2Cl) = C_2H_4Cl_2$

16. (a) When 10g of sulphur is burnt in 10g of oxygen 20g of sulphur dioxide is produced? Find the mass of sulphur dioxide formed on burning 20g of sulphur in 30g of oxygen ? Justify your answer by stating the law which governs your answer.

(b) State the postulate of Dalton's atomic theory which explains the above law.

Ans. (a) According to the question sulphur and oxygen combine in the ratio of 1:1 by mass. 20 g sulphur will combine only with 20 g of oxygen and 40 g of sulphur dioxide will be produced. This is in accordance with law of constant proportion.

Which says in a chemical substance elements are always present in definite proportion by mass.

(b) The relative number and kinds of atoms are constant in a given compound.

17. Calculate the number of aluminum ions present in 0.051 g of aluminium oxide [Atomic Mass of Al = 27 u]

Ans. 1 Mole of Al_2O_3 = 27 x 2 + 16 x 3

= 54 + 48 = 102 grams

Now 102 g of Al_2O_3 contains = 54 g of Al

So, 0.051 g of Al_2O_3 contains = $\frac{54}{102}$ x 0.051 g of Al = 0.027 g of Al

Now 27 g of Al has ions = $6.022\ 10^{23}$

So 0.027 g of Al has ions = $\frac{6.0223\ 10^{23}}{27}$ x 0.027 = 6.022×10^{20}

18. (i) State the law of constant proportions.

(ii) Show that water illustrates the law of constant proportions.

Ans. (i) In a chemical substance, the elements are always present in definite proportion by mass.

(ii) In pure water, the ratio between the masses of hydrogen and oxygen is 2 : 16 or 1 : 8, which will remain the same, whatever the source of water.

19. When 3.0 g of magnesium is burnt in 2.00 g of oxygen, 5.00 g of magnesium oxide is produced. What mass of magnesium oxide will be formed when 3.00 g magnesium is burnt in 5.00 g of oxygen? Which law of chemical combination will govern your answer? State the law.

Ans. When 3.0 g of magnesium is burnt in 2.00 g of oxygen, 5.00 g of magnesium oxide is produced. It means that magnesium and oxygen are combined in the ratio 3:2 to form magnesium oxide.So when 3.00 g magnesium is burnt in 5.00 g of oxygen, 5.00 g magnesium oxide will be formed and remaining oxygen will not be used up. Law of definite proportion. It states that in a chemical substance, the elements are always present in definite proportions by mass.

20. Verify by calculating that :

(a) 5 moles of CO_2 and 5 moles of H_2O do not have the same mass. Atomic mass of Carbon, Oxygen and Hydrogen are 12u, 16u, 14u respectively

(b) 240 g of calcium and 240 g magnesium elements have a mole number ratio of 3 : 5. Atomic mass of Calcium and magnesium are 40u and 24u respectively.

Ans. (a) Mass of 5 moles of CO_2 = 5 (Molar mass of C + 2 x Molar mass of O)

= 5(12 + 2 x 16) = 5(12 + 32)

= 5(44) = 220 g

Mass of 5 moles of H_2O = 5(2 x Molar mass of H + Molar mass of O)

= 5(2 x 1 x 16) = 5(2 x 16) = 5(18) = 90 g

$\Rightarrow$ 5 moles of CO_2 and 5 moles of H_2O have different mass.

(b) 240 g of Ca contain = 240/40 moles = 6 moles

240 g of Mg contain = 240/24 moles = 10 moles

ratio = $\frac{6}{10}$ = 3 : 5

21. Which has more number of atoms 100g of sodium or 100g of iron (At mass Na = 23 u,Fe = 56u)

Ans.

For Sodium	**For Iron**
m = 100 g	m = 100g

M	=	23 g	M	=	56 g
No	=	6.022×10^{23}	No	=	6.022×10^{23}
n	=	$\frac{m}{M} = \frac{100}{23}$	n	=	$\frac{m}{M} = \frac{100}{56}$
N	=	n x No	N	=	n x No
	=	$\frac{100}{23} \times 6.022 \times 10^{23}$		=	$\frac{100}{56} \times 6.022 \times 10^{23}$
	=	$4.348 \times 6.022 \times 10^{23}$		=	$1.785 \times 6.022 \times 10^{23}$

Hence 100 g sodium has more number of atoms.

22. The percentages of three elements calcium, carbon and oxygen in a sample of calcium carbonate are given as:

Calcium = 40%, Carbon = 12.0%, Oxygen = 48%

If the law of constant proportion is true, what weights of these elements will be present in 1.5 g of another sample of calcium carbonate?

(Atomic masses of Ca = 40 u, C = 12 u, O = 16 u)

Ans. Molecular mass of $CaCO_3 = 40 + 12 + 16 \times 3 = 100$ u

weight of calcium present in 1.5 g of calcium carbonate $= 1.5 \text{ g} \times \frac{40}{100} = 0.6\text{g}$

weight of carbon present in 1.5 g of calcium carbonate $= 1.5 \text{ g} \times \frac{12}{100} = 0.18\text{g}$

weight of oxygen present in 1.5 g of calcium carbonate $= 1.5 \text{ g} \times \frac{48}{100} = 0.72\text{g}$

23. (a) State any four postulates of Dalton's atomic theory of matter.

(b) Which of his postulates does not hold correct at present?

Ans. (a) 1. All matter is made of very tiny particles called atoms.

2. Atoms are indivisible particles which can neither be created nor destroyed in a chemical reaction.

3. Atoms of a given element are identical in mass and chemical properties.

4. Atoms combine in the ratio of small whole numbers to form compounds.

(b) According to him, atoms are indivisible. But atoms can be further divided into electrons, protons and neutrons.

24. (a) Calculate the mass of 0.5 mole of oxygen atoms.

(b) Calculate the number of molecules of glucose present in its 90 grams (molecular mass of glucose is 180u)

(c) Calculate number of moles of water in 2 grams of water.

Ans. (a) Mass = Molar mass x no.of moles

$m = M \times n = 16 \times 0.5 = 8\text{g}.$

(b) $\text{n (moles)} = \frac{\text{m (mass)}}{\text{M (molar mass)}}$

$$n = \frac{90}{180} = 0.5 \text{ mole.}$$

$\because$ 1 mole contains = 6.022×10^{23} molecules

$\because$ 0.5 moles will contain = $0.5 \times 6.022 \times 10^{23}$

= 3.011×10^{23} molecules

(c) Mole (n) = $\frac{\text{m (mass)}}{\text{M (molar mass)}}$

= $\frac{2}{18} = \frac{1}{9}$ = 0.11 moles

25. " Equal volumes of gases at the same temperature and pressure should contain equal number of molecules".

(a) Name the above law.

(b) The amount of sodium hydroxide required to prepare 500 mL of 0.1 M sodium hydroxide solution is

[(a) 4g (b) 2g (c) 1g (d) 5g]

(c) An organic compound contains 80%C and 20%H. Determine its empirical formula.

Ans. (a) Avogadros Law (b) 2 g (c) CH_3

26. When a solution is diluted its concentration changes.

If 500mL of a 5M solution is diluted to 1500mL, What will be the molarity of the solution obtained?

(i) 1.5 M (ii) 1.66 M (iii) 0.017 M (iv) 1.59 M

Ans. (ii)

27. How are 0.5 ml of NaOH differents from 0.5 M of NaOH?

Ans. 0.5 ml of NaOH means 0.5 mole ie. 20.0 g of NaOH, 0.5 M of NaOH means that 0.5 mole ie. 20.0g of NaOH are dissolved in 1L of its solution.

28. A compound on analysis was found to contain C = 34.6%, H = 3.85% and O = 61.55% Calculate the empirical formula.

Ans.

Element	Percentage	Atomic Mass	Gram atoms (Moles)	Atomic ratio (Molar ratio)	Simplest whole no. ratio
C	34.6	12	$\frac{34.6}{12} = 2.88$	$\frac{2.88}{2.88} = 1$	3
H	3.85	1	$\frac{3.85}{1} = 3.85$	$\frac{3.85}{2.88} = 1.337$ or $\frac{4}{3}$	4
O	61.55	16	$\frac{61.55}{16} = 3.85$	$\frac{3.85}{2.88} = 1.337$ or $\frac{4}{3}$	4

So the simplest ratio are C : H : O = 3 : 4 : 4

So the empirical formula is $C_3H_4O_4$

29. What is the relation between vapour density and molecular mass?

Ans. Vapour density = $\frac{\text{molecular mass}}{2}$

30. Calculate the molarity of a solution of NaOH containing 10g of NaOH in 200 cm^3 solution. (Mol. mass of NaOH = 40).

Ans. Molarity of NaOH = $\frac{10 \times 1000}{40 \times 200}$ = 1.25 M.

31. How many electrons are present in 224 mL of CO_2 at STP ?

Ans. No. of moles = $\frac{224}{22400}$ = 0.01

No. of melecules = 0.01 x N_A = 0.01 x 6.02 x 10^{23} = 6.02 x 10^{21}

No of electrons in one molecule = 6 + 8 × 2 = 22

$\therefore$ Total number of electrons = 22 x 6.02 x 10^{21} = 1.324 x 10^{23}

32. Calculate the molality of a solution containing 10 g of NaOH in 200 cm^3 of solution. Density of solution = 1.04 g per mL. (Molecular mass of NaOH = 40).

Ans. Mass of solution = 200 x 1.04 = 208 g.

Mass of NaOH = 10 g

$\therefore$ Mass of water = 208 - 10 = 198 g

Molality = $\frac{10 \times 1000}{40 \times 198}$ = 1.262 M.

33. Calculate the number of moles present in 200 g $CaCO_3$?

Ans. No. of moles (n) = $\frac{\text{Given mass } (m)}{\text{Molar mass}} = \frac{200}{100}$ = 2

34. What is the significant figures in 1.050 x 10^4?

Ans. Four

35. What is the law called which deals with the ratios of the volumes of the gaseous reactants and products?

Ans. Gay Lussac's law of gaseous volumes.

36. Copper oxide obtained by heating copper carbonate or copper nitrate contains copper and oxygen in the same ration by mass. Which law is illustrated by this observation? State the law.

Ans. Law of Definite Proportions This law states that: A chemical compound always consists of the same elements combined together in the same ratio, irrespective of the method of preparation or the source from where it is taken.

37. Write the empirical formula of the following:

(a) N_2O_4 (b) $C_6H_{12}O_6$ (c) H_2O (d) H_2O_2

Ans. (a) NO_2 (b) CH_2O (c) H2O (d) HO

38. State the number of significant figures in each of the following:

(i) 208.91 (ii) 0.00456 (iii) 453 (iv) 0.346

Ans. (i) 208.91 has five significant figures.

(ii) 0.00456 has three significant figures.

(iii) 453 has three significant figures.

(iv) 0.346 has three significant figures.

39. What is the percentage of carbon, hydrogen and oxygen in ethanol?

Ans. Molecular formula of ethanol is : C_2H_5OH

Molar mass of ethanol is : (2×12.01 + 6×1.008 + 16.00) g = 46.068 g

Mass per cent of carbon $= \frac{24.02\ g}{46.068\ g} \times 100 = 52.14\%$

Mass per cent of hydrogen $= \frac{6.048\ g}{46.068\ g} \times 100 = 13.13\%$

Mass per cent of oxygen $= \frac{16.00\ g}{46.068\ g} \times 100 = 34.73\%$

ENTRANCE CORNER

1. 1.0g of magnesium is burnt with 0.56 g of oxygen in a closed vessel. Which reactant is left in excess and how much?

(At.weight of mg= 24, O=16)

(a) Mg, 0.16g (b) O_2, 0.16 g

(c) Mg, 0.44 g (d) O_2, 0.28 g **(2014)**

2. 6.02 x10^{20} molecules of urea are present in 100mL of its solution. The concentration of solution is

(a) 0.02M (b) 0.01M

(c) 0.001M (d) 0.1M **(2013)**

3. What volume of oxygen gas (O_2) measured at 0^0C and 1 atm, is needed to burn completely 1L of propane gas ($C_3\ H_6$) measured under the same conditions?

(a) 7L (b) 6L

(c) 5L (d)10L **(2008)**

4. The number of moles of $KMnO_4$ reduced by one mole of KI in alkaline medium is

(a) One fifth (b) five

(c) one (d) two **(2005)**

5. Which has maximum number of molecules?

(a) $7gN_2$ (b) $2g\ H_2$

(c) $16\ g\ NO_2$ (d) $16\ gO_2$ **(2002)**

6. Assuming fully decomposed, the volume of CO_2 released at STP on heating 9.85 g of $BaCO_3$ (at mass of Ba = 137) will be

(a) 1.12L (b) 0.84L

(c) 2.24L (d) 4.96L **(2000)**

7. In the reaction,

$4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(l)$

When 1 mole of ammonia and 1 mole of O_2 are made to react to completion, then

(a) 1.0 mole of H_2O is produced (b) 1.0mole of NO will be produced

(c) all the oxygen will be consumed (d) all the ammonia will be consumed **(1998)**

8. The percentage weight of Zn in white vitriol ($ZnSO_4 . 7H_2O$) is approximately equal to (at. Mass of Zn = 65, S= 32, O =16 and H =1)

(a) 33.65% (b) 32.56%

(c) 23.65% (d) 22.65% **(1995)**

9. An organic compound contains C,H and S. The minimum molecular weight of the compound containing 8% sulphur is (atomic weight of S= 32amu)

(a) 600 g mol^{-1} (b) 200 g mol^{-1}

(c) 400 g mol^{-1} (d) 300 g mol^{-1} **(2016)**

10. The volume of 0.1N dibasic acid sufficient to neutralize 1g of a base that furnishes 0.04 mole of OH^- in aqueous solution is

(a) 400 mL (b) 600 mL

(c) 200 mL (d) 800 mL **(2016)**

11. The ratio of masses of oxygen and nitrogen in a particular gaseous mixture is 1:4. The ratio of number of their molecules is

(a) 3 : 16 (b) 1 : 4

(c) 7 : 32 (d) 1 : 8 **(2014)**

12. The molarity of a solution obtained by mixing 750mL of 0.5M HCl with 250mL of 2M HCl will be

(a) 0.975 M (b) 0.875 M

(c) 1.00 M (d) 1.75 M **(2013)**

13. How many moles of magnesium phosphate, $Mg_3(PO_4)_2$ will contain 0.25 mole of oxygen atoms?

(a) 0.02 (b) 3.125×10^{-2}

(c) 1.25×10^{-2} (d) 2.5×10^{-2} **(2006)**

14. With increase of temperature, which of these changes?
(a) Molality
(b) Weight fraction of solute
(c) Fraction of solute present in water
(d) Mole fraction **(2002)**

15. The number of oxygen atoms in 4,4 g of CO_2 is
(a) 1.2×10^{23} (b) 6×10^{22}
(c) 6×10^{23} (d) 12×10^{23} **(1990)**

16. What is the weight of oxygen required for the complete combustion of 2.8 kg of ethylene?
(a) 2.8kg (b)6.4kg
(c) 9.6 kg (d)96 kg **(1989)**

17. How many moles of electrons weigh one kilogram?
(a) 6.023×10^{23} (b) $\frac{1}{9.108}\times10^{31}$
(c) $\frac{6.023}{9.108}\times10^{54}$ (d) $\frac{1}{9.108\times6.023}\times10^{8}$ (e) 0.55

18. Which among the following is the heaviest?
(a) One mole of oxygen (b) One molecule of sulphur trioxide
(c) 100 amu of uranium (d) 44g of CO_2

19. The mass of nitrogen per gram hydrogen in the compound Hydrazine is exactly one and a half times the mass of Nitrogen in the compound Ammonia. The fact illustrates the law of:
(a) conservation of mass
(b) multiple proportions
(c) definite proportions
(d) reciprocal proportions
(e) Constant composition.

20. The mol. formula of chloride of a metal is MCl_3. The formula of the metal sulphate is:
(a) MSO_4 (b) M_2SO_4 (c) $M_2(SO_4)_3$
(d) $M_3(SO_4)_2$ (e) $M(SO_4)_2$

21. The molarity of pure water is:
(a) 1 (b) 10 (c) 7 (d) 55.56 (e) 100

22. The density of 3M solution of NaCl is 1.25g mL^{-1}. The molarity of the solution is ....m

(a) 27.9 (b) 2.79 (c) 1.79 (d) 0.79 (e) 17.9

Answers

1 (a)	2 (b)	3 (c)	4 (c)	5 (b)	6 (a)
7 (c)	8 (d)	9 (c)	10 (a)	11 (c)	12 (b)
13 (b)	14 (c)	15 (a)	16 (c)	17 (d)	18 (d)
19 (b)	20 (c)	21 (d)	22 (b)		

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