

Solutions to Workbook-1 [Chemistry] | Stoichiometry-I

Daily Tutorial Sheet

Level - 0

Very Short Answer Type (1 Mark)

1. Mass of 1 mole C-12 atoms is 12g
Therefore, mass of 1 C-12 atom is $12/N_A = 1.992648 \times 10^{-23} \text{ g} \approx 1.99 \times 10^{-23} \text{ g}$

2. Mass percent of calcium = $\frac{3 \times (\text{Atomic mass of calcium})}{\text{molecular mass of } \text{Ca}_3(\text{PO}_4)_2} \times 100 = \frac{120\text{u}}{310\text{u}} \times 100 = 38.71\%$

 Mass percent of phosphorous = $\frac{2 \times (\text{Atomic mass of phosphorus})}{\text{molecular mass of } \text{Ca}_3(\text{PO}_4)_2} \times 100 = \frac{2 \times 31\text{u}}{310\text{u}} \times 100 = 20\%$

 Mass percent of oxygen = $\frac{8 \times (\text{Atomic mass of oxygen})}{\text{molecular mass of } \text{Ca}_3(\text{PO}_4)_2} \times 100 = \frac{8 \times 16\text{u}}{310\text{u}} \times 100 = 41.29\%$

3. (A) No of atoms = $n \times N_A = 0.25 \times 6.023 \times 10^{23} = 1.506 \times 10^{23}$ atoms
 (B) No of molecules = $0.2 \times 6.023 \times 10^{23}$ molecules
 So number of atoms = $2 \times 0.2 \times 6.023 \times 10^{23} = 2.4092 \times 10^{23}$ atoms

4. Molecules = $0.1 N_A$
 Atoms = $4 \times 0.1 N_A$

5. Number of H-atoms = $0.2 \times 2 \times N_A = 0.4 N_A = 2.409 \times 10^{23}$
 Number of O-atoms = $0.2 \times 4 \times N_A = 4.818 \times 10^{23}$
 Number of S-atoms = $0.2 \times 1 \times N_A = 1.204 \times 10^{23}$

6. $n = \frac{w}{M_0} = \frac{143}{286} = 0.5 \text{ mole}$

 Na atoms = $0.5 \times 2 \times N_A = 6.022 \times 10^{23}$
 H atoms = $0.5 \times 20 \times N_A = 10^{24}$
 C atoms = $0.5 N_A = 3.011 \times 10^{23}$
 O atoms = $0.5 \times 13 \times N_A = 3.91 \times 10^{24}$

Short Answer Type-I (2 Marks)

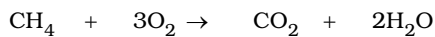
7. (A) The least precise term has 3 significant figures. Hence, the answer should have 3 significant figures.
 (B) Leaving the exact number (5), the second term has 4 significant figures. Hence, the answer should have 4 significant figures.
 (C) In the given addition, the least number of decimal places in the term is 4. Hence, the answer should have 4 significant figures.

8.

Element	%	At. wt	n	Simple Ratio	
C	34.6 %	12	$\frac{34.6}{12} = 2.88$	$\frac{2.88}{2.88} = 1 \times 3$	3
H	3.85 %	1	$\frac{3.85}{1} = 3.85$	$\frac{3.85}{2.88} = 1.33 \times 3$	4
O	61.55 %	16	$\frac{61.55}{16} = 3.846$	$\frac{3.85}{2.88} = 1.33 \times 3$	4

So, empirical formula $C_3H_4O_4$

9.



1 Mole

1 Mole = 44 gm

0.5 Mole = 22 gm

10.

(A) 1 mol of N_2 , i.e., 28 g react with 3 mol of H_2 , i.e., 6 of H_2 . 2000 g of N_2 will react with

$$H_2 = \frac{6}{28} \times 2000 \text{ g} = 428.6 \text{ g.}$$

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

$$\text{produce } NH_3 = 2 \text{ mol} = 34 \text{ g} \quad 2000 \text{ g of } N_2 \text{ will produce } NH_3 = \frac{34}{28} \times 2000 \text{ g} = 2428.57 \text{ g}$$

(B) H_2 will remain unreacted.

(C) Mass left unreacted = 1000 g - 428.6 g = 571.4 g

11.

$$M_1V_1 + M_2V_2 = M_3(V_1 + V_2)$$

$$10 \times V_1 + 3 \times (1 - V_1) = 6 \times 1$$

$$V_1 = 428 \text{ ml}$$

$$V_2 = 572 \text{ ml}$$

12.

Element	% wt	At. wt	n	Simple Rate of Mole
K	42.39	39	$\frac{42.39}{39} = 1.087$	$\frac{1.087}{0.27} = 4$
Fe	15.217	56	$\frac{15.217}{56} = 0.27$	$\frac{0.27}{0.27} = 1$
C	19.5652	12	$\frac{19.565}{12} = 1.63$	$\frac{1.63}{0.27} = 6$
N	22.826	14	$\frac{22.826}{12} = 1.63$	$\frac{1.63}{0.27} = 6$

Empirical formula = $K_4Fe(CN)_6$

Molecular Mass = 368u = Empirical formula Mass $\therefore$ Molecular formula = $K_4[Fe(CN)_6]$

13. (A) $C + O_2 \rightarrow CO_2$
 1 Mole = 44 gm.
- (B) $C + O_2 \rightarrow CO_2$, O_2 is limiting reagent
- | | | |
|------------------------|-------|------|
| Initial 1 mole = 12 gm | 16 gm | |
| After reaction = 6 gm | 0 | 22 g |
- (C) $C + O_2 \rightarrow CO_2$
 2 mole 16 gm
 (O_2 is limiting reagent)
 Mass of CO_2 formed = 22g

Short Answer Type-II (3 Marks)

14. $2A + 4B \rightarrow 3C + 4D$
 According to the above equation, 2 moles of 'A' require 4 moles of 'B' for the reaction.
 Hence, for 5 moles of 'A', the moles of 'B' required = $\times \frac{4 \text{ mol of B}}{2 \text{ mol of A}} = 10 \text{ mol B}$
 But we have only 6 moles of 'B', hence, 'B' is the limiting reagent. So amount of 'C' formed is determined by amount of 'B'.
 Since 4 moles of 'B' give 3 moles of 'C'. Hence 6 moles of 'B' will give
 $6 \text{ mol of B} \times \frac{3 \text{ mol of C}}{4 \text{ mol of B}} = 4.5 \text{ mol of C}$
15. (A) $p_1 = 1 \text{ atm}$, $T_1 = 273 \text{ K}$, $V_1 = ?$
 32 g of oxygen occupies 22.4 L of volume at STP*
 Hence, 1.6 g of oxygen will occupy, $1.6 \text{ g oxygen} \times \frac{22.4 \text{ L}}{32 \text{ g oxygen}} = 1.12 \text{ L}$
 $V_1 = 1.12 \text{ L}$
 $p_2 = \frac{p_1}{2} = \frac{1}{2} = 0.5 \text{ atm}$ $V_2 = ?$
 According to Boyle's law: $p_1 V_1 = p_2 V_2$
 $V_2 = \frac{p_1 \times V_1}{p_2} = \frac{1 \text{ atm} \times 1.12 \text{ L}}{0.5 \text{ atm}} = 2.24 \text{ L}$
- (B) Number of molecules of oxygen in the vessel = $\frac{6.022 \times 10^{23} \times 1.6}{32} = 3.011 \times 10^{22}$
16. The balanced equation for the combustion of carbon in dioxygen /air is
- | | | |
|--------|------------|-----------------------|
| $C(s)$ | $+ O_2(g)$ | $\rightarrow CO_2(g)$ |
| 1 mole | 1 mole | 1 mole |
| (12 g) | (32 g) | (44 g) |
- (A) In air, combustion is complete. Therefore, CO_2 produced from the combustion of 1 mole of carbon = 44g
- (B) As only 16 g of dioxygen is available, it can combine only with 0.5 mole of carbon, i.e., dioxygen is the limiting reactant. Hence, CO_2 produced = 22 g.
- (C) Here again, dioxygen is the limiting reactant. 16 g of dioxygen can combine only with 0.5 mole of carbon. CO_2 produced again is equal to 22g.

17. Fractional abundance of ^{35}Cl = 0.7577, Molar mass = 34.9689
 Fractional abundance of ^{37}Cl = 0.2423, Molar mass = 36.9659
 Average atomic mass = $(0.7577)(34.9689 \text{ amu}) + (0.2423)(36.9659 \text{ amu})$
 $= 26.4959 + 8.9568 = 35.4527$

18. Average Atomic Mass

$$= \frac{\left\{ \begin{aligned} &(\text{Natural abundance of } ^1\text{H} \times \text{molar mass}) \\ &+ (\text{Natural abundance of } ^2\text{H} \times \text{molar mass of } ^2\text{H}) \end{aligned} \right\}}{100} = \frac{99.985 \times 1 + 0.015 \times 2}{100}$$

$$= \frac{99.985 + 0.030}{100} = \frac{100.015}{100} = 1.00015 \text{ u}$$

19. Number of moles of HCl = $250 \text{ mL} \times \frac{0.76\text{M}}{1000} = 0.19 \text{ mol}$

Mass of CaCO_3 = 1000 g

$$\text{Number of moles of } \text{CaCO}_3 = \frac{1000\text{g}}{100\text{g}} = 10 \text{ mol}$$

According to given equation 1 mol of CaCO_3 (s) requires 2 mol of HCl (aq). Hence, for the reaction of 10 mol of CaCO_3 (s) number of moles of HCl required would be:

$$10 \text{ mol } \text{CaCO}_3 \times \frac{2 \text{ mol } \text{HCl}(\text{aq})}{1 \text{ mol } \text{CaCO}_3(\text{s})} = 20 \text{ mol } \text{HCl}(\text{aq})$$

But we have only 0.19 mol HCl (aq), hence, HCl (aq) is limiting reagent.

So amount of CaCl_2 formed will depend on the amount of HCl available.

Since, 2 mol HCl (aq) forms 1 mol of CaCl_2 , therefore, 0.19 mol of HCl (aq) would give:

$$0.19 \text{ mol } \text{HCl}(\text{aq}) \times \frac{1 \text{ mol } \text{CaCl}_2(\text{aq})}{2 \text{ mol } \text{HCl}(\text{aq})} = 0.095 \text{ mol} \quad \text{or} \quad 0.095 \times \text{molar mass of } \text{CaCl}_2 = 0.095 \times 111 = 10.54 \text{ g}$$

Long Answer Type (5 Marks)

20. Mass of NaOH = 4 g

$$\text{Number of moles of } \text{NaOH} = \frac{4\text{g}}{40\text{g}} = 0.1 \text{ mol}$$

Mass of H_2O = 36 g

$$\text{Number of moles of } = \frac{36\text{g}}{18\text{g}} = 2 \text{ mol}$$

$$\text{Mole fraction of water} = \frac{\text{Number of moles of } \text{H}_2\text{O}}{\text{No. of moles of water} + \text{No. of moles of } \text{NaOH}} = \frac{2}{2 + 0.1} = \frac{2}{2.1} = 0.95$$

$$\text{Mole fraction of } \text{NaOH} = \frac{\text{Number of moles of } \text{NaOH}}{\text{No. of moles of } \text{NaOH} + \text{No. of moles of Water}} = \frac{0.1}{2 + 0.1} = \frac{0.1}{2.1} = 0.047$$

Mass of solution = mass of water + mass of NaOH = 36 g + 4 g = 40 g

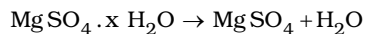
Volume of solution = $40 \times 1 = 40 \text{ mL}$

(Since specific gravity of solution is = 1 g mL^{-1})

$$\text{Molarity of solution} = \frac{\text{Number of moles of solute}}{\text{Volume of solution in litre}} = \frac{0.1 \text{ mol } \text{NaOH}}{0.04 \text{ L}} = 2.5 \text{ M}$$

21.

Element	%	At .Mass	Ratio	Simple Ratio
Mg	20	24	0.833	= 1
S	20.66	32	0.833	= 1
O	53.33	16	3.33	= 4



$$\% \text{ of H}_2\text{O} = 51.2 = \frac{18x \times 100}{120 + 18x} \quad x = 7.$$

Molecular formula for hydrated salt is $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$

22.

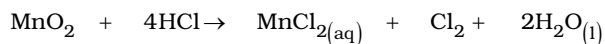
(A) $N = \frac{3.65}{36.5} \times \frac{1000}{500} = 0.2 \text{ gm. equivalent/L}$

(B) $N_1 V_1 = N_2 V_2$
 $0.2 \times V = 0.05 \times 200$
 $V = 62.5 \text{ ml.}$

23.

Element	% wt	At. wt	N	Simple Rate of Mole
Na	14.31	23	0.622	2
S	9.97	32	0.311	1
O	69.50	16	4.34	14
H	6.22	1	6.22	20

24.



If 87 4×36.5

So, 5 x

$$x = \frac{146 \times 5}{87} = 8.39 \text{ gm}$$

25.

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) = $1 + 14 + 48 = 63 \text{ g mol}^{-1}$

$$\text{Moles in 69 g HNO}_3 = \frac{69 \text{ g}}{63 \text{ g mol}^{-1}} = 1.095 \text{ mole}$$

$$\text{Volume of 100 g nitric acid solution} = \frac{100 \text{ g}}{1.41 \text{ g mL}^{-1}} = 70.92 \text{ mL} = 0.07092 \text{ L}$$

$$\text{Conc. of HNO}_3 \text{ in moles per litre} = \frac{1.095}{0.07092 \text{ L}} = 15.44 \text{ M}$$