# CHEMISTRY (CODE - 043)

## CHAPTER TEST - STEP-BY-STEP SOLUTIONS AND MARKING SCHEME

**Class XI (2026-27) | Chapter 1: Some Basic Concepts of Chemistry**

**Max. Marks: 35**

### SECTION A (12 × 1 = 12 marks)

*Each correct answer carries 1 mark.*

**Q1. Ans: A. 4 and 5 [1]**

0.04050: the leading zeros only fix the decimal point (not significant); 4, 0, 5 and the final 0 after the decimal point are significant → **4**.

2.0034: zeros between non-zero digits are significant → **5**.

**Q2. Ans: C. P – sugar solution; Q – mixture of salt and sugar; R – copper; S – ammonia [1]**

X = mixtures (separable by physical methods); Y = pure substances (fixed composition).

P = homogeneous mixture (uniform) → sugar solution. Q = heterogeneous mixture (non-uniform) → mixture of salt and sugar.

R = element (only one type of atom) → copper. S = compound (different atoms in a fixed ratio) → ammonia.

**Q3. Ans: B. 9 g [1]**

Moles = $3.011 \times 10^{23} \div 6.022 \times 10^{23} = 0.5$ mol

Mass = 0.5 mol × 18 g $mol^{-1}$ = **9 g**

**Q4. Ans: D. 298.15 K [1]**

From the figure, the Fahrenheit thermometer reads 77 °F (the same level on the Celsius thermometer).

°C = (°F − 32) × 5/9 = (77 − 32) × 5/9 = 25 °C

K = 25 + 273.15 = **298.15 K**

**Q5. Ans: C. $2.5 \times 10^{-4}$ $m^3$ [1]**

250 mL = 250 $cm^3$; 1 $cm^3 = 10^{-6}$ $m^3$

$250 \times 10^{-6}$ $m^3$ = **$2.5 \times 10^{-4}$ $m^3$**

**Q6. Ans: A. Law of Multiple Proportions [1]**

Mass of carbon is fixed (3 g). The masses of oxygen that combine with it are 4 g and 8 g.

Ratio = 4 : 8 = 1 : 2, a simple whole-number ratio → Law of Multiple Proportions.

**Q7. Ans: B. $B_2$ is the limiting reagent [1]**

1 atom of A needs 1 molecule of $B_2$. 200 atoms of A would need 200 molecules of $B_2$, but only 100 are available. So $B_2$ is consumed first and limits the product.

**Q8. Ans: D. $1.2 \times 10^{24}$ [1]**

Each $NH_3$ molecule has 1 N + 3 H = 4 atoms.

Atoms = $0.5 \times 4 \times 6.022 \times 10^{23} = 1.2044 \times 10^{24} \approx$ **$1.2 \times 10^{24}$**

**Q9. Ans: C. 0.2 m [1]**

Moles of NaCl = 5.85 ÷ 58.5 = 0.1 mol; mass of solvent = 500 g = 0.5 kg

Molality = 0.1 ÷ 0.5 = **0.2 m**

**Q10. Ans: A. Both A and R are true, and R is the correct explanation of A. [1]**

Molality = moles of solute ÷ mass of solvent (kg); mass does not change with temperature. Molarity uses volume of solution, which changes with temperature. R correctly explains A.

**Q11. Ans: C. A is true but R is false. [1]**

Chlorine has two isotopes, $^{35}Cl$ (75.77%) and $^{37}Cl$ (24.23%); its average atomic mass is about 35.5 u, so A is true. R is false because chlorine is not a single isotope.

**Q12. Ans: D. A is false but R is true. [1]**

Terminal zeros are significant only if they are on the right of a decimal point. 100 (no decimal point) has only one significant figure, so A is false. R is a correct rule.

## SECTION B (4 × 2 = 8 marks)

**Q13. [2]**

**(i) Precision** is the closeness of various measurements of the same quantity to each other. **Accuracy** is the agreement of a measured value with the true value. [1]

**(ii)** The centre of the board is the true value. [1]

**A**: the four results lie close together but far from the centre → **precise but not accurate**.

**B**: the results are scattered and away from the centre → **neither precise nor accurate**.

**C**: the results lie close together at the centre → **both precise and accurate**.

**Q14. [2]**

**(i)** 0.0000786 = **$7.86 \times 10^{-5}$** (decimal moved 5 places to the right). Leading zeros are not significant, so there are **3 significant figures**. [1]

**(ii)** 6.35 + 1.2 + 0.042 = 7.592. The result cannot have more digits after the decimal point than the number with the fewest (1.2 has one). Rounded: **7.6**. [1]

**Q15. [2]**

Average atomic mass = Σ (fractional abundance × atomic mass) [½]

= (0.6915 × 62.93 u) + (0.3085 × 64.93 u) [½]

= 43.52 u + 20.03 u [½]

= **63.55 u** [½]

**Q16. [2]**

**(i)** From the figure, the volumes $H_2 : O_2 : H_2O$ (vapour) = 2 : 1 : 2, a simple whole-number ratio. This illustrates **Gay-Lussac's Law of Gaseous Volumes**. [½]

**Statement:** When gases combine or are produced in a chemical reaction, they do so in a simple ratio by volume, provided all gases are at the same temperature and pressure. [½]

**(ii)** $H_2 : O_2 : H_2O = 2 : 1 : 2$ [½]

Oxygen needed = 60 × 1/2 = **30 mL**; water vapour formed = 60 × 2/2 = **60 mL** [½]

**OR**

For a fixed mass (32 g) of sulphur, the masses of oxygen are 32 g and 48 g. Ratio = 32 : 48 = **2 : 3**, a simple whole-number ratio, so the Law of Multiple Proportions is illustrated. [1]

**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. [1]

## SECTION C (2 × 3 = 6 marks)

**Q17. [3]**

**Step 1 – Moles of each element** (take a 100 g sample: C = 40.00 g, H = 6.71 g, O = 53.29 g)

C = 40.00 ÷ 12.01 = 3.331 mol

H = 6.71 ÷ 1.008 = 6.657 mol

O = 53.29 ÷ 16.00 = 3.331 mol [1]

**Step 2 – Simplest ratio** (divide by the smallest value, 3.331)

C : H : O = 1 : 2 : 1 → **Empirical formula = $\mathbf{CH_2O}$** [1]

**Step 3 – Molecular formula**

Empirical formula mass = 12.01 + 2(1.008) + 16.00 = 30.03 g

n = molar mass ÷ empirical formula mass = 180 ÷ 30.03 ≈ 6

**Molecular formula = $\mathbf{(CH_2O)_6 = C_6H_{12}O_6}$** [1]

**Q18. [3]**

Moles of Al = 5.4 ÷ 27 = 0.2 mol; moles of $Cl_2$ = 14.2 ÷ 71 = 0.2 mol

**(i)** From the equation, 2 mol Al needs 3 mol $Cl_2$. So 0.2 mol Al needs 0.2 × 3/2 = 0.3 mol $Cl_2$, but only 0.2 mol is available. **$\mathbf{Cl_2}$ is the limiting reagent.** [1]

**(ii)** 3 mol $Cl_2$ gives 2 mol $AlCl_3$, so 0.2 mol $Cl_2$ gives 0.2 × 2/3 = 0.133 mol $AlCl_3$.

Molar mass of $AlCl_3$ = 27 + 3(35.5) = 133.5 g $mol^{-1}$

Mass of $AlCl_3$ = 0.1333 × 133.5 = **17.8 g** [1]

**(iii)** Al used = 0.1333 mol × 27 = 3.6 g

Al left = 5.4 − 3.6 = **1.8 g** [1]

*Check:* 3.6 g Al + 14.2 g $Cl_2$ = 17.8 g $AlCl_3$ ✓

## SECTION D (1 × 4 = 4 marks)

**Q19. [4]**

Mass of solution = 90 + 500 = 590 g; moles of glucose = 90 ÷ 180 = 0.5 mol

**(i)** Mass per cent = (90 ÷ 590) × 100 = **15.25 % (≈ 15.3 %)** [1]

**(ii)** Molality = 0.5 mol ÷ 0.500 kg = **1.00 m** [1]

**(iii)** Volume of solution = mass ÷ density = 590 g ÷ 1.18 g $mL^{-1}$ = 500 mL = 0.500 L

Molarity = 0.5 mol ÷ 0.500 L = **1.00 M** [1]

**(iv) Molarity** would change (decrease). It depends on the volume of the solution, which changes with temperature. Molality depends only on masses, which do not change with temperature. [1]

## SECTION E (1 × 5 = 5 marks)

**Q20. [5]**

**(i) Balancing step by step** [2]

Unbalanced: $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$

*Balance C:* 3 C on the left, so 3 $CO_2$: $C_3H_8 + O_2 \rightarrow 3CO_2 + H_2O$

*Balance H:* 8 H on the left, so 4 $H_2O$: $C_3H_8 + O_2 \rightarrow 3CO_2 + 4H_2O$

*Balance O:* right side has 3 × 2 + 4 × 1 = 10 O atoms, so 5 $O_2$.

**$C_3H_8$ (g) + 5 $O_2$ (g) → 3 $CO_2$ (g) + 4 $H_2O$ (l)**

Check: C = 3, H = 8, O = 10 on each side ✓

**(ii)** Molar mass of $C_3H_8$ = 3(12) + 8(1) = 44 g $mol^{-1}$; 44 g = 1 mol. [2]

1 mol $C_3H_8$ gives 3 mol $CO_2$ → 3 × 44 = **132 g $CO_2$**

1 mol $C_3H_8$ gives 4 mol $H_2O$ → 4 × 18 = **72 g $H_2O$**

**(iii)** Oxygen used = 5 mol × 32 = 160 g. [1]

Mass of reactants = 44 + 160 = 204 g

Mass of products = 132 + 72 = 204 g

Total mass of reactants = total mass of products, so the Law of Conservation of Mass is obeyed.

**OR**

**(i)** Moles = 49 ÷ 98 = **0.5 mol** [1]

**(ii)** Molecules = 0.5 × 6.022 × $10^{23}$ = **3.011 × $10^{23}$** [1]

**(iii)** Each $H_2SO_4$ molecule has 2 H, 1 S and 4 O atoms. [2]

H atoms = 2 × 3.011 × $10^{23}$ = **6.022 × $10^{23}$**

S atoms = 1 × 3.011 × $10^{23}$ = **3.011 × $10^{23}$**

O atoms = 4 × 3.011 × $10^{23}$ = **1.2044 × $10^{24}$**

**(iv)** Moles of O atoms = 0.5 × 4 = 2 mol; mass of oxygen = 2 × 16 = **32 g** [1]

***** END OF SOLUTIONS *****