NCERT TEXTUAL EXERCISE (SOLVED)

1. Calculate the molecular mass of the following :

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

Ans. (i) Molecular mass of $H_2O = 2\,(1.008\text{ amu}) + 16.00\text{ amu} = 18.016\text{ amu}$.

(ii) Molecular mass of $CO_2 = 12.01\text{ amu} + 2 \times 16.00\text{ amu} = 44.01\text{ amu}$.

(iii) Molecular mass of $CH_4 = 12.01\text{ amu} + 4\,(1.008\text{ amu}) = 16.042\text{ amu}$.

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

Ans. Mass % of an element $= \dfrac{\text{Mass of that element in the compound}}{\text{Molar mass of the compound}} \times 100$

Now, molar mass of

$Na_2SO_4 = 2\,(23.0) + 32.0 + 4 \times 16.0 = 142\text{ g/mol}$

Mass percent of sodium $= \dfrac{46}{142} \times 100 = 32.39\%$

Mass percent of sulphur $= \dfrac{32}{142} \times 100 = 22.54\%$

Mass percent of oxygen $= \dfrac{64}{142} \times 100 = 45.07\%$

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

Ans.

Element	*Symbol*	*% by mass*	*Atomic Mass*	*Moles of the Elements*	*Simplest Molar Ratio*	*Simplest Whole Number Molar Ratio*
Iron	Fe	69.9	55.85	$\dfrac{69.9}{55.85} = 1.25$	$\dfrac{1.25}{1.25} = 1$	2
Oxygen	O	30.1	16.0	$\dfrac{30.1}{16.0} = 1.88$	$\dfrac{1.88}{1.25} = 1.5$	3

$\therefore$ Empirical formula $= Fe_2O_3$.

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

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

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

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

Ans. The balanced equation for the combustion of carbon in dioxygen/air is

$$\begin{array}{lll} C(s) & + \quad O_2(g) \longrightarrow & CO_2(g) \\ 1 \text{ mol} & \quad 1 \text{ mol} & 1 \text{ mol} \\ & \quad (32 \text{ g}) & (44 \text{ g}) \end{array}$$

(i) In air, combustion is complete. Therefore, CO_2 produced from the combustion of 1 mol of carbon = 44 g.

(ii) As only 16 g of dioxygen is available, it can combine only with 0.5 mol of carbon, i.e., dioxygen is the limiting reactant. Hence, CO_2 produced = 22 g.

(iii) Here again, dioxygen is the limiting reactant. 16 g of dioxygen can combine only with 0.5 mol of carbon. CO_2 produced again is equal to 22 g.

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.

Ans. 0.375 M aqueous solution means that 1000 mL of solution contains sodium acetate = 0.375 mol.

500 mL of the solution should contain sodium acetate = 0.375/2 mole

Molar mass of sodium acetate = 82.0245 g/mol

Mass of sodium acetate required = 0.375/2 mol × 82.0245 g/mol = 15.380 g.

6. Calculate the concentration of nitric acid in moles per litre in a sample which has a density, 1.41 g/mL and the mass per cent of nitric acid in it being 69%.

Ans. Mass percent of 69% means that 100 g of nitric acid solution contain 69 g of nitric acid by mass. Molar mass of nitric acid (HNO_3) = 1 + 14 + 48 = 63 g/mol

Moles in 69 g of HNO_3 = 69/63 = 1 .095 mol

Volume of 100 g nitric acid solution = 100/1.41 = 70.92 mL = 0.07092 L

Conc. of HNO_3 in moles per litre = 1.095/0.0792 L = 15.44 M

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

Ans. Mole of $CuSO_4$ contains 1 mol of Cu

Molar mass of $CuSO_4$ = 63.5 + 32 + 4 × 16 = 159.5 g/mol.

Thus, Cu that can be obtained from 159.5 g of $CuSO_4$ = 63.5 g

Cu that can be obtained from 100 g of $CuSO_4 = \dfrac{63.5}{159.5} \times 100 \text{ g} = 39.81 \text{ g}$

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.

Ans. Calculation of Empirical Formula. See Q.3.

Empirical formula mass of $Fe_2O_3 = 2 \times 55.85 + 3 \times 16.00 = 159.7$ g/mol

$$n = \frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{159.8}{159.5} = 1$$

Hence, molecular formula is same as empirical formula, viz., 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. Fractional abundance of $^{35}C1 = 0.7577$, Molar mass = 34.9689

Fractional abundance of $^{37}C1 = 0.2423$, Molar mass = 36.9659

Average atomic mass = (0.7577) (34.9689 amu) + (0.2423) (36.9659 amu)

= 26.4959 + 8.9568 = 35.4527

10. In 3 mol 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 mol of C_2H_6 contains 2 mol of carbon atoms. Hence, 3 mol of C_2H_6 will contain C-atoms = 6 mol.

(ii) 1 mol of C_2H_6 contains 6 mol of hydrogen atoms. Hence, 3 mol of C_2H_6 will contain H – atoms = 18 mol.

(iii) 1 mol of C_2H_6 contains 6.02×10^{23} molecules. Hence, 3 mol of C_2H_6 will contain ethane molecules = $3 \times 6.02 \times 10^{23} = 18.06 \times 10^{23}$ molecules.

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

Ans. Molar mass of sugar ($C_{12}H_{22}O_{11}$) = $12 \times 12 + 22 \times 1 + 11 \times 16 = 342$ g/mol

No. of moles in 20 g of sugar $= \dfrac{20\text{ g}}{342\text{ g / mol}} = 0.0585\text{ mol}$

Molar concentration $= \dfrac{\text{Moles of solute}}{\text{Volume of solution}} = \dfrac{0.0585}{2} = 0.0293$ M

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12. If the density of methanol is 0.793 kg/L, what is its volume needed for making 2.5 L of its 0.25 M solution?

Ans. Molar mass of methanol (CH_3OH) = 32 g/mol = 0.032 kg/mol

$$\text{As Molarity} = \frac{\text{mass of solute}}{\text{molar mass of solute} \times \text{volume of solution}}$$

We can write $\text{molarity} = \dfrac{\text{density of solution}}{\text{Molar mass of solute}}$

Molarity of the given solution will be

$$\frac{0.793\,\text{Kg/L}}{0.032\,\text{Kg}} = 24.78\ \text{mol/L}$$

Applying $\underset{\text{(Given solution)}}{M_1} \times V_1 = \underset{\text{(Solution to be prepared)}}{M_2} \times V_2$

$24.78 \times V_1 = 0.25 \times 2.5$ L or $V_1 = 0.02522$ 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 :

1 Pa = 1 N/m²

If mass of air at sea level is 1034 g/cm², calculate the pressure in pascal.

Ans. Pressure = Force/Area

Force = $m \times g$ = 1034 g × 9.8 m/s²

= 1.034 kg × 9.8 m/s²

= 10.1332 N

$$\text{Area} = 1\ \text{cm}^2 = \frac{1}{100}\,m \times \frac{1}{100}\,m = 0.0001\ \text{m}^2$$

$$\therefore\ P = \frac{10.1332\ \text{N}}{0.0001\ \text{m}^2} = 1.01332 \times 10^5\ \text{Pa}$$

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

Ans. SI unit of mass is kilogram (kg). It is defined as the mass of platinum–iridium cylinder that is stored in air-tight jar at International Bureau of Weights and Measures in France.

15. Match the following prefixes with their multiples:

Prefixes	**Multiples**
(i) micro	**10^6**
(ii) deca	**10^9**

(iii)	**mega**	$\mathbf{10^{-6}}$
(iv)	**giga**	$\mathbf{10^{-15}}$
(v)	**femto**	$\mathbf{10^{1}}$

Ans. Micro = 10^{-6}, deca = 10^1, mega = 10^6, giga = 10^9, femto = 10^{-15}.

16. What do you mean by significant figures?

Ans. The total number of digits in a number including the last digit whose value is uncertain is called significant figures.

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 per cent by mass.

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

Ans. (i) 15 ppm means 15 parts in million (10^6) parts,. % by mass $\frac{15}{10^6} \times 100$
$= 15\times10^{-4} = 1.5 \times 10^{-3}$ %

(ii) Molar mass of chloroform ($CHC1_3$) = 12 + 1 + 3 × 35.5 = 119.5 g/mol

100 g of the sample contain chloroform = 1.5×10^{-3} g

∴ 1000 g (1 kg) of the sample will contain chloroform = 1.5×10^{-2} g

Now molality $= \frac{1.5\times10^{-2}}{119.5} = 1.255\times10^{-4}$ m

∴ Molality = 1.255×10^{-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) 4.8×10^{-3} (ii) 2.34×10^5

(iii) 8.008×10^3 (iv) 5.000×10^2

(v) 6.0012×10^0

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) 6

(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**
14 g	**16 g**
14 g	**32 g**
28 g	**32 g**
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 =....................dm³

Ans. (a) Fixing the mass of dinitrogen as 28 g, masses of dioxygen combined will be 32, 64, 32 and 80 g in the given four oxides. These are in the ratio 2:4:2:5 which is a simple whole number ratio.

Hence, the given data obey the law of multiple proportions.

Statement – When two elements combine to form two or more chemical compounds, then the masses of one of the elements which combine with a fixed mass of the other, bear a simple ratio to one another.

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

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

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

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

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

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

22. If the speed of light is 3.0 × 10^8 m/s, calculate the distance covered by light in 2.00 ns.

Ans. Distance covered = Speed × Time = 3.0×10^8 m/s × 2.00 ns

$1\text{ns} = 10^{-9}$ s

$\therefore\ 3 \times 10^8\text{ m/s} \times 2 \times 10^{-9}\text{ s}$

= 0.6 m

23. In a reaction $A + B_2 \longrightarrow AB_2$

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. Therefore, 200 molecules of B will react with 200 atoms of A and 100 atoms of A will be left unreacted. Hence, B is the limiting reagent while A is the excess reagent.

(ii) According to the given reaction, 1 mol of A reacts with 1 mol of B. Therefore, 2 mol of A will react with 2 mol of B. Hence, A is the limiting reactant.

(iii) No limiting reagent.

(iv) 2.5 mol of B will react with 2.5 mol of A. Hence, B is the limiting reagent.

(v) 2.5 mol of A will react with 2.5 mol of B . Hence, A is the limiting reagent.

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

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

(i) Calculate the mass of ammonia produced if 2.00 × 10^3 g dinitrogen reacts with 1.00 × l0^3 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) 1 mol of N_2, i.e., 28 g react with 3 mol of H_2, i.e., 6 g of H_2, 2000 g of N_2 will react with $H_2 = \frac{6}{28} \times 2000$ g = 428. 6 g but the amount of H_2 actually present is 1000 g, thus H_2 is in excess and remain unreacted.

Thus, N_2 is the limiting reagent while H_2 is the excess reagent.

2 mol of N_2, i.e., 28 g of N_2 produce NH_3 = 2 mol = 34 g

2000 g of N_2 will produce $NH_3 = \frac{34}{28} \times 2000 \text{ g} = 2428.57 \text{ g}$

(ii) H_2 will remain unreacted.

(iii) Mass left unreacted = 1000 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/mol

0.50 mol Na_2CO_3 means = 0.50 × 106 g = 53 g

0.50 M Na_2CO_3 means 0.50 mol i.e., 53 g Na_2CO_3 are present in 1 L of the solution.

26. If 10 volumes of dihydrogen gas reacts with 5 volumes of dioxygen gas, how many volumes of water vapour would be produced?

Ans. H_2 and O_2 react according to the equation $2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$

Thus, 2 volumes of H_2 react with 1 volume of O_2 to produce 2 volumes of water. Hence, 10 volumes of H_2 will react completely with 5 volumes of O_2 to produce 10 volumes of water vapour.

27. Convert the following into basic units:

(i) 28.7 pm **(ii) 15.15 μm**

(iii) 25,365 mg

Ans. (i) $28.7 \text{ pm} = 28.7 \text{ pm} \times \frac{10^{-12} \text{ m}}{1 \text{pm}} = 2.87 \times 10^{-11} \text{ m}$

(ii) $15.15 \text{ μs} = 15.15 \text{ μs} \times \frac{10^{-6} \text{ s}}{1 \mu} = 1.515 \times 10^{-5} \text{ s}$

(iii) $25{,}365 \text{ mg} = 25{,}365 \text{ mg} \times \frac{1 \text{g}}{1000 \text{ mg}} \times \frac{1 \text{ kg}}{1000 \text{g}} = 2.5365 \times 10^{-2} \text{kg}$

28. Which one of the followings 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) $1\,g\,Au = \frac{1}{197}\,mol = \frac{1}{197} \times 6.02 \times 10^{23}$ atoms

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

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

(iv) $1\,g\,Cl_2 = \frac{1}{71}\,mol = \frac{1}{71} \times 6.02 \times 10^{23}$ molecules $= \frac{2}{71} \times 6.02 \times 10^{23}$ atoms

Thus, 1 g of Li has 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.

Ans. $X_{C_2H_5OH} = \frac{n(C_2H_5OH)}{n(C_2H_5OH) + n(H_2O)} = 0.040 \ldots$(i)

The aim is to find number of moles of ethanol in 1 L of the solution which is nearly = 1 L of water

(because solution is dilute)

No.of moles in 1 L water $= \frac{1000\,g}{18\,g/mol} = 55.55$ mol

Substituting $n(H_2O) = 55.55$ in eq. (i) we get

$$\frac{n(C_2H_5OH)}{n(C_2H_2OH) + 55.55} = 0.040$$

or $0.96\,n\,(C_2H_5OH) = 55.55 \times 0.040$ or $n(C_2H_5OH) = 2.31$ mol

Hence, molarity of the solution = 2.31 M

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

Ans. 1 mol of ^{12}C atoms $= 6.022 \times 10^{23}$ atoms = 12 g

thus, 6.022×10^{23} atoms of ^{12}C have mass = 12 g

$\therefore$ 1 atom of ^{12}C will have mass $= \frac{12}{6.022 \times 10^{23}}\,g = 1.9927 \times 10^{-23}\,g$

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

(i) $\dfrac{0.02856 \times 298.15 \times 0.112}{0.5785}$ **(ii)** **55.364 × 5**

(iii) **0.0125 + 0.7864 + 0.0215**

Ans. (i) The least precise term has 3 significant figures (in 0.112). Hence , the answer should have 3 significant figures.

(ii) Leaving the exact number (5), the second term has 4 significant figures. Hence, the answer should have 4 significant figures.

(iii) In the given addition the least number of decimal places in the term is 4. Hence, the answer should have the 4 significant figures.

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

Isotope	Isotopic molar mass	Abundance
^{36}Ar	335.96755 g mol/L	0.337%
^{38}Ar	37.96272 g mol/L	0.063%
^{40}Ar	39.9624 g mol/L	9.600%

Ans. Molar mass of Ar = 35.96755×0.00337 + 37.96272 × 0.00063 + 39.96924 × 0.99600 = 39.948 g/mol

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

(i) 52 mol of Ar **(ii) 52 u of He**

(iii) 52 g of He.

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

52 mol of Ar = 52 × 6.022×10^{23} atom = 3.131×10^{25} atoms

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

4 u of He =1 atom of He

52 u of He = ¼ × 52 atoms = 13 atoms

(iii) 1 mol of He = 4 g = 6.022×10^{23} atoms

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

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. Amount of carbon in 3.38 g of $CO_2 = \dfrac{12}{44} \times 3.38 \text{ g} = 0.9218 \text{ g}$

Amount of hydrogen in 0.690 g of $H_2O = \frac{2}{18} \times 0.690$ g = 0.0767 g

Compound contains only C and H.

Mole ratio C:H = $\frac{0.9218}{12} : \frac{0.0767}{1}$ = 0.0768:0.0767 = 1:1

Empirical formula = CH

$\therefore$ 10.0 L of the gas at STP weigh = 11.6 g

22.4 L of the gas at STP will weigh = $\frac{11.6}{10.0} \times 22.4$ = 25.984 g or 26 approx

Molar mass = 26 g/mol

Empirical formula mass of CH = 12 + 1 = 13

$\therefore \; n = \frac{\text{Molecular mass}}{\text{E.F.mass}} = \frac{26}{13} = 2$ $\therefore$ Molecular formula = 2 × CH = C_2H_2

35. Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction,

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

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

Ans. **Step 1.** Calculate mass of HCl in 25 mL of 0.75 M HCl

1000 mL of 0.75 M HCl contain HCl = 0.75 mol = 0.75 × 36.5 g =24.375 g

25 mL of 0.75 HCl will contain HCl = $\frac{24.375}{1000} \times 25$ g = 0.6844 g.

Step 2. Calculate mass of $CaCO_3$ reacting completely with 0.6844 g of HCl

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(L)$

2 mol of HCl, i.e., 2×36.5g = 73g HC1react completely with $CaCO_3$ = 1 mol = 100 g

$\therefore$ 0.6844 g HCl will react completely with $CaCO_3 = \frac{100 \times 0.6844}{73} g = 0.938 g$

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 HC1 react with 5.0 g of manganese dioxide?

Ans. 1 mol of MnO_2, i.e., 55 + 32 = 87 g MnO_2 react with 4 mol of HC1, i.e., 4 × 36.5 g = 146 g of HC1.

Hence, 5.0 g of MnO_2 will react with HC1 $= \frac{146}{87} \times 5.0 g = 8.40 g$