MCQ 50 marks

1. A. Group 17, Period 3

Explanation: $l = 1 \Rightarrow$ p-subshell, $n = 3 \Rightarrow 3p$. By Hund's rule, the 3p filling order is: first parallel spin ($+\frac{1}{2}$) at $m = -1, 0, +1$, then paired spin ($-\frac{1}{2}$) again at $m = -1, 0, +1$.

Getting $m = 0, s = -\frac{1}{2}$ means this is the fifth electron, i.e., $3p^5$.

So $X = [Ne]3s^2 3p^5 = Cl\,(Z = 17)$ — Group 17, Period 3.

2. D. $Cr\,(Z = 24)$

Explanation: By the normal Aufbau order, 6 valence electrons should give $3d^4 4s^2$, but that leaves 4s full, which the condition rules out. Due to the extra stability of a half-filled 3d, $Cr = [Ar]3d^5 4s^1$ forms as an exception — here 4s is not full ($4s^1$), yet the total valence electrons are $5 + 1 = 6$.

3. A. $n = 2, l = 2, m = 0, s = -\frac{1}{2}$

Explanation: The value of l can range from 0 to $(n - 1)$. For $n = 2$, the maximum $l = 1$; so $l = 2$ (with $n = 2$) is impossible.

4. A. Co, 9 electrons

Explanation: In forming M^{2+}, electrons are usually removed from 4s first.

So the parent metal's configuration is $[Ar]3d^7 4s^2$, which is $Co\,(Z = 27)$.

Therefore, the valence-shell electron count = $7 + 2 = 9$.

5. B. Much higher

Explanation: 3d electrons provide weak screening for 4s electrons. As a result, despite Zn having 10 extra protons, this weak screening means the Z_{eff} on the 4s electron is significantly higher than in Ca — this is exactly why the radius of Zn is much smaller than that of Ca.

6. A. $1s^2 2s^2 2p^4 3s^1$

Explanation: According to the Aufbau principle, an electron cannot go to 3s before 2p is completely full ($2p^6$), so $2p^4 3s^1$ is an excited-state configuration (the ground state would be $2p^5$).

7. C. It has only a single unpaired $4s^1$ **electron**

Explanation: The full $3d^{10}$ subshell effectively shields the $4s^1$ electron from the nucleus's attraction, so the single 4s electron is removed relatively easily.

8. C. $Al\,(Z = 13)$

Explanation: $Al = [Ne]3s^2 3p^1$, with 3 valence electrons.

After removing the first 3, $Al^{3+} = [Ne]$ (noble-gas structure) forms.

The fourth electron must be removed from the stable inner $n = 2$ shell, causing a huge jump in E_4.

9. A. $Z = 26$

Explanation: To get $3d^5$ (maximum unpaired, 5), $M^{3+} = [Ar]3d^5$.

Since forming M^{3+} usually removes $4s^2$ and one 3d electron, the parent $M = [Ar]3d^6 4s^2 = Fe\,(Z = 26)$.

10. B. $5s < 4d < 6s < 4f$

Explanation: $(n + l)$ values — $5s = (5 + 0) = 5$; $4d = (4 + 2) = 6$; $6s = (6 + 0) = 6$; $4f = (4 + 3) = 7$.

4d and 6s have the same $(n + l)$ value (6), so the one with the smaller n (i.e., 4d, $n = 4$) fills first.

Order: $5s < 4d < 6s < 4f$.

11. D. Group 2, oxidation number +2

Explanation: A huge jump is seen from E_2 to E_3 (1145 → 4912, about 4.3 times), meaning a noble-gas structure is achieved after losing 2 electrons.

So the element belongs to Group 2 and will show an oxidation number of +2 (e.g., Ca).

12. D. $Be \rightarrow B$ **and** $N \rightarrow O$

Explanation: $Be(2s^2) \rightarrow B(2s^2 2p^1)$: going from a full $2s^2$ to a newly starting $2p^1$ causes a dip at B.

$N(2p^3) \rightarrow O(2p^4)$: going from a half-filled $2p^3$ to a paired $2p^4$ causes a dip at O due to inter-electronic repulsion.

13. **C.** $P = O, Q = N, R = C, S = F$

Explanation: In order from Group 14 to 17: C, N, O, F.
Normally E_1 would increase steadily, but because of the extra stability of N's half-filled $2p^3$, an exception occurs where $E_1(N) > E_1(O)$
Matching the values: $C = 1086, N = 1402, O = 1314, F = 1681$.
So $P = O, Q = N, R = C, S = F$.

14. **B. 4**

Explanation: A huge jump from E_4 to E_5 means a noble-gas structure is achieved after removing 4 electrons, i.e., it has 4 valence electrons (e.g., C or Si).

15. **D. First Group 2, second Group 1**

Explanation: In the first element, the jump occurs from E_2 to E_3 (Group 2, e.g., Mg).
In the second element, the huge jump occurs right at $E_1 \rightarrow E_2$ (496 → 4562, about 9 times), meaning a noble-gas structure is reached after losing just 1 electron — Group 1 (e.g., Na).

16. **A. P's stability is greater than S's**

Explanation: As with N and O, going from P ($3p^3$, half-filled, stable) to S($3p^4$) in period 3 causes E_1 to slightly decrease due to inter-electronic repulsion, even though Z_{eff} increases with increasing atomic number.

17. **D. There is no jump from $E_1 \rightarrow E_2$**

Explanation: In a Group 1 element (1 valence electron), a huge jump occurs from $E_1 \rightarrow E_2$. But here E_1, E_2, E_3 increase smoothly in sequence — there is no huge jump, which is instead characteristic of a Group 13 element like B ($Z = 5$).

18. **C.** $Na \gg Mg > Al$

Explanation: $Na^+ = [Ne]$ (a fully noble-gas structure); removing an electron from here is extremely difficult, so $E_2(Na)$ is the highest. $Mg^+ = [Ne]3s^1$ and $Al^+ = [Ne]3s^2$ (full 3s, relatively stable) — so $E_2(Al) > E_2(Mg)$, but both are much lower than $E_2(Na)$.

19. **D. Group 15, Period 3 (P)**

Explanation: A huge jump (about 3.4 times) is seen from E_5 to E_6 — a noble-gas structure is achieved after removing 5 electrons, meaning 5 valence electrons.
This is the actual successive ionization energy data for the period-3, Group 15 element P.

20. **A. A completely full electron configuration is extremely stable**

Explanation: Despite the noble gas's measured Van der Waals radius appearing relatively large, its completely full electron structure and high Z_{eff} make its ionization energy the highest in the period.

21. **A. The extra stability of the full $3s^2$ subshell**

Explanation: Alongside the general increase in E_1 from rising Z_{eff}, the extra stability of Mg's full $3s^2$ subshell adds further, so $E_1(Mg)$ rises even more than expected.

22. **C.** Ne

Explanation: A huge jump (4.2 times) from E_3 to E_4 — 3 valence electrons, the element is Al. $Al^{3+} = [Ne]$.

23. **D. 4s; in the cation, 3d becomes lower in energy**

Explanation: In the neutral atom, 4s is slightly lower in energy than 3d, but once 3d electrons are added, they poorly shield 4s, so in the cation the effective energy of 4s becomes higher than 3d — hence in cation formation, 4s electrons are always removed first.

24. **C. Group 1**

Explanation: In Group 1, with only 1 valence electron, an immediate huge jump occurs at $E_1 \rightarrow E_2$; so having the first 7 values with no jump is least consistent with this description.

25. **C. Greater electron repulsion in oxygen's small size**

Explanation: This is analogous to the F/Cl pair — because of its small size, oxygen has higher electron density, causing greater repulsion toward an incoming electron, so S's electron affinity is higher than O's.

26. **B.** Mg (Z = 12)
Explanation: l = 0 ⇒ s-subshell, n = 3 ⇒ 3s. A jump at $E_2 \rightarrow E_3$ means 2 valence electrons — $3s^2$, i.e., Mg.

27. **D. Adding an electron to a negative ion causes Coulombic repulsion**
Explanation: After the first electron is added, the ion is already negatively charged; adding a second electron to this negative ion requires energy input to overcome repulsion between like charges.

28. **C.** N **and the noble gases**
Explanation: N's half-filled $2p^3$ and the noble gases' completely full structure are both reluctant to accept an extra electron, so both have an electron affinity that is zero or positive.

29. **C. Electronegativity is an averaged property, while affinity is an individually measured energy**
Explanation: Pauling-scale electronegativity is a statistical/averaged property derived from bond-energy data, where the small-size-related inter-electronic repulsion of an isolated atom is not directly reflected, but it does directly affect the electron affinity measured in the gaseous state.

30. **A.** O **and** S
Explanation: In both this pair and the analogous F/Cl pair, the smaller second-period element (O or F) shows a lower electron affinity than the corresponding third-period element (S/Cl), contrary to expectation.

31. **B. Lower than both**
Explanation: X = P ($3p^3$), and just like nitrogen, due to the half-filled p^3 configuration, P's electron affinity is lower than its neighbors Si ($3p^2$) and S ($3p^4$).

32. **A. The second step will be strongly positive (endothermic)**
Explanation: This principle is universal, just like $O^- \rightarrow O^{2-}$ — adding a second electron to any negative ion is always an energy-absorbing process.

33. **A.** N, O, F
Explanation: Because of the small size of period-2 elements, all three of N, O, F show electron affinity lower than expected (i.e., lower than the next member of their respective group), which violates the simplified rule 'smaller radius = higher affinity'.

34. **D.** $Mg^{2+} < Na^{+} < F^{-} < O^{2-} < N^{3-}$
Explanation: With the same number of electrons (10), as the proton count decreases (from N = 7 to Mg = 12), the effective nuclear attraction decreases and the radius increases. So N^{3-}, with the fewest protons, is the largest. The mathematical basis for this is the Z/e ratio — it is lowest for N^{3-} and highest for Mg^{2+}.

35. **C. Reluctance to accept a new electron**
Explanation: ns^2np^3 (Group 15) always shows unusually low electron affinity compared to other group members due to the half-filled p subshell — this is true for both N and P.

36. **C. Going down a group, screening increases and electron affinity generally decreases**
Explanation: Going down a group, atomic radius increases and effective nuclear attraction decreases (due to increased screening), so electron affinity generally decreases (Cl > Br > I; the exception occurs only among the small period-2 elements).

37. **B. Contraction due to poor screening by** 4f **electrons**
Explanation: The 4f orbital is diffuse but poorly penetrating in shape, so it provides weak screening both among itself and for outer electrons.
As a result, the Z_{eff} on Hf's valence electrons increases much more than expected, which contracts and offsets the expansion caused by the added shell.

38. **A. Greatest in** Li^+
Explanation: The smaller ion has a higher charge density (z/r), so it holds on to more water molecules more tightly, meaning the effective (hydrated) radius of $Li^{+}_{(aq)}$ increases the most compared to its gaseous radius.

39. **B. Diagonal relationship**
Explanation: When elements diagonally positioned in the periodic table (of periods 2 and 3) have similar charge densities ($\phi = z/r$), they show a deep similarity in chemical behavior (diagonal relationship) — Li/Mg is the classic example.

40. **B. Group 13**
Explanation: Amphoteric oxides (Al_2O_3, ZnO, BeO, PbO) are generally found in the metal–nonmetal borderline groups of the periodic table (such as Group 13) and in certain specific d-block elements.

41. **D. Ne's Van der Waals radius is larger than a covalent radius**
Explanation: Since Ne has no chemical bonds, its Van der Waals radius has to be used, which is inherently larger than a covalent radius due to weak intermolecular forces — this causes a sudden jump in the graph that doesn't actually represent an increase in atomic size.
Because Van der Waals and covalent radii are defined differently and measured from different (weak/strong) interactions, a direct comparison is misleading.

42. **A. In d^0, d^{10}, the radius follows the trend smoothly**
Explanation: In symmetric configurations like $Ca^{2+}(d^0)$, $Zn^{2+}(d^{10})$, with no directional-repulsion-related complications, the radius follows the expected smooth trend; a slight deviation is seen for intermediate d configurations.

43. **C.** $Ca^{2+} < K^+ < Cl^- < S^{2-}$
Explanation: Proton counts: S(16) < Cl(17) < K(19) < Ca(20).
In an isoelectronic series, fewer protons means a larger radius.
So Ca^{2+} (with the most protons) is the smallest.

44. **D. Variability of 4f configuration and inconsistency in screening**
Explanation: Although the overall trend is contraction, some lanthanides have exceptional electron configurations (sharing between 4f and 5d), which makes the rate of contraction not perfectly linear.

45. **B.** Fe^{2+} **is larger**
Explanation: $Fe^{2+} = 3d^6$ and $Fe^{3+} = 3d^5$ — with the same proton count, the electron-poorer Fe^{3+} has a higher Z/e ratio, so it is smaller (hence Fe^{2+} is larger).

46. **B. Metallic radius**
Explanation: In metallic bonding, the electron cloud is more spread out and the bond weaker, so the internuclear distance is greater than in covalent bonding, making the metallic radius value relatively larger.

47. **D. Increasing Z_{eff} makes the M − O bond covalent**
Explanation: A small cation with high charge density strongly polarizes oxygen's electron cloud, forming a covalent bond, which dissolves in water to release H^+, showing acidic character; whereas the low-charge-density Na^+ forms a fully ionic oxide, which is basic.

48. **C. CO and N_2O**
Explanation: CO and N_2O (along with NO, H_2O) are neutral oxides, which react with water to form neither an acid nor a base.

49. **B.** $HClO_4 > HClO_3 > HClO_2 > HClO$
Explanation: As the oxidation number of the central Cl increases, more electron-withdrawing oxygen atoms pull electron density from the O − H bond, making H^+ easier to dissociate and increasing acidity — so $HClO_4$ is the strongest.

D. $HClO_4$ **is stronger, because it has higher**

50. **electronegativity**

Explanation: When oxidation numbers are equal, the size of the central atom becomes decisive — a smaller, more electronegative central atom (Cl) weakens the $O-H$ bond, making H^+ release easier, so $HClO_4 > HIO_4$. Since Cl is smaller than I, it has higher electronegativity and attracts the $O-H$ bond's electron density more strongly.