Some Basic Concepts of Chemistry

Standard difficulty · 20 questions · 40 marks · Time: 36 minutes

Name: _____ Date: _____ Score: ____ / 40

Show your working for Sections B and C. Answers are on the last page — try every question first!

Section A — Choose the correct answer — 8 × 1 = 8 marks

1. One mole of a substance contains:

(a) 22.7 particles **(b)** 1 gram of particles

(c) 6.022×10^{22} particles **(d)** 6.022×10^{23} particles

2. How many significant figures are there in 90007?

(a) 5 **(b)** 6 **(c)** 4 **(d)** 7

3. How many moles are there in 18.5 g of $Ca(OH)_2$? (Atomic masses: H = 1 u, O = 16 u, Ca = 40 u)

(a) 0.125 mol **(b)** 0.5 mol **(c)** 1 mol **(d)** 0.25 mol

4. How many atoms in total are there in 4 mol of CH_4? ($N_A = 6.022 \times 10^{23}$ mol^{-1})

(a) 1.204×10^{25} **(b)** 2.409×10^{24} **(c)** 1.445×10^{25} **(d)** 4.818×10^{24}

5. 18 g of glucose ($C_6H_{12}O_6$, molar mass 180 g/mol) is dissolved in water to make 250 mL of solution. Its molarity is:

(a) 0.4 M **(b)** 0.8 M **(c)** 4 M **(d)** 0.04 M

6. The percentage by mass of O in Na_2CO_3 is: (Atomic masses: C = 12 u, O = 16 u, Na = 23 u)

(a) 54.72% **(b)** 15.09% **(c)** 43.40% **(d)** 45.28%

7. The empirical formula of C_6H_6 is:

(a) CH **(b)** C_3H_3 **(c)** C_6H_6 **(d)** C_2H_2

8. Assertion (A): 44 g of CO_2 and 30 g of C_2H_6 contain the same number of molecules. (Atomic masses: H = 1 u, C = 12 u, O = 16 u)
Reason (R): Equal numbers of moles of any substances contain equal numbers of molecules.

(a) Both A and R are true, and R is the correct explanation of A.

(b) Both A and R are true, but R is not the correct explanation of A.

(c) A is true, but R is false.

(d) A is false, but R is true.

Section B — Short answer — 8 × 2 = 16 marks

9. State the law of multiple proportions with an example.

10. How many moles and how many atoms are there in 310.5 g of lead? (Pb = 207 u; N_A = 6.022×10^{23} mol^{-1})

11. 27.38 g of hydrogen chloride (HCl, molar mass 36.5 g/mol) is dissolved in 500 g of water. Find the molality of the solution.

12. 230 g of ethanol (C_2H_5OH) is mixed with 90 g of water. Find the mole fraction of each. (Molar masses: C_2H_5OH 46 g/mol, water 18 g/mol)

13. A compound contains P 43.66%, O 56.34% by mass. Find its empirical formula. (Atomic masses: O = 16 u, P = 31 u)

14. What volume of water must be added to 20 mL of 2 M nitric acid (HNO_3) to dilute it to 0.1 M?

15. What mass of copper sulphate ($CuSO_4$, molar mass 159.5 g/mol) is needed to make 50 mL of a 1 M solution?

16. A compound has the empirical formula CH_2 and a molar mass of 84 g/mol. Find its molecular formula. (Atomic masses: H = 1 u, C = 12 u)

Section C — Word problems — 4 × 4 = 16 marks

17. State the laws of chemical combination. Explain the law of definite proportions with an example.

18. 56 g of Fe reacts with 64 g of S by Fe + S → FeS. (a) Which is the limiting reagent? (b) What mass of FeS forms? (c) What mass of the other reactant is left over? (Atomic masses: S = 32 u, Fe = 56 u)

19. 58 g of butane (C_4H_{10}) burns completely in oxygen. (a) Write the balanced equation. (b) Find the masses of carbon dioxide and water formed. (c) What mass of oxygen is used? (Atomic masses: H = 1 u, C = 12 u, O = 16 u)

20. A solution of silver nitrate ($AgNO_3$, molar mass 170 g/mol) is 30% by mass and has a density of 1.4 g/mL. Find (a) its molarity, (b) its molality. (2 decimal places where needed.)

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Answer key

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Q	Answer	Working
1	**(d) 6.022×10^{23} particles**	the Avogadro constant
2	**(a) 5**	non-zero digits and zeros between them count; leading zeros never count; trailing zeros count after a decimal point
3	**(d) 0.25 mol**	n = mass ÷ molar mass = 18.5 ÷ 74
4	**(a) 1.204×10^{25}**	molecules = 4 × N_A; each CH_4 has 5 atoms
5	**(a) 0.4 M**	moles = 18 ÷ 180 = 0.1; ÷ 0.25 L
6	**(d) 45.28%**	3 × 16 ÷ 106 × 100
7	**(a) CH**	divide every subscript by their highest common factor
8	**(a) Both A and R are true, and R is the correct explanation of A.**	A is true: both are 1 mol. R is true and explains A.
9	**When two elements form more than one compound, the masses of one element that combine with a fixed mass of the other are in a simple whole-number ratio. Example: 12 g of carbon combines with 16 g of oxygen in CO and 32 g in CO_2, a ratio of 1 : 2.**	—
10	**1.5 mol; 9.033×10^{23} atoms**	n = mass ÷ molar mass; atoms = n × N_A
11	**1.5 mol/kg**	moles = 0.75; molality = moles ÷ 0.5 kg
12	**x(ethanol) = 0.5; x(water) = 0.5**	moles: 5 and 5; total 10
13	**P_2O_5**	divide each percentage by the atomic mass, then by the smallest result, and round to whole numbers
14	**380 mL (final volume 400 mL)**	$M_1V_1 = M_2V_2$
15	**7.975 g**	moles = 1 × 0.05
16	**C_6H_{12}**	n = 84 ÷ 14 = 6
17	**Laws: conservation of mass; definite proportions; multiple proportions; Gay-Lussac's law of gaseous volumes; Avogadro's law. Definite proportions: a compound always contains the same elements combined in the same proportion by mass, whatever its source. Example: pure water from any source has hydrogen and oxygen in the ratio 1 : 8 by mass.**	—
18	**(a) Fe; (b) 88 g; (c) 32 g of S**	moles ÷ coefficient: Fe 1, S 2; the smaller is limiting

Q	Answer	Working
19	**(a) $2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$; (b) 176 g CO_2, 90 g H_2O; (c) 208 g O_2**	1 mol of C_4H_{10} reacts; use the mole ratio 2 : 13 : 8 : 10
20	**(a) 2.47 M; (b) 2.52 mol/kg**	100 g of solution has 30 g solute (0.18 mol) and 70 g water; its volume is 100 ÷ 1.4 mL