Some Basic Concepts of Chemistry

CHAPTER 1

EXAM DRILL

ANSWERS

1. **(c) :** Planck's constant = J s

Force = kg m s^{-2}

2. **(d) :** Molecular mass of HCl = 36.5

No. of gram moles in 0.365 g of HCl = $\frac{0.365}{36.5} = 0.01$

Volume of solution in L = $\frac{100}{1000} = 0.1$ L

Molarity = $\frac{n}{V \text{ in L}} = \frac{0.01}{0.1} = 0.1$ M

3. **(c) :** 22.4 L of methane reacts with 44.8 L of oxygen to give 22.4 L of CO_2 and 44.8 L of water.

4. **(a) :** Mass percent of $X = \frac{\text{Mass of } X}{\text{Mass of solution}} \times 100$

$= \frac{5}{5+45} \times 100 = 10\%$

5. **(a) :** Volume of steam = 1 litre = 10^3 cm^3

Mass of 10^3 cc steam = Density × Volume

$= 0.0006 \times 10^3 = 0.6$ g

Actual volume occupied by H_2O molecules of steam is equal to volume of water of same mass.

∴ Actual volume of H_2O molecules in 0.6 g steam

$= \frac{\text{Mass of steam}}{\text{Density of } H_2O} = \frac{0.6 \text{ g}}{1 \text{ g/cm}^3} = 0.6 \text{ cm}^3$

6. **(c) :** $2NaOH + CO_2 \rightarrow Na_2CO_3 + H_2O$

Moles of NaOH = $\frac{20}{40} = \frac{1}{2}$, moles of CO_2 = $\frac{1}{2} \times \frac{1}{2} = \frac{1}{4}$

Moles of CO = $1 - \frac{1}{4} = \frac{3}{4}$

Moles of CO_2 produced = $\frac{3}{4}$ (from CO)

∴ Extra moles of NaOH required = $\frac{3}{4} \times 2 = \frac{3}{2}$

∴ Mass of extra NaOH = $\frac{3}{2} \times 40 = 60$ g

OR

(b) $2KClO_3 \rightarrow 2KCl + 3O_2$

Thermal decomposition of 2 moles of $KClO_3$ produces 3 moles of O_2.

$$4Al + 3O_2 \rightarrow 2Al_2O_3$$

3 moles of O_2 forms 2 moles of Al_2O_3.

7. **(c) :** No. of moles of $BaCl_2$ in 100 mL of solution

$= \frac{100 \times 0.5}{1000} = 0.05$

No. of moles of Cl^- ions in $BaCl_2$ solution

$= 2 \times 0.05 = 0.1$

No. of moles of KCl in 100 mL solution = $\frac{100 \times 0.2}{1000} = 0.02$

No. of moles of Cl^- ions in KCl solution = 0.02

Total volume after mixing = 100 + 100 + 100 = 300 mL

Total moles of Cl^- ions = 0.1 + 0.02 = 0.12

∴ Moles of Cl^- present per mL = $\frac{0.12}{300}$

∴ No. of Cl^- ions per mL = $\frac{0.12}{300} \times 6.022 \times 10^{23}$

$= 2.408 \times 10^{20}$ ions

8. Avogadro's Law

9. CH_2O

10. $M_1V_1 = M_2V_2 = V_2 = \frac{M_1V_1}{M_2} = \frac{2 \times 5}{10} = 1$ L

11. **(d) :** 22.4 L of H_2 at S.T.P. contains 1 mole,

Therefore, 0.224 L of H_2 at S.T.P contains 0.01 moles.

12. **(b)**

13. **(c) :** Atomic masses of most of the elements are fractional.

14. **(a) :** Zeros between non-zero digits are significant.

OR

(c) Zero at the end or right of a number are significant provided they are on the right side of the decimal point.

15. Molality involves only masses which do not change with temperature whereas molarity involves volume which changes with temperature. Hence, molality is preferred over molarity. The SI unit of molarity is mol dm^{-3}.

16. A molecule is a group of two or more atoms of same element held together by chemical bonds *e.g.*, oxygen, hydrogen etc. A compound is a substance which is formed by two or more different types of elements which are united chemically in a fixed proportion, *e.g.*, sugar, water etc.

17. Relative Abundance (%) $\Rightarrow N_2 = 79\%$, $O_2 = 21\%$

$x_{N_2} = 0.79, x_{O_2} = 0.21, M_{N_2} = 28u, M_{O_2} = 32u$

$\therefore$ Average molecular mass of air

$= (0.79 \times 28 + 0.21 \times 32)$ u = 28.84 u

OR

When a heap of straw is ignited in air then the sum of the masses of the straw and the reacting oxygen will certainly be the same as the sum of the masses of ash, carbon dioxide and water vapour produced, *i.e.*, law of conservation of mass is applicable here. The gaseous carbon dioxide and water vapour escape from the system and only the ash remains as residue.

18. (i) $\frac{5.79 \times 0.07326}{9.003} = 0.047115 = 0.0471$

Number of significant figures in the term 5.79 (least number of significant figures) is 3. Therefore, result should have 3 significant figures.

(ii) $845 \times 0.00219 + 202 = 1.85055 + 202 = 203.85055$

= 204 (after rounding off)

Answer should have 3 significant figures.

19. (i) Same molecular formula and empirical formula – CH_4 (methane).

(ii) Different molecular formula and empirical formula – $C_6H_{12}O_6$ glucose, its empirical formula is CH_2O.

OR

Molar mass of ethanol (C_2H_5OH)

$= 2 \times 12 + 5 \times 1 + 16 + 1 = 46$

Mass percent of carbon $= \frac{24}{46} \times 100 = 52.17\%$

Mass percent of hydrogen $= \frac{6}{46} \times 100 = 13.04\%$

Mass percent of oxygen $= \frac{16}{46} \times 100 = 34.78\%$

20. Experiment 1

0.3 g of *X* combines with 0.4 g of *Y*

$\therefore$ 1 g of *X* combines with $\frac{0.4}{0.3} = 1.33$ g of *Y*

Experiment 2

18 g of *X* combines with 48 g of *Y*

$\therefore$ 1 g of *X* combines with $\frac{48}{18} = 2.67$ g of *Y*

Experiment 3

40 g of *X* combines with 160 g of *Y*

$\therefore$ 1g of *X* combines with $\frac{160}{40} = 4$ g of *Y*

For a given fixed mass of *X*, the element *Y* bears a ratio of 1.33 : 2.66 : 4 = 1 : 2 : 3. Hence, the law of multiple proportion is proved here. The law states, 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 element bear a simple ratio to one another.

21. (i) Total money to be spent = Avogadro's number

= Rs. 6.022×10^{23}

Time taken to spend Rs. $6.022 \times 10^{23} = \frac{6.022 \times 10^{23}}{10^5}$ sec

$= 6.022 \times 10^{18}$ sec

$= \frac{6.022 \times 10^{18}}{60 \times 60 \times 24 \times 365}$ yrs $= 1.91 \times 10^{11}$ yrs

(ii) One molal solution is a solution in which one mole of solute is present in 1000 g of solvent.

OR

(i) 88 g of $CO_2 = \frac{88}{44} = 2$ moles

1 mole of CO_2 has $2 \times N_A$ atoms of oxygen

2 moles of CO_2 will have $2 \times 2 \times N_A = 4N_A$ atoms of oxygen

In one mole of CO, oxygen present $= 1 \times N_A$ atoms

Thus, $4N_A$ atoms of oxygen ≡ 4 moles of CO

Mass of 4 moles of CO = 4 × 28 = 112 g

(ii) Mass of one molecule of $^{14}CO_2$

$= 14 + 2 \times 16$

$= 14 + 32 = 46$ u

22. (i) No. of moles of urea $= \frac{120}{60} = 2$

Total mass of the solution = (1000 + 120)g = 1120 g

Volume of the solution $= \frac{1120}{1.15} = 974$ mL

$\therefore$ Molarity $= \frac{2}{974} \times 1000 = 2.05$ M

(ii) 1 L solution contains 230 g H_2SO_4.

($\because$ Solution is 23% *w/v*)

No. of moles of $H_2SO_4 = \frac{230}{98} = 2.35$

Mass of solution $= v \times d = 1000 \times 1.4 = 1400$ g

Mass of solvent = 1400 – 230 = 1170 g

$\therefore$ Molality $= \frac{2.35}{1170} \times 1000 = 2$ molal

23. (a) 108 g of Ag contains 6.022×10^{23} atoms.

$\therefore$ 12.6×10^{-3} g of Ag contains

$= \frac{6.022 \times 10^{23}}{108} \times 12.6 \times 10^{-3}$ atoms

$= 7.02 \times 10^{19}$ atoms

7.02×10^{19} atoms of Al have mass $= \dfrac{27 \times 7.02 \times 10^{19}}{6.022 \times 10^{23}}$

$= 3.15 \times 10^{-3}$ g

(b) $CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$

	$CaCO_3$	$2HCl$		CO_2
	1 mole	2 mole		1 mole
or	100 g	71 g		44 g

Here, $CaCO_3$ will be limiting reagent.

20 g of $CaCO_3$ will react with HCl to produce $= \dfrac{44}{100} \times 20$

$= 8.8$ g of CO_2

24. $M_1V_1 = M_2V_2$ [M_1 = molarity of NaOH solution, V_1 = volume of NaOH solution, M_2 = molarity of HCl solution, V_2 = volume of HCl solution]

or, $1 \times V_1 = 2 \times 200$

or, $V_1 = 400$ mL

Amount of NaOH

$= \dfrac{1 \times 400}{1000} = 0.4$ moles

$NaOH + HCl \longrightarrow NaCl + H_2O$

1 mole of NaOH produced 1 mole of NaCl.

∴ Moles of NaCl produced = 0.4

∴ Mass of NaCl produced $= 0.4 \times 58.5 = 23.4$ g

25. H_2SO_4 is 98% by mass.

∴ Mass of H_2SO_4 = 98 g

Mass of solution = 100 g

∴ Volume of solution

$= \dfrac{100}{1.84}$ mL $= \dfrac{100}{1.84 \times 1000}$ L $= 0.054$ L

∴ $M_{H_2SO_4} = \dfrac{w_B}{M_B \times V(\text{in L})} = \dfrac{98}{98 \times 0.054}$

$= 18.51$ M

Let V mL of this H_2SO_4 are used to prepare 5 L of 0.5 M H_2SO_4 solution.

Then, millimoles of conc. H_2SO_4 = millimoles of dil. H_2SO_4

(∵ millimoles does not change on dilution.)

$V \times 18.51 = 5000 \times 0.5 \Rightarrow V = 135.06$ mL

26. (a) Empirical formula mass of the compound = 30 u

Molecular formula = n × Empirical formula

$n = \dfrac{\text{Molecular mass}}{\text{Empirical formula mass}} = \dfrac{180}{30} = 6$

∴ Molecular formula $= (CH_2O)_6 = C_6H_{12}O_6$

(b) Mass of solution $= 1000 \text{ mL} \times 1.25 \text{ g mL}^{-1} = 1250$ g

Mass of solute

= Molarity × Molar mass of solute × Volume (in L)

$= 3 \times 58.5 \times 1 = 175.5$ g

Mass of solvent $= 1250.0 - 175.5 = 1074.5$ g

$m = \dfrac{n}{w_A(\text{in g})} \times 1000 = \dfrac{3}{1074.5} \times 1000$

$= \dfrac{3000}{1074.5} = 2.79$ mol/kg

27. (a)

Element	%	Atomic mass	Moles	Mole ratio or Atomic ratio	Simplest whole no. ratio
C	57.8	12	$\frac{57.8}{12} = 4.82$	$\frac{4.82}{2.41} = 2$	4
H	3.6	1	$\frac{3.6}{1} = 3.60$	$\frac{3.60}{2.41} = 1.49$	3
O	38.6	16	$\frac{38.6}{16} = 2.41$	$\frac{2.41}{2.41} = 1$	2

∴ Empirical formula $= C_4H_3O_2$

Empirical formula mass

$= (4 \times 12) + (3 \times 1) + (2 \times 16) = 83$ u

Molecular mass = 2 × Vapour density

$= 2 \times 83 = 166$ u

$n = \dfrac{\text{Molecular mass}}{\text{Empirical formula mass}} = \dfrac{166}{83} = 2$

Hence, molecular formula = n × (empirical formula)

$= 2 \times (C_4H_3O_2) = C_8H_6O_4$

(b) $2BCl_3 + 3H_2 \longrightarrow 2B + 6HCl$

3 moles of H_2 is consumed to give 2 moles of B.

or, 3 × 22.4 L or 67.2 L H_2 is consumed to give 2 × 10.8 = 21.6 g of B

∴ 21.6 g of B is produced by 67.2 L H_2

∴ 108 g of B is produced by $= \dfrac{67.2 \times 108}{21.6}$ L of H_2

= 336 L of H_2

∴ 336 L of H_2 will be consumed.

28. (a) (i) Molar mass of $MgSO_4.7H_2O$

$= 24 + 32 + 4 \times 16 + 7 \times 18 = 246$ g/mol

Percentage by mass of Mg $= \dfrac{24}{246} \times 100 = 9.76\%$

(ii) Molar mass of $KAl(SO_4)_2.12H_2O$

$= 39 + 27 + 2 \times (32 + 64) + 12 \times 18 = 474$ g/mol

Percentage by mass of Al $= \dfrac{27}{474} \times 100 = 5.69\%$

(b) 100 mL of air at S.T.P contains = 21 mL of O_2

∴ 5000 mL of air at S.T.P contains $= \dfrac{21 \times 5000}{100}$ mL of O_2

= 1050 mL of O_2

No. of moles of $O_2 = \frac{1050}{22400} = 0.0469$ moles

29. (i) Molar mass of $CuSO_4 \cdot 5H_2O$
= (63.5 + 32 + 64 + 90) g/mol = 249.5 g/mol
So, mass of 6.023×10^{23} molecules of $CuSO_4 \cdot 5H_2O$
= 249.5 g
Therefore,
mass of 1×10^{22} molecules of $CuSO_4 \cdot 5H_2O$

$$= \frac{249.5 \times 1 \times 10^{22}}{6.023 \times 10^{23}} = 4.14 \text{ g}$$

(ii) 16 g of methane = 6.022×10^{23} molecules

3.2 g of methane = $\frac{6.022 \times 10^{23} \times 3.2}{16}$ molecules

1 molecule of CH_4 contains = 6 + 4 = 10 electrons
∴ 3.2 g of methane contains

$= \frac{10 \times 6.022 \times 10^{23} \times 3.2}{16}$ electrons

= 1.2044×10^{24} electrons

(iii) 100 g of chlorophyll contains Mg = 2.68 g

∴ 2.5 g of chlorophyll contains Mg = $\frac{2.68 \times 2.5}{100}$ g = 0.067 g

1 mole of Mg = 24 g = 6.022×10^{23} atoms

∴ 0.067 g of Mg = $\frac{6.022 \times 10^{23} \times 0.067}{24}$ atoms

= 1.68×10^{21} atoms

OR

(a) (i) $CH_3NO_2 + 3Cl_2 \longrightarrow CCl_3NO_2 + 3HCl$
1 mol = 61 g ; 1 mol = 164.5 g

Thus, mass of CH_3NO_2 required for

150 g of $CCl_3NO_2 = \frac{61}{164.5} \times 150 = 55.6$ g

(ii) Mg + S ⟶ MgS

	Mg	S
Mole :	1 mole	1 mole
Mass :	24 g	32 g

Thus, when 2 g S (*i.e.*, 2/32 mole) reacts with Mg, S acts as limiting reagent while Mg is in excess.

Thus, amount of product will be decided by moles of S and not by moles of Mg, therefore

1 mole S yields 1 mole MgS

∴ Moles of MgS formed = $\frac{2}{32}$

and mass of MgS formed = $\frac{2}{32} \times 56 = 3.5$ g

(b) $2KClO_3 \rightarrow 2KCl + 3O_2$
2 moles ; 2 moles ; 3 moles

Volume of O_2 produced by 2 moles or 245 g of $KClO_3$
= 3 × 22.4 L

∴ Volume of O_2 produced by 5.25 g of $KClO_3$

$= \frac{3 \times 22.4 \times 5.25}{245} = 1.44$ L

30. **(i)** 1 amu or 1 u = $\frac{1}{12} \times \frac{12 \text{ g}}{6.022 \times 10^{23}} = 1.66 \times 10^{-24}$ g

(ii) 1 mole O_2 contains = 6.022×10^{23} molecules
= $2 \times 6.022 \times 10^{23}$ atoms
= $2 \times 8 \times 6.022 \times 10^{23}$ electrons
= 9.6352×10^{24} electrons

(iii) 6.022×10^{23} molecules of methane have mass = 16 g

∴ 10^{23} molecules of methane have mass

$= \frac{16 \times 10^{23}}{6.022 \times 10^{23}}$ g = 2.657 g

(iv) 1 : 1

(v) 22.4 L of the gas at S.T.P will weigh
= 1.97 × 22.4 g = 44.13 g

Therefore, molecular mass of the gas is 44.13 u

Hence, vapour density will be $\frac{44.13}{2} = 22.065$