CHEMISTRY (CODE - 043)

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

Class XI (2026-27) | Chapter 3: Classification of Elements and Periodicity in Properties

Max. Marks: 35

SECTION A (12 × 1 = 12 marks)

Each correct answer carries 1 mark.

Q1. Ans: A. period 7, group 14 [1]

Period 7 starts at Z = 87 (group 1). Counting on: Z = 104 is in group 4, … Z = 112 in group 12, Z = 113 in group 13, so Z = 114 is in **group 14** (below Pb) of **period 7**.

Q2. Ans: A. S and R [1]

Atomic radius increases down a group and decreases across a period. S is in group 1, period 4 (the lowest and leftmost) → largest radius.

Electronegativity increases across a period and decreases down a group. R (group 17, period 2) is the top-right element shown → highest electronegativity.

Q3. Ans: A. Uue [1]

Digits of 119: 1, 1, 9 → un, un, enn → ununennium; symbol = u + u + e = **Uue**.

Q4. Ans: A. F^- [1]

All four are isoelectronic (10 electrons). The ion with the smallest nuclear charge (F^-, Z = 9) pulls the electrons least, so it has the **largest radius**. Order: $Al^{3+} < Mg^{2+} < Na^+ < F^-$.

Q5. Ans: A. period 6, f-block [1]

The valence shell is n = 6, so the period is 6. The last electron enters the 4f subshell [(n − 2)f], so the element is an f-block (inner transition, lanthanoid) element (Z = 64, Gd).

Q6. Ans: A. (first option) [1]

P: ns^2np^4 → 6 valence electrons → group 16. Q: $(n-1)d^5ns^2$ → d-block, 5 + 2 = 7 → group 7 (Mn). R: ns^1 → group 1. S: ns^2np^5 → 7 valence electrons → group 17.

Q7. Ans: B. Ionisation enthalpy [1]

Ionisation enthalpy increases along a period (Arrow 1) and decreases down a group (Arrow 2). Atomic radius and metallic character show the opposite pattern, and ease of losing electrons follows metallic character.

Q8. Ans: B. Al_2O_3 [1]

Al_2O_3 reacts with both acids and bases → amphoteric. Na_2O is basic, Cl_2O_7 is acidic, CO is neutral.

Q9. Ans: B. Al_2S_3 [1]

Al is in group 13 (valence 3); S is in group 16 (valence 2). Criss-cross: Al_2S_3.

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

Penetration of a 2s electron is greater than that of a 2p electron, so the 2p electron of B is better shielded and is removed more easily than the 2s electron of Be.

Q11. Ans: B. Both A and R are true, and R is not the correct explanation of A. [1]

Electron gain enthalpy of F (−328) is less negative than that of Cl (−349) because the added electron goes to the small n = 2 shell and faces strong electron–electron repulsion. F being the most electronegative element is true, but it is not the reason.

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

Atomic radius **decreases** across a period because the effective nuclear charge increases while electrons are added to the same shell. So A is false and R is true.

SECTION B (4 × 2 = 8 marks)

Q13. [2]

(i) In Be the electron removed is a 2s electron; in B it is a 2p electron. A 2s electron penetrates closer to the nucleus than a 2p electron, so the 2p electron of boron is more shielded by the inner (1s and 2s) electrons and is removed more easily. [1]

(ii) In nitrogen ($2p^3$) the three 2p electrons are in different orbitals. In oxygen ($2p^4$) two electrons share one 2p orbital, so there is extra electron–electron repulsion and the fourth electron is removed more easily. [1]

Q14. [2]

(i) Atoms and ions that contain the same number of electrons are isoelectronic species. Example: O^{2-}, F^-, Na^+, Mg^{2+} (10 electrons each) (any one). [1]

(ii) All have 10 electrons. For the same number of electrons, the larger the nuclear charge the smaller the radius. [1]

$Al^{3+} < Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$

Q15. [2]

(i) Electronegativity **increases** from Na (0.9) to Cl (3.0). Across a period the atomic radius decreases and the effective nuclear charge increases, so the attraction for shared electrons increases. [1]

(ii) Na is the most metallic (lowest electronegativity, loses electrons most easily); **Cl** is the most non-metallic (highest electronegativity, strong tendency to gain electrons). [1]

Q16. [2]

(i) Atomic radius **increases** from Li (152 pm) to Cs (262 pm). Down the group the principal quantum number increases, so the valence electron is farther from the nucleus, and the inner shells shield it from the nuclear pull. [1]

(ii) Cs – it has the largest radius and its outermost electron is the farthest from the nucleus and most shielded, so it is removed most easily. [1]

OR

Digits 1, 2, 0 → un, bi, nil → **unbinilium**, symbol **Ubn**. [1]

It will be placed in **group 2** (alkaline earth metals), **s-block**, with outer configuration **$8s^2$** ([Uuo]$8s^2$). [1]

SECTION C (2 × 3 = 6 marks)

Q17. [3]

(i) Mendeleev's periodic law: the properties of the elements are a periodic function of their atomic weights. [1]

(ii) The predicted atomic weights, densities and formulae of the oxides and chlorides of eka-aluminium and eka-silicon agree closely with those found for gallium and germanium (e.g. 68 vs 70; EO_2 vs GeO_2). This shows that his periodic table could predict the properties of undiscovered elements, and that leaving gaps was justified. [1]

(iii) Modern Periodic Law: the physical and chemical properties of the elements are periodic functions of their **atomic numbers** (Moseley), not atomic weights. [1]

Q18. [3]

(i) Ionisation enthalpy per mole = 2.18×10^{-18} J $\times$ 6.022×10^{23} mol^{-1} = 1.313×10^{6} J mol^{-1} = **1312 kJ mol^{-1}** ($\approx 1.31 \times 10^{3}$). [1]

(ii) Closer to **575** kJ mol^{-1}. The electron removed from Al is a 3p electron, which is more shielded by the 3s electrons and penetrates less than the 3s electron of Mg, so Al has a lower value than Mg. [1]

(iii) Na has one 3s electron outside a neon core, so its first ionisation enthalpy is low. Na^{+} has a stable noble-gas (neon) configuration, so removing a second electron from it is very difficult. Mg has two 3s electrons, so its first value is higher than Na's, but its second electron is also a 3s electron, removed from Mg^{+} without breaking a noble-gas core, so its second value is lower than Na's. [1]

SECTION D (1 × 4 = 4 marks)

Q19. [4]

(i) Element V – very high Δ_iH_1 (2372) and positive $\Delta_{eg}H$ (+48): a noble gas. [1]

(ii) Element II – lowest Δ_iH_1 (419) and a very high Δ_iH_2: an alkali metal that loses one electron very easily. [1]

(iii) Element III – very high Δ_iH_1 and the most negative $\Delta_{eg}H$ (−328): a halogen that gains an electron readily. [1]

(iv) Element VI – low Δ_iH_1 (738) and Δ_iH_2 (1451), with no large jump between them: it loses two electrons easily to give M^{2+} and forms MX_2 (an alkaline earth metal). [1]

SECTION E (1 × 5 = 5 marks)

Q20. [5]

(i) For the fifth period (n = 5), the orbitals filled are 5s, 4d and 5p, with the order of energy 5s < 4d < 5p. [1]

Number of orbitals: 5s = 1, 4d = 5, 5p = 3 → total = 9 orbitals. Each orbital holds 2 electrons, so 9 × 2 = **18 electrons** → 18 elements. [1]

(ii) s-block: **ns^{1-2}** [½]; p-block: **ns^2np^{1-6}** [½]

d-block: **$(n-1)d^{1-10}ns^{0-2}$** [½]; f-block: **$(n-2)f^{1-14}(n-1)d^{0-1}ns^2$** [½]

(iii) n = 4 → **period 4**. The last electrons enter 3d → **d-block**. Group number = 2 + 2 = **4**. (The element is titanium, Z = 22.) [1]

OR

(i) The enthalpy change when an electron is added to a neutral gaseous atom to convert it into a negative ion [$X(g) + e^{-} \rightarrow X^{-}(g)$]. [1]

(ii) In F and O the added electron enters the small n = 2 shell, where there is strong repulsion from the electrons already present. In Cl and S the electron goes to the larger n = 3 shell, where the repulsion is much smaller, so the values are more negative. [2]

(iii) In noble gases the valence shell is completely filled, so the incoming electron has to enter the next higher principal quantum level; this gives a very unstable configuration, so energy has to be supplied. [1]

(iv) Chlorine (−349 kJ mol^{-1}). [1]

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