

TRIUMPH MHT-CET CHEMISTRY SOLUTIONS to MCQs

Salient Features

- ☞ Detailed solutions provided for difficult MCQs as per the concepts emphasized in the syllabus
- ☞ **Smart Keys** (Smart Code, Caution, Thinking Hatke, Shortcut) - Multiple Study Techniques to enhance understanding of concepts and problem solving skills
- ☞ Solutions to Evaluation Test for each chapter
- ☞ Solutions to two Model Question Papers
- ☞ Solutions to Two MHT-CET 2023 Question Papers

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PREFACE

Target's **Triumph MHT-CET Chemistry Solutions to MCQs** book provides students with holistic comprehension of principles of chemistry through solutions to MCQs based on the concepts emphasized in the syllabus.

It includes **Smart Keys** (Smart Code, Caution, Thinking Hatke, and Shortcut), which offer supplemental explanations for the tricky questions and are intended to help students how to approach problems in novel ways in the shortest possible time with accuracy.

- **Smart Code** showcases simple and smart mnemonic.
- **Caution** apprises students about mistakes often made while solving MCQs.
- **Shortcuts** comprise formulae based short cuts considering their usage in solving MCQ.
- **Thinking Hatke** reveals quick witted approach to crack the specific question.

Roadmaps for the sequences of organic reactions are drawn in the solutions to the newly added chapter "**Organic Reactions: Compilation of Organic Reaction Based MCQs**" making them a helpful novelty in learning organic chemistry.

Solutions to two **Model Question Papers** and two **MHT-CET 2023 Question Papers** are also included in this book.

All the features of this book are designed keeping the following elements in mind:

Time management, easy memorization or revision, and non-conventional yet simple methods for MCQ solving.

We hope the book benefits the learner as we have envisioned.

Publisher

Edition: First

The journey to create a complete book is strewn with triumphs, failures and near misses. If you think we've nearly missed something or want to applaud us for our triumphs, we'd love to hear from you.

Please write to us on: mail@targetpublications.org

Disclaimer

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Classical Thinking

1.1 Introduction

1. (C)

1.2 Nature of chemistry

1. (C) 2. (C) 3. (C)

4. (D) 5. (D) 6. (D)

1.3 Properties of matter and their measurement

1. (B) 2. (B) 3. (D)

4. (A) 5. (B) 6. (A)

7. (B)

1.4 Laws of chemical combination

1. (C) 2. (C)

3. (B) $\text{BaCl}_2 + \text{H}_2\text{SO}_4 \longrightarrow \text{HCl} + \text{BaSO}_4$
∴ $20.8 + 9.8 = 7.3 + x$
 $x = 23.3$

4. (C) 5. (A) 6. (B)

7. (C) 8. (B)

1.5 Avogadro law

1. (C)

1.6 Dalton's atomic theory

1. (A)

1.7 Atomic and molecular masses

1. (A)

2. (A) Isotopes are the atoms of the same element having same atomic number (i.e., containing same number of protons and electrons) but different mass number (i.e., different number of neutrons).

3. (A)

1.8 Mole concept and molar mass

1. (D) 2. (C) 3. (C)

4. (C)

5. (D) Molecular formula of benzene is C_6H_6 .
∴ Molar mass = $12 \times 6 + 6 \times 1$
 $= 72 + 6 = 78 \text{ g mol}^{-1}$ ∴ 1 mole of benzene is equal to 78 g of C_6H_6 .6. (B) Molar mass of $\text{H}_2 = 2 \text{ g mol}^{-1}$
2 g will contain 6.022×10^{23} molecules of H_2 .∴ 1 g of H_2 will contain $\frac{6.022 \times 10^{23}}{2}$ molecules
 $= 3.011 \times 10^{23}$ molecules $\approx 3 \times 10^{23}$ molecules

7. (A) 8. (C)

9. (D) Atomic mass of the element
 $= 1.792 \times 10^{-22} \times 6.022 \times 10^{23}$
 $= 108$ 10. (C) 1 mole of ozone (O_3) = 48 g
∴ 0.5 mole of ozone (O_3) = $\frac{0.5 \times 48}{1} = 24 \text{ g}$ 11. (B) Number of molecules = $n \times 6.022 \times 10^{23}$
Now, $n = \frac{\text{mass of oxygen}}{\text{molar mass of oxygen}} = \frac{16}{32} = 0.5 \text{ mol}$
∴ Number of molecules = $0.5 \times 6.022 \times 10^{23}$
 $= 3.011 \times 10^{23}$ **1.9 Moles and gases**

1. (A)

2. (B) At S.T.P,
22.4 dm³ of any gas $\equiv 6.022 \times 10^{23}$ molecules
 $\equiv 6.022 \times 10^{23}$ SO_2 molecules
 $\equiv 6.022 \times 10^{23}$ S atoms

Critical Thinking

1.2 Nature of chemistry

1. (A) 2. (D)

3. (D) The constituents of a compound cannot be easily separated by physical method.

4. (C) Mixture of any two liquids may be homogeneous or heterogeneous mixtures.

5. (D) A rusty nail is a mixture.

6. (D)



1.3 Properties of matter and their measurement

1. (D) 2. (A)
3. (C) $1 \text{ L} = 10^{-3} \text{ m}^3 = 10^3 \text{ cm}^3 = 1 \text{ dm}^3 = 10^3 \text{ mL}$.

4. (D)

(A)	The mass of a body does not vary as its position changes.
(B)	The SI unit of length is metre.
(C)	A volumetric flask is used to prepare a known volume of a solution.

5. (A) $^{\circ}\text{F} = \frac{9}{5} (^{\circ}\text{C}) + 32$
 $= \frac{9}{5} (40) + 32$
 $= 72 + 32$
 $= 104 ^{\circ}\text{F}$

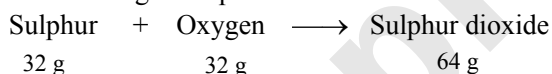
6. (B) $^{\circ}\text{F} = \frac{9}{5} (^{\circ}\text{C}) + 32$
 $50 = \frac{9}{5} (^{\circ}\text{C}) + 32$
 $^{\circ}\text{C} = \frac{(50 - 32) \times 5}{9} = 10 ^{\circ}\text{C}$

1.4 Laws of chemical combination

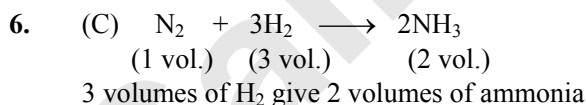
1. (C) 2. (D) 3. (B)

4. (A)

5. (B) 32 g of sulphur combine with 32 g oxygen to form 64 g of sulphur dioxide as follows:



Hence, $(0.5 \times 32 = 16 \text{ g})$ of sulphur will combine with $(0.5 \times 32 = 16 \text{ g})$ of oxygen to give $(0.5 \times 64 = 32 \text{ g})$ sulphur dioxide.



\therefore 2 L of H_2 will give $= \frac{2 \times 2}{3}$ L of ammonia
 $= 1.33 \text{ L}$ of ammonia

1.6 Dalton's atomic theory

1. (D)

1.7 Atomic and molecular masses

1. (D)

2. (B) One atomic mass unit is defined as a mass exactly equal to one-twelfth of the mass of one C-12 atom.
 $1 \text{ a.m.u.} = 1.66 \times 10^{-24} \text{ g}$.

$$\begin{aligned} 1 \text{ atom of } ^{12}\text{C} &= 12 \text{ a.m.u.} \\ &= 12 \times 1.66 \times 10^{-24} \text{ g} \\ &= 1.9923 \times 10^{-23} \text{ g} \end{aligned}$$

3. (B)

4. (A) Average atomic mass of Boron (B) =
 $\frac{(\text{At. mass of } ^{10}\text{B} \times \% \text{Abundance}) + (\text{At. mass of } ^{11}\text{B} \times \% \text{Abundance})}{100}$
 $= \frac{(10.13 \text{ u} \times 19.60) + (11.009 \text{ u} \times 80.40)}{100} = 10.84 \text{ u}$

5. (A) Average atomic mass of X
 $= \frac{200 \times 90 + 199 \times 8 + 202 \times 2}{100} = 199.96 \approx 200 \text{ u}$

6. (A) Molecular mass of $\text{C}_6\text{H}_5\text{Cl}$
 $= (6 \times \text{Average atomic mass of C})$
 $+ (5 \times \text{Average atomic mass of H})$
 $+ (1 \times \text{Average atomic mass of Cl})$
 $= (6 \times 12.0 \text{ u}) + (5 \times 1.0 \text{ u}) + (1 \times 35.5 \text{ u})$
 $= 112.5 \text{ u}$

7. (D) Molecular mass of $\text{O}_2 = 32 \text{ u}$
 \therefore Mass of 1 molecule = 32 u
 \therefore Mass of 1 molecule of O_2
 $= 32 \times 1.66 \times 10^{-24} \text{ g} = 53.1 \times 10^{-24} \text{ g}$

8. (C) Formula mass of KCl
 $= \text{Average atomic mass of K}$
 $+ \text{Average atomic mass of Cl}$
 $= 39.1 + 35.5 = 74.6 \text{ u}$

1.8 Mole concept and molar mass

1. (B) Molecular weight of sodium oxide (Na_2O)
 $= 46 + 16 = 62 \text{ u}$
 $62 \text{ g of Na}_2\text{O} = 1 \text{ mole}$
 $620 \text{ g of Na}_2\text{O} = 10 \text{ moles}$.

2. (A)

3. (C) 6.022×10^{23} atoms of H weighs 1 g.
 \therefore Mass of 1 atom of hydrogen = $\frac{1}{6.022 \times 10^{23}}$
 $= 1.6 \times 10^{-24} \text{ g}$

4. (D) 1 mole $\equiv 6.022 \times 10^{23}$ electrons
 One electron weighs $9.108 \times 10^{-31} \text{ kg}$
 \therefore 1 mole of electrons weighs
 $6.022 \times 10^{23} \times 9.108 \times 10^{-31} \text{ kg}$
 \therefore Number of moles that will weigh 1 kg
 $= \frac{1}{6.022 \times 10^{23} \times 9.108 \times 10^{-31}} \text{ moles}$
 \therefore $\frac{1}{9.108 \times 6.022} \times 10^8$ moles of electrons will weigh one kilogram.

5. (D) Molar mass of $\text{NH}_3 = 14 + (3 \times 1) = 17 \text{ g mol}^{-1}$
 Number of moles = $\frac{4.25}{17} = 0.25 \text{ mol}$



- Number of molecules of NH_3
 $= 0.25 \times N_A = 1.506 \times 10^{23}$ molecules
 One molecule of NH_3 contains 4 atoms.
 $\therefore 1.506 \times 10^{23}$ molecules will contain
 $= 1.506 \times 10^{23} \times 4$
 $= 6.024 \times 10^{23}$ atoms $\approx 6 \times 10^{23}$ atoms.
6. (D) Number of atoms $= n \times N_A \times \text{Atomicity}$
 Where atomicity is the number of atoms in one molecule
 $18 \text{ g of H}_2\text{O} \equiv 1 \text{ mole} = 3 \times N_A$ atoms
 $16 \text{ g of O}_2 \equiv \frac{1}{2} \text{ mole} = 2 \times \frac{1}{2} N_A$ atoms
 $4.4 \text{ g of CO}_2 \equiv \frac{1}{10} \text{ mole} \equiv 3 \times \frac{1}{10} \times N_A$ atoms
 $16 \text{ g of CH}_4 \equiv 1 \text{ mole} = 5 \times N_A$ atoms
 \therefore Maximum number of atoms is present in 16 g of CH_4 .
7. (C) Number of atoms $= n \times N_A \times \text{Atomicity}$
 Number of S atoms $= 6.022 \times 10^{23} \times 0.2 \times 8$
 $\approx 9.63 \times 10^{23}$
8. (A) Number of moles in 4.4 g of CO_2
 $= \frac{4.4}{44} = 0.1$
 Number of oxygen atoms in 1 mole of CO_2
 $= 2 \times N_A$
 \therefore Number of oxygen atoms in 0.1 mole of CO_2
 $= 0.1 \times 2 \times N_A = 0.2 \times 6.022 \times 10^{23} = 1.20 \times 10^{23}$
9. (C) Total number of atoms in a given amount of H_2O
 $= n \times N_A \times 3$
 $= \frac{0.05}{18} \times 6.022 \times 10^{23} \times 3 = 5.05 \times 10^{21}$
10. (B) 6.022×10^{23} dioxygen molecules are present in 1 mole i.e., 32 g of dioxygen.
 $\therefore 1.8 \times 10^{22}$ dioxygen molecules will be present in
 $\frac{1.8 \times 10^{22} \times 32}{6.022 \times 10^{23}} = 0.96 \text{ g of dioxygen}$
11. (A) Molecular weight of $\text{C}_{60}\text{H}_{122}$
 $= 12 \times 60 + 122 \times 1 = 720 + 122 = 842 \text{ u}$
 $\therefore 6.022 \times 10^{23}$ molecules $= 842 \text{ g}$
 1 molecule $= \frac{842}{6.022 \times 10^{23}}$
 $= 139.82 \times 10^{-23}$
 $\approx 1.4 \times 10^{-21} \text{ g}$
12. (A) 1 mole of BaCO_3 contains 3 moles of oxygen atoms
 $\therefore 1.5$ moles of oxygen $\equiv \frac{1}{3} \times 1.5 = \frac{1}{2}$
 $= 0.5$ moles of BaCO_3

13. (A) 1 L of air $= 1000 \text{ mL} = 1000 \text{ cc}$.
 1000 cc of air contains 210 cc of O_2
 1 mole $= 22.4 \text{ L} = 22400 \text{ cc}$.
 \therefore Number of moles of $\text{O}_2 = \frac{210}{22400} = 0.0093$ moles
14. (A) 16 g O_2 has number of moles $= \frac{16}{32} = \frac{1}{2}$
 14 g N_2 has number of moles $= \frac{14}{28} = \frac{1}{2}$
 Number of moles is same, so number of molecules is same.
15. (D) $d = \frac{M}{V}$ ($d = \text{density}$, $M = \text{mass}$, $V = \text{volume}$)
 Since $d = 1 \text{ g/mL}$, So, $M = V$
 $18 \text{ g} = 18 \text{ mL}$
 $18 \text{ mL} = N_A$ molecules ($N_A = \text{Avogadro's number}$)
 $1000 \text{ mL} = \frac{N_A}{18} \times 1000 = 55.55 N_A$.
16. (B) 1 mole of water
 $= 18 \text{ g of water}$
 $= 6.022 \times 10^{23}$ molecules of water
 $\therefore 18$ moles of water
 $= 18 \times 6.022 \times 10^{23}$ molecules of water
 $= 1.08396 \times 10^{25}$ molecules of water

1.9 Moles and gases

1. (C) 1 mole of nitrogen gas $\equiv 22.4 \text{ L of N}_2$
 (molar volume at S.T.P.)
 0.5 mole of nitrogen gas $= 11.2 \text{ L of N}_2$ at S.T.P.
2. (C) Volume occupied by 1 mole of any gas at STP
 $= 22.4 \text{ dm}^3$
 \therefore Volume occupied by 4.4 g of CO_2 i.e., 0.1 mole of CO_2 at STP $= 2.24 \text{ dm}^3 = 2.24 \text{ L}$
3. (A) At STP, $22.4 \text{ L} (22400 \text{ cm}^3)$ oxygen gas
 $= 1 \text{ mole oxygen gas}$
 Hence, 11.2 cm^3 corresponds to
 $\frac{11.2}{22400} = 0.0005 \text{ mole}$
4. (B) Number of moles $= \frac{\text{Mass of a substance}}{\text{Molar mass of the substance}}$
 $= \frac{60.0 \text{ g}}{30 \text{ g mol}^{-1}} = 2 \text{ mol}$
 Number of moles of a gas (n)
 $= \frac{\text{Volume of the gas at STP}}{\text{Molar volume of the gas}}$
 \therefore Volume of the gas at STP
 $= n \times \text{Molar volume of the gas}$
 $= 2 \text{ mol} \times 22.4 \text{ dm}^3 \text{ mol}^{-1}$
 $= 44.8 \text{ dm}^3$



Concept Fusion

1. (D)
2. (A) $V \propto n$
 Number of moles (n) = $\frac{\text{Mass of the substance}}{\text{Molar mass of the substance}}$
 $\therefore n = \frac{\text{mass}}{\text{atomic mass}(M)} \quad \therefore V \propto n \propto \frac{1}{M}$
 Atomic Mass of O = 16
 Atomic Mass of N = 14
 $\therefore \frac{V_{(O)}}{V_{(N)}} = \frac{n_{(O)}}{n_{(N)}} = \frac{M_{(N)}}{M_{(O)}}$
 $\frac{V_{(O)}}{V_{(N)}} = \frac{14}{16} = \frac{7}{8}$
 \therefore The ratio is 7 : 8
3. (C) Baking soda or sodium hydrogen carbonate (NaHCO_3) is a compound. Diamond and charcoal are different forms of the element carbon. 22 carat gold is an alloy of gold with other metals (mainly copper). Hence, it is a mixture.
4. (C) In compound B, 32 parts of X react with 84 parts of Y.
 \therefore In compound B, 16 parts of X react with 42 parts of Y.
 In compound C, 16 parts of X react with x parts of Y.
 The ratio of masses of Y, which combine with fixed mass of X in compounds B and C, is 3:5.
- | | | |
|---|-----|---|
| B | 42 | 3 |
| C | x | 5 |
- $\therefore x = \frac{42 \times 5}{3} = 70$

5. (C) 100 g of haemoglobin contains 0.33 g of Fe.
 \therefore 67200 g of haemoglobin contains
 $= \frac{67200 \times 0.33}{100} = 221.76$ g of Fe
 Number of atoms of Fe = $\frac{221.76}{56} = 3.96 \approx 4$
6. (C) Let the mass of CH_4 and SO_2 be w_1 and w_2 , respectively.
 $\frac{w_1}{w_2} = \frac{1}{2}$
 $\therefore \frac{n_1}{n_2} = \frac{w_1}{M_1} \times \frac{M_2}{w_2} = \frac{1}{16} \times \frac{64}{2} = \frac{2}{1}$
 $\therefore \frac{n_2}{n_1} = \frac{1}{2}$
 Therefore, the ratio of number of molecules of SO_2 to CH_4 is 1:2.
7. (A) Mass of drop = volume of drop \times Density
 $= \frac{1}{25} \times 1 = \frac{1}{25}$ g
 Number of water molecules
 $= \text{Moles of water} \times N_0 = \frac{1}{25 \times 18} N_0 = \frac{0.02}{9} N_0$
8. (A)
9. (A) 5.6 L at S.T.P. weighs 7.5 g.
 \therefore 22.4 L at S.T.P weighs $\frac{7.5 \times 22.4}{5.6} = 30$ g
 \Rightarrow Molar mass of gas = 30 g mol^{-1}
 Hence, the gas is NO

MHT-CET Previous Years' Questions

1. (B) $2\text{C}_2\text{H}_5\text{OH}_{(l)} + 2\text{Na}_{(s)} \longrightarrow 2\text{C}_2\text{H}_5\text{ONa} + \text{H}_{2(g)} \uparrow$
 2 mole of Na = 1 mole of H_2
 $= 2 \text{ g} = 2 \times 10^{-3} \text{ kg}$
2. (A) 3. (C)
4. (C) Average atomic mass =
 $\frac{\text{atomic mass of } ^{10}\text{B} \times \text{percentage} + \text{atomic mass of } ^{11}\text{B} \times \text{percentage}}{100}$
 Let the % abundance of ^{10}B isotope = x .
 % abundance of ^{11}B isotope = $100 - x$.
 Average atomic mass = 10.81
 From formula, Average atomic mass
 $= \frac{10 \times x + 11 \times (100 - x)}{100} = 10.81$
 $10x + 1100 - 11x = 10.81 \times 100$
 $-x = -1100 + 1081$
 $x = 19$

- Percentage abundance of lighter isotope,
 $^{10}\text{B} = 19\%$.
5. (D) $^{\circ}\text{F} = \frac{9}{5} (^{\circ}\text{C}) + 32 = 32 \times \frac{9}{5} + 32 = 89.6 ^{\circ}\text{F}$
6. (A) 100 mL = 100 g (Density = 1 g/mL)
 Number of moles = $\frac{100}{18} = 5.55$ mol
 Number of molecules = $5.55 \times 6.022 \times 10^{23}$
 $= 33.45 \times 10^{23}$
7. (B) At STP, volume = $22.4 \text{ dm}^3 = 0.022414 \text{ m}^3$
8. (D)
9. (C) Since the mass is same,
 Number of atoms $\propto \frac{1}{\text{M.W.}}$
 Among the given, Na has the smallest atomic mass.



10. (A)
11. (C) Law of multiple proportions is applicable when two or more elements combine in more than one form.
12. (A) Moles of Ar = $\frac{3.99}{39.9} = 0.1$ mol
1 mol of Ar = 6.022×10^{23} atoms
 \therefore 0.1 mol of Ar = 6.022×10^{22} atoms
13. (A) At STP, $22.4 \text{ dm}^3 = 1$ mol of NH_3
 5.6 dm^3 at STP = $\frac{5.6 \text{ dm}^3}{22.4 \text{ dm}^3} = 0.25$ mol
14. (D) $2\text{KClO}_{3(s)} \longrightarrow 2\text{KCl}_{(s)} + 3\text{O}_{2(g)}$
 $2 \times 74.5 \text{ g} \quad 3 \times 22.4 \text{ L}$
 $= 149 \text{ g} \quad = 67.2 \text{ L}$
Now, 67.2 L of $\text{O}_2 = 149$ g of KCl at STP
33.6 l of $\text{O}_2 = x$ g of KCl
 $\therefore x = \frac{149 \times 33.6}{67.2} = 74.5$ g of KCl
15. (D)
i. $n_{\text{Ar}} = \frac{13.3}{39.9} = 0.33$ mol
ii. $n_{\text{O}_2} = \frac{24}{32} = 0.75$ mol
iii. $n_{\text{CO}_2} = \frac{11}{44} = 0.25$ mol
iv. $n_{\text{CH}_4} = \frac{16}{16} = 1$ mol
 \therefore Maximum no. of moles = Maximum no. of molecules.
16. (C) Vol. of NH_3 gas at STP = $5.6 \text{ cm}^3 = 5.6 \times 10^{-3} \text{ dm}^3$
Now, 22.4 dm^3 of $\text{NH}_3 = 1$ mole of NH_3 at STP
 $\therefore 5.6 \times 10^{-3} \text{ dm}^3$ of $\text{NH}_3 = \frac{5.6 \times 10^{-3}}{22.4} = 2.5 \times 10^{-4}$ mol of NH_3
No. of atoms = $2.5 \times 10^{-4} \times 6.022 \times 10^{23} \times 4 = 6.022 \times 10^{20}$ atoms
17. (C) $\text{CH}_4 + 2\text{O}_2 \longrightarrow \text{CO}_2 + 2\text{H}_2\text{O}$
1 mole of methane required = $2 \times 22.4 \text{ dm}^3$ of O_2
 \therefore 0.25 mole of methane required = $2 \times 22.4 \times 0.25 = 11.2 \text{ dm}^3$ of O_2
18. (B) 1 mol of an element = 6.022×10^{23} atoms
 $\therefore 3.01 \times 10^{24}$ atoms = $\frac{3.01 \times 10^{24}}{6.022 \times 10^{23}} = 4.998 \text{ mol} \approx 5$ mol
No. of moles = $\frac{\text{Mass}}{\text{atomic mass}}$
 \therefore Mass = $5 \times 21.13 = 105.65 \text{ g mol}^{-1}$
19. (D)
20. (D) $224 \text{ cm}^3 = 0.224 \text{ dm}^3$
 $22.4 \text{ dm}^3 = 1 \text{ mol} = 6.022 \times 10^{23}$ molecules
 $\therefore 0.224 \text{ dm}^3$ of a gas = 0.01 mol = 6.022×10^{21} molecules
21. (B) 1 mol urea = 60 g urea = 6.022×10^{23} molecules
 $5.4 \text{ g urea} = \frac{5.4 \text{ g} \times 6.022 \times 10^{23}}{60 \text{ g mol}^{-1}} = 5.4 \times 10^{22}$ molecules
22. (B) $22.4 \text{ dm}^3 \text{ CH}_4$ at STP = 1 mol = 16 g of CH_4
 $\therefore 44.8 \text{ dm}^3$ of CH_4 at STP = 16 g $\times 2 = 32$ g of CH_4
23. (B) Law of multiple proportions is applicable when two or more elements combine in more than one form.
24. (B) $\text{K} = ^\circ\text{C} + 273 = -197^\circ\text{C} + 273 = 76 \text{ K}$
25. (B)
26. (C) 1 mol of $\text{H}_2\text{O} = 18$ g
 $\therefore 0.25$ mol of $\text{H}_2\text{O} = 18 \text{ g} \times 0.25 = 4.5$ g
27. (C) 1 mole of Ar = 39 g of Ar = 6.022×10^{23} atoms of Ar
 $\therefore 52$ moles of Ar = $6.022 \times 10^{23} \times 52 = 3.1 \times 10^{25}$ atoms of Ar
28. (A) Molar mass of $\text{CH}_4 = 16 \text{ g mol}^{-1}$
16 g of $\text{CH}_4 = 22.4 \text{ dm}^3$ at STP
 $\therefore 24$ g of $\text{CH}_4 = \frac{22.4 \times 24}{16} = 33.6 \text{ dm}^3$
29. (A) $\text{H}_{2(g)} + \frac{1}{2} \text{O}_{2(g)} \longrightarrow \text{H}_2\text{O}_{(l)}$
11.2 dm^3 of O_2 gives 18 g of water at STP.
 $\therefore 9$ g water = $\frac{11.2 \times 9}{18} = 5.6 \text{ dm}^3$ of O_2
30. (B) C_2H_6 : Molar mass = 30 g mol^{-1}
Ethane
30 g of ethane = 22.4 dm^3 at STP
 $\therefore 75$ g of ethane = $\frac{22.4 \times 75}{30} = 56.0 \text{ dm}^3$
31. (B) No. of moles of urea = $\frac{\text{Mass of urea}}{\text{Molar mass of urea}} = \frac{5.4}{60} = 0.09$ moles
32. (B) $\frac{0.863 \text{ g}}{\text{cm}^3} \times \frac{1000 \text{ cm}^3 \text{ dm}^{-3}}{1000 \text{ g kg}^{-1}} = 0.863 \text{ kg dm}^{-3}$
33. (C) 1 mol urea = 60 g urea = 6.022×10^{23} molecules = $4 \times 6.022 \times 10^{23}$ H atoms
 $\therefore 6$ g urea = $4 \times 6.022 \times 10^{22} = 2.41 \times 10^{23}$ H atoms



34. (B) Molecular mass of O_2 is 32 u.
Mass of 1 molecule of O_2
 $= 32 \times 1.6606 \times 10^{-24} \text{ g} = 53.13 \times 10^{-24} \text{ g}$
35. (D) 22.4 L of O_2 at STP = 1 mole of O_2
 $\therefore 5.6 \text{ L of } O_2 = \frac{5.6 \text{ L}}{22.4 \text{ L mol}^{-1}} = \frac{1}{4} \text{ mole } O_2$
36. (C)
- | | |
|--------------|--|
| (A) | 1 mL of water = 1 g water
(Density of water = 1 g/mL) |
| \therefore | 10 mL of water = 10 g |
| (B) | 1 mol of CH_4 = 16 g |
| \therefore | $\frac{1}{2}$ mol of CH_4 = 8 g |
| (C) | 1 mole of C atom = 12 g |
| (D) | 6.022×10^{23} atoms of oxygen = 16 g |
| \therefore | 3.011×10^{23} atoms of oxygen = 8 g |
37. (C) Nitrogen_(g) + Hydrogen_(g) \longrightarrow Ammonia_(g)
[1 L] [3 L] [2 L]
38. (B) Volume of a drop = 0.05 mL
Since density of water is 1 g/mL, the mass of a drop of water is 0.05 g.
Now, 1 mol of H_2O = 18 g
 $= 6.022 \times 10^{23}$ molecules
 $\therefore 0.05 \text{ g of water} = \frac{0.05 \text{ g} \times 6.022 \times 10^{23}}{18 \text{ g mol}^{-1}}$
 $= 1.67 \times 10^{21}$ molecules
39. (A) At STP 22.4 L = 1 mol
 $\therefore 1 \text{ L} = \frac{1 \text{ L}}{22.4 \text{ L mol}^{-1}} = 0.0446 \text{ mol}$
Now,
Mole = $\frac{\text{Mass}}{\text{Molar Mass}}$
 $\therefore \text{Molar Mass} = \frac{\text{Mass}}{\text{Mole}} = \frac{1.16 \text{ g}}{0.0446 \text{ mol}} = 26 \text{ g mol}^{-1}$
Among the given, C_2H_2 has molar mass of 26 g mol^{-1} .
40. (A)
- | | |
|--------------|--|
| (A) | 1 mol of CH_4 = 16 g |
| \therefore | $\frac{1}{4}$ mol of CH_4 = 4 g |
| (B) | 6.022×10^{23} atoms of oxygen = 16 g |
| \therefore | 3.011×10^{23} atoms of oxygen = 8 g |
| (C) | 1 g atom C = 1 mole of C atom = 12 g |
| (D) | 6.022×10^{23} molecules of water = 18 g |
41. (C) $2H_2 + O_2 \longrightarrow 2H_2O$
2 Vol 1 Vol 2 Vol
10 Vol 5 Vol 10 Vol
10 Volume of H_2 when reacts with 5 volume of O_2 , it forms 10 volume of H_2O .

42. (C) $N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$
1 Vol 3 Vol 2 Vol
 \therefore Volume ratio = (1 : 3 : 2)
43. (B) Given: $t^\circ C = 60$
 $^\circ F = \frac{9}{5} \times t^\circ C + 32 = \frac{9}{5} \times 60 + 32 = 140^\circ F$
44. (C) At STP, 1 mole of any gas occupies 22.4 L of volume.
45. (A) No. of Na atoms = 200 atoms
 $n = \frac{\text{No. of Na atoms}}{N_A} = \frac{200}{N_A}$
and $n = \frac{W}{M}$
 $\therefore \frac{W}{M} = \frac{200}{N_A}$
 $\therefore W = \frac{200}{N_A} \times M = \frac{200}{6.022 \times 10^{23}} \times 23$
 $= 7.64 \times 10^{-21} \text{ g}$
46. (B) Nitrogen_(g) + Hydrogen_(g) \longrightarrow Ammonia_(g)
[1 L] [3 L] [2 L]
[10 dm³] [30 dm³] [20 dm³]
47. (B) For NH_3
 $22.4 \text{ dm}^3 = 1 \text{ mol}$
 $= 6.022 \times 10^{22}$ molecules
 $= 6.022 \times 10^{22} \times 4$ atoms
 $\therefore 2.24 \text{ dm}^3 = 0.1 \text{ mol}$
 $= 0.6022 \times 10^{22}$ molecules
 $= 0.6022 \times 10^{22} \times 4$ atoms
 $= 2.4088 \times 10^{23}$ atoms
48. (B) $2KClO_3 \longrightarrow 2KCl + 3O_2 \uparrow$
[2 moles] [3 moles]
2 moles of $KClO_3 = 2 \times 122.5 = 245 \text{ g}$
3 moles of O_2 at STP occupy = $(3 \times 22.4 \text{ dm}^3)$
Thus, 245 g of potassium chlorate will liberate 67.2 dm^3 of oxygen gas.
Let 'x' gram of $KClO_3$ liberate 22.4 dm^3 of oxygen gas at S.T.P.
 $\therefore x = \frac{245 \times 22.4}{3 \times 22.4} = 81.67 \text{ g}$
49. (B) 1 mol $N_2 = 6.022 \times 10^{23}$ molecules
At STP, 1 mol $N_2 = 22.4 \text{ dm}^3 = 22.4 \times 10^3 \text{ cm}^3$
 $\therefore 22.4 \times 10^3 \text{ cm}^3 = 6.022 \times 10^{23}$ molecules
 $\therefore 22.4 \text{ cm}^3 = 6.022 \times 10^{20}$ molecules
50. (B) Atomic mass is the mass of an atom of the element.
Mass of 1 atom of the element = 10 u
Now, $1 \text{ u} = 1.66056 \times 10^{-24} \text{ g}$
Therefore, $10 \text{ u} = 1.66056 \times 10^{-23} \text{ g}$
51. (C)



52. (B) $C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$
 $1 \text{ mol C} \equiv 1 \text{ mol CO}_{2(g)}$
 $6 \text{ g C} = 0.5 \text{ mol C}$
 $\therefore 0.5 \text{ mol C} \equiv 0.5 \text{ mol CO}_{2(g)}$
 At STP, $1 \text{ mol CO}_{2(g)} \equiv 22.4 \text{ dm}^3$
 $\therefore 0.5 \text{ mol CO}_{2(g)} = 11.2 \text{ dm}^3$
53. (D) Molar mass of methane (CH_4) = 16 g mol^{-1}
 \therefore No. of moles of $CH_4 = \frac{3.2}{16} = 0.2 \text{ mol}$
 No. of atoms in a molecule of $CH_4 = 5$
 \therefore Moles of atoms in $0.2 \text{ mol CH}_4 = 0.2 \times 5 = 1 \text{ mol}$
54. (C) Molecular mass ethane (C_2H_6) = 30 g mol^{-1}
 $1 \text{ mol of ethane} = 30 \text{ g} = 22.4 \text{ dm}^3$ at STP
 $\therefore 60 \text{ g ethane} = 22.4 \text{ dm}^3 \times 2 = 44.8 \text{ dm}^3$ of C_2H_6
55. (C) $^{\circ}F = \left(^{\circ}C \times \frac{9}{5}\right) + 32 = -40.0 \times \frac{9}{5} + 32$
 $= -72 + 32$
 $= -40.0^{\circ}F$
56. (D)
57. (B) Chemical formula of ammonium nitrate is NH_4NO_3 .
 $1 \text{ mol NH}_4NO_3 \equiv 2 \text{ mol N-atoms}$
 $\therefore 80 \text{ g NH}_4NO_3 = 2 \text{ mol N-atoms}$
 $\therefore 8 \text{ g NH}_4NO_3 = 0.2 \text{ mol N-atoms}$
58. (A) $4.25 \text{ g NH}_3 = \frac{4.25}{17} \text{ mol NH}_3 = 0.25 \text{ mol NH}_3$
 $1 \text{ mol NH}_3 = 1 \text{ mol N-atoms} + 3 \text{ mol H-atoms}$
 $\therefore 0.25 \text{ mol NH}_3 = 0.25 \times 4$
 $= 1 \text{ mol atoms}$
59. (D) $^{\circ}F = \frac{9}{5} (^{\circ}C) + 32$
 $= \frac{9}{5} (50) + 32 = 90 + 32 = 122^{\circ}F$
60. (B) Number of moles of a gas (n)
 $= \frac{\text{Volume of gas at STP}}{22.4 \text{ dm}^3 \text{ mol}^{-1}}$

- $\therefore n = \frac{4.48 \text{ dm}^3}{22.4 \text{ dm}^3 \text{ mol}^{-1}} = 0.2 \text{ mol}$
- $Mg_{(s)} + 2HCl_{(aq)} \longrightarrow MgCl_2 + H_{2(g)} \uparrow$
 $1 \text{ mol Mg} \equiv 1 \text{ mol H}_2 \text{ gas}$
 \therefore Mg required to liberate $0.2 \text{ mol H}_2 \text{ gas}$
 $= 0.2 \text{ mol} = 0.2 \times 24 = 4.8 \text{ g}$
61. (A) Structure of methoxymethane:
 $CH_3 - O - CH_3$
 Its molecular formula is C_2H_6O .
 Molar mass = 46 g mol^{-1}
 \therefore No. of moles of $C_2H_6O = \frac{46 \text{ g}}{46 \text{ g mol}^{-1}}$
 $= 1 \text{ mol}$
 One molecule of C_2H_6O contains 2 C-atoms and 6 H-atoms.
 $\therefore 1 \text{ mol C}_2H_6O$ contains 2 mol C-atoms and 6 mol H-atoms.
62. (B) Number of atoms
 $= \text{Number of moles} \times \text{Avogadro's constant}$
 $= \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}} \times 6.022 \times 10^{23} \text{ atoms/mol}$
 Since, the mass is same for all elements, the number of atoms will be inversely proportional to atomic mass. Among the given, Na has the smallest atomic mass.
63. (D) $CaCO_{3(s)} \longrightarrow CaO_{(s)} \longrightarrow CO_{2(g)}$

[1 mol]	[1 mol]	[1 mol]
[100 g]	[56 g]	[44 g]
[10 g]	[5.6 g]	[4.4 g]
64. (C)
65. (B) Number of moles (n)
 $= \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}} = \frac{100 \text{ g}}{40 \text{ g mol}^{-1}} = 2.5 \text{ mol}$
 Number of molecules
 $= \text{Number of moles} \times \text{Avogadro's constant}$
 $= 2.5 \text{ mol} \times 6.022 \times 10^{23} \text{ molecules/mol}$
 $= 1.5055 \times 10^{24}$

◆ ◆ ◆ Evaluation Test ◆ ◆ ◆

1. (C)
2. (C) Molecular mass of $N_2O_4 = 28 + 64 = 92 \text{ g}$
 \therefore number of moles = $\frac{54}{92} = 0.59 \text{ moles}$
 Molecular mass of $CO_2 = 12 + 32 = 44 \text{ g}$
 \therefore number of moles = $\frac{28}{44} = 0.64 \text{ moles}$
 Molecular mass of $H_2O = 2 + 16 = 18 \text{ g}$
- \therefore number of moles = $\frac{36}{18} = 2 \text{ moles}$
 Molecular mass of $C_2H_5OH = 24 + 6 + 16 = 46$
 \therefore number of moles = $\frac{46}{46} = 1 \text{ mole}$
 Among the given, water has more moles.
 \therefore Largest number of molecules is present in 36 g of water.



3. (A) Mass of 6×10^{23} molecules of water = 18 g
 Mass of 1 molecule of water = $\frac{18}{6 \times 10^{23}}$
 $= 3 \times 10^{-23}$ g
 $= 3 \times 10^{-26}$ kg.
4. (A)
5. (B) Mass of 1 mole of Ne = 20 g
 Number of moles = $\frac{\text{Mass of a substance}}{\text{Molar mass of the substance}}$
 $= \frac{52 \text{ g}}{20 \text{ g mol}^{-1}} = 2.6 \text{ mol}$
6. (B)
7. (B) Molar mass of NaOH = g mol⁻¹
 $\therefore 1 \text{ mol NaOH} = 40 \text{ g}$
 $0.01 \text{ mol NaOH} = 0.4 \text{ g}$
8. (D) 1 molecule of PCl₃ \equiv 4 atoms
 $\therefore 1 \text{ mole i.e., Avogadro number (N}_A\text{) of PCl}_3\text{ molecules will contain } 4 \times N_A \text{ atoms.}$
 $\therefore 1.4 \text{ moles of PCl}_3 = 4 \times 1.4 \times N_A \text{ atoms}$
 $= 3.372 \times 10^{24} \text{ atoms}$
9. (A) $14 \text{ g of CO} = \frac{14}{28} = 0.5 \text{ mole.}$
 1 mole of CO occupies 22.4 L at NTP
 $\therefore 0.5 \text{ mole will occupy } 11.2 \text{ L}$
10. (D) 1 molecule of CO contains 1 oxygen atom.
 $\therefore 6.02 \times 10^{24} \text{ CO molecules contain } 6.02 \times 10^{24} \text{ oxygen atoms.}$
11. (C) 12. (C)
13. (D) The SI unit of mass is 'kilogram' and not 'gram'.
14. (B) At S.T.P,
 $22.4 \text{ dm}^3 \text{ of any gas} \equiv 6.022 \times 10^{23} \text{ molecules}$
15. (D) In first experiment:
 $2.70 \text{ g of copper oxide contain } 2.16 \text{ g of copper.}$
 $\% \text{ of copper} = 2.16 / 2.70 \times 100 = 80 \%$
 In second experiment:
 $1.83 \text{ g of copper oxide contain } 1.46 \text{ g of copper.}$
 $\% \text{ of copper} = 1.46 / 1.83 \times 100 = 79.8 \%$
 Therefore, percentage of copper in copper oxide is approximately 80 %.
 Since the percentage of copper in both the sample of copper oxide is nearly same, the above data illustrates the law of definite proportion.
16. (D) 17. (B) 18. (A)
19. (B) $\text{BaCl}_2 + \text{H}_2\text{SO}_4 \longrightarrow 2\text{HCl} + \text{BaSO}_4$
 $\therefore 20.8 + 9.8 = 7.3 + x$
 $\therefore x = 23.3 \text{ g}$
20. (A) 21. (C) 22. (A)
23. (D)
24. (D) Atomic mass of the given element
 $= 6.022 \times 10^{23} \times 10.86 \times 10^{-26} \text{ kg}$
 $= 65.4 \times 10^{-3} \text{ kg}$
 $= 65.4 \text{ g}$
 \therefore The element whose atom has mass of $10.86 \times 10^{-26} \text{ kg}$ is zinc.
25. (A) Contribution of $^{10}\text{B} = 10.0 \times 0.19$
 $= 1.9 \text{ amu} \dots(\text{i})$
 Contribution of $^{11}\text{B} = 11.0 \times 0.81$
 $= 8.91 \text{ amu} \dots(\text{ii})$
 Adding (i) and (ii) = $1.9 + 8.91 = 10.81 \text{ amu}$
 Thus, the average atomic mass of boron is 10.81 amu.
26. (D) 27. (D)
28. (A) N³⁻ ion has 8 valence electrons.
 $14 \text{ g N}^{3-} \text{ ions have } 8N_A \text{ valence electrons}$
 $\therefore 4.2 \text{ g of N}^{3-} \text{ ions have valence electrons}$
 $= \frac{8N_A \times 4.2}{14} = 2.4N_A$
29. (A) $16 \text{ g of O}_2 \text{ has number of moles} = \frac{16}{32} = \frac{1}{2}$
 $14 \text{ g of N}_2 \text{ has number of moles} = \frac{14}{28} = \frac{1}{2}$
 Number of moles is same, so number of molecules are same.
30. (B) Molecular mass of BaCl₂.2H₂O
 $= 137 + 35.5 \times 2 + 2 \times 18 = 244 \text{ g}$
 $244 \text{ g of BaCl}_2.2\text{H}_2\text{O} = 2 \text{ moles of water}$
 $\therefore 488 \text{ g of BaCl}_2.2\text{H}_2\text{O} = \frac{488 \times 2}{244}$
 $= 4 \text{ moles of water}$
31. (A)
32. (A) $1 \text{ cm}^3 = 0.001 \text{ L}$
 $\therefore 11.2 \text{ cm}^3 = 0.001 \times 11.2 = 0.0112 \text{ L}$
 $22.4 \text{ L of gas at STP} = 1 \text{ mol}$
 \therefore Number of moles in 11.2 cm^3 of H₂ is
 $= \frac{0.0112}{22.4} = 0.0005 \text{ mol}$
33. (A) Helium atom has 2 electrons.
 $\therefore 1 \text{ mol He} \equiv 2 N_A \text{ electrons}$
 $\therefore 2 \text{ mol He} \equiv 4 N_A \text{ electrons}$
34. (D) 1 mole of K₄[Fe(CN)₆] = 12 \times 6 g of carbon
 $\therefore 0.5 \text{ mole of K}_4[\text{Fe}(\text{CN})_6] = \frac{0.5 \times 12 \times 6}{1} \text{ g}$
 $= 36 \text{ g of carbon}$
35. (B)