Indian School Al Wadi Al Kabir
Assessment – I (2025-2026)

Class: XI | Subject: Chemistry(043) | Max. marks: 70
Date: 16/09/2025 | Set- I-MS | Time: 3 Hours

	SECTION A	
1	B. 27.27 %	1
2	A. $mvr = nh/2\pi$	1
3	C. 4th period and 5th group	1
4	A. Carbon = -1, Oxygen = +1	1
5	D. 0.5 M	1
6	A. $\Delta E = h\upsilon$	1
7	C. F > N > C > B > Si	1
8	C. NO^+	1
9	A. 1.8066×10^{23} atoms.	1
10	C. Magnetic quantum number	1
11	D. Ununquadium	1
12	C. 2	1
13	B. Both A and R are true, and R is not the correct explanation of A.	1
14	D. A is false but R is true	1
15	A. Both A and R are true, and R is the correct explanation of A.	1
16	B. Both A and R are true, and R is not the correct explanation of A.	1
17	A. I. One atomic mass unit is defined as a mass exactly equal. to one-twelfth of the mass of one carbon 12 atom. II. If a substance 'A' dissolves in substance 'B' and their number of moles are n_A and n_B, respectively, then the mole fractions of A and B are given as: Mole fraction of A $= \frac{\text{No. of moles of A}}{\text{No. of moles of solutions}} = \frac{n_A}{n_A + n_B}$ Mole fraction of B $= \frac{\text{No. of moles of B}}{\text{No. of moles of solutions}} = \frac{n_B}{n_A + n_B}$ Thus, total mole fractions = 1 B. I. One mole is the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 g (or 0.012 kg) of the $_{12}C$ isotope. II.	1×2

	$\text{Molarity (M)} = \dfrac{\text{No. of moles of solute}}{\text{Volume of solution in litres}}$ Volume changes with temperature.	
18	I. Diagram with labelling II. Diagram with labelling	1×2
19	I. Na attains a stable noble gas configuration after losing the first outermost electron, hence it is difficult and requires more IE_2. Mg has its outermost electron in 3s orbital after first ionization, and removing this second electron leads to a stable noble gas configuration, hence less IE_2. II. Anomalous behaviour. Reasons: Small size, large charge to radius ratio, high electronegativity/high ionization enthalpy, absence of d orbitals	1×2
20	Any two limitations with eg if any.	1×2
21	$\lambda = \dfrac{h}{mv}$ $\lambda = \dfrac{6.626 \times 10^{-34}\ \text{Js}}{(0.2\ \text{kg}) \times (15\ \text{m/s})}$. $= 2.2 \times 10^{-34}$ m	½ ½ 1
22	I. 10 Volumes of water vapour (Gay Lussac's law) II. $\text{Moles of ethanol} = \dfrac{23\ \text{g}}{46\ \text{g mol}^{-1}} = 0.5\ \text{mol.}$ $\text{Moles of water} = \dfrac{54\ \text{g}}{18\ \text{g mol}^{-1}} = 3\ \text{mol.}$ $\text{Total moles} = 0.5\ \text{mol} + 3\ \text{mol} = 3.5\ \text{mol.}$ $\text{Mole fraction of ethanol} = \dfrac{0.5\ \text{mol}}{3.5\ \text{mol}} = 0.143$ Mole fraction of water = 1- 0.143 = 0.857	1 ½ ½ ½ ½
23	$\text{Energy} = -2.18 \times 10^{-18}\ Z^2/n^2$ $= -2.18 \times 10^{-18} \times 3^2/1^2$ $= -19.62 \times 10^{-18}$ J $r = a_0 n^2 / Z$ $= 52.9\ \text{pm} \times 1/3$ $= 17.6$ pm	½ ×3 ½ × 3
24	I. Due to the partial double character of C-O bonds due to resonance. II. I being a larger anion has greater polarizability, hence greater covalent nature. III. The covalent radius is half of the distance between two similar atoms **joined by a covalent bond**. Van der Waals radius is half the distance between the nuclei of two **nonbonded atoms**.	1 1 1
25	(a) I. The energy required to remove an electron from an isolated gaseous atom (X) in its ground state.	1×3

	II. Atoms and ions which contain the same number of electrons are called isoelectronic species*. (b) The physical and chemical properties of the elements are periodic functions of their atomic numbers.	
26	$E = 2.18 \times 10^{-18} \times \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$ $= 2.18 \times 10^{-18} \times \left(\frac{1}{4^2} - \frac{1}{2^2}\right)$ $= 2.18 \times 10^{-18} \times (-0.1875)$ $= - 0.408 \times 10^{-18}$ J Energy released = 0.408×10^{-18} J $\upsilon = \Delta E/h$ $= 0.408 \times 10^{-18} / 6.626 \times 10^{-34}$ $= 0.062 \times 10^{16}\ s^{-1}$	½ ½ 1 ½ ½
27	I. $Mg^{2+} < F^- < O^{2-}$ II. Si < S < P III. Br < Cl < F	1×3
28	I. Molarity $= \frac{\text{No: of moles}}{\text{Volume of solution in litres}}$ $= \frac{\text{mass/molar mass}}{\text{Volume}}$ 1 M $= \frac{\text{mass/56}}{\text{1 L}}$ Mass = 56 g II. Moles = Mass/ Molar mass 2 = Mass / 44 Mass = 88 g III. Mass = $1/6.022 \times 10^{23} \times 17$ $= 2.823 \times 10^{-23}$ g IV. Mass% % = Mass of solute/Total mass × 100 5 = Mass/30 × 100 Mass = 1.5 g	1×3

29	I. No. of moles of CaO = mass/molar mass = 10/56 = 0.178 mol No. of moles of $CaCO_3$ = 0.178 mol Mass of $CaCO_3$ = 17.8 g II. No. of moles of $CaCO_3$ = mass/molar mass = 10/100 = 0.1 mol No. of moles of CO_2 = 0.1 mol Volume = 2.24 L **OR** II. A given compound always contains exactly the same proportion of elements by weight. III. 200 g of Ca has a higher no of atoms. No of Ca atom in 200 u = 200/40 = 5 atoms No of Ca atoms in 200 g of Ca = $200/40 \times 6.022 \times 10^{23}$ $= 30.11 \times 10^{23}$ atoms	1 1 ½ ½ 1
30	I. $(n-1)d^{1-10}ns^{0-2}$ **OR** I. No, due to completely filled d orbitals. II. The ionization enthalpy of the extreme left element in a period is the least and the electron gain enthalpy of the element on the extreme right is the highest negative. III. (i) F: High Z_{eff} , small size etc. (ii) S: When an electron is added to O, the added electron goes to the smaller *n* = 2 quantum level and suffers significant repulsion from the other electrons present in this level. For the *n* = 3 quantum level (S), the added electron occupies a larger region of space and the electron-electron repulsion is much less.	1 1 1×2
31	A. I. In beryllium, the electron removed during the ionization is an *s*-electron whereas the electron removed during ionization of boron is a *p*-electron. The penetration of a 2*s*-electron to the nucleus is more than that of a 2*p*-electron; hence the 2*p* electron of boron is more shielded from the nucleus by the inner core of electrons than the 2*s* electrons of beryllium. II. Absence of d orbitals III. The electron has to enter the next higher principal quantum level leading to a very unstable electronic configuration. IV. A cation is smaller than its parent atom because it has fewer electrons while its nuclear charge remains the same. (greater force of attraction. Zeff etc) V. Electronegativity varies depending on the element to which it is bound.	1×5

	OR I. T II. T<R<Q<S<P III. S IV. Acidic oxide V. T	
32	A. I. Half filled configuration of Fe^{3+}. II. In the ground state of the atoms, the orbitals are filled in order of their increasing energies. III. n=3, l=2, ml = -2, ms=-1/2 IV. No n and l cannot be same. L=0 to n-1. V. n =5 **OR** I. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$ II. 4, 4s, 4p, 4d, 4f III. mvr = nh/2Π (angular momentum) 2Πr = nh/mv 2Πr = nλ IV. It states that it is impossible to determine simultaneously, the exact position and exact momentum (or velocity) of an electron. V. 9 electrons	1×5
33	A. I. (i) Moles of H_2 = 200/2=100 moles Moles of O_2 = 100/32= 3.125 moles O_2 is the LR. Moles of H_2O = 6.25 moles Mass of H_2O formed = 112.5 g (ii) No of moles of unreacted H_2 = 93.75 moles. II. Molarity = Mass% × density × 10/Molar mass of solute = 10 × 3.65 ×10/36.5 = 10M **OR** I. $\text{Moles of NaOH} = \frac{\text{Mass of NaOH}}{\text{Molar mass of NaOH}} = \frac{4\text{ g}}{40\text{ g/mol}} = 0.1\text{ mol}$ $\text{Mass of water in kg} = \frac{36\text{ g}}{1000\text{ g/kg}} = 0.036\text{ kg.}$ $\text{Molality} = \frac{\text{Moles of NaOH}}{\text{Mass of water in kg}} = \frac{0.1\text{ mol}}{0.036\text{ kg}} \approx 2.78\text{ mol/kg.}$ $\text{Volume of solution in L} = \frac{100\text{ ml}}{1000\text{ ml/L}} = 0.1\text{ L.}$ $\text{Molarity} = \frac{\text{Moles of NaOH}}{\text{Volume of solution in L}} = \frac{0.1\text{ mol}}{0.1\text{ L}} = 1.0\text{ M.}$	½ ½ ½ ½ 1 2 1 1 1

	II. <table><tr><td></td><td>C</td><td>H</td></tr><tr><td>Mass in g</td><td>86</td><td>14</td></tr><tr><td>Moles</td><td>86/12= 7.16</td><td>14/1=14</td></tr><tr><td>Divide by the smallest value</td><td>1</td><td>2</td></tr></table>EF: CH_2 Empirical formula mass= 14 u n = 28/14 = 2 Molecular formula = C_2H_4	½ ½ ½ ½