MCQ 50 marks

1. **D.** P (15)
Explanation: $l = 1 \Rightarrow 3p$. In the first round (parallel spin $+\frac{1}{2}$), the fill order is $m = -1, 0, +1$. With $m = +1, s = +\frac{1}{2}$, this is the third (last parallel-spin) electron, meaning all three orbitals hold one electron each — $3p^3$. So $Z = [Ne]3s^2 3p^3 = P$ ($Z = 15$).

2. **B.** Cu ($Z = 29$)
Explanation: 18 electrons in the M shell means $3s^2 3p^6 3d^{10}$ is completely full. 1 in the N shell means $4s^1$. Total = K(2) + L(8) + M(18) + N(1) = 29 $\Rightarrow$ Cu.

3. **A.** $(3, 2, -1, +\frac{1}{2})$ **and** $(3, 2, -1, -\frac{1}{2})$
Explanation: The same orbital means n, l, m are all identical, with only s opposite — by the Pauli principle this is the only possible combination.

4. **A.** $[Ar]3d^3 4s^2$, V
Explanation: Forming M^{2+} removes the 2 4s electrons first, so M = $[Ar]3d^3 4s^2$ = V ($Z = 23$).

5. **B. Why 4s electrons are removed first**
Explanation: Even though Z_{eff} on 3d might seem low, the universal rule for cation formation (ns removed first) still applies here — $Sc^{3+} = [Ar]$.

6. **C.** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
Explanation: (A) is Ca's correct Aufbau-following configuration. (B) is wrong — an electron can't go to 3s before 2p is full. (C) is impossible — 2p can hold at most 6 electrons, not 7. (D) is wrong — an electron can't go to 2p before 2s is full.

7. **C.** $Z = 24$ **(Cr)**
Explanation: $Cr^{3+} = [Ar]3d^3$ is a very well-known, stable ion (e.g. in $[Cr(H_2O)_6]^{3+}$). Cr's actual ground-state configuration is exceptional ($[Ar]3d^5 4s^1$); forming the +3 ion removes the single $4s^1$ electron plus two more 3d electrons, giving $3d^3$ ($Z = 24$).

8. **D. Group 1**
Explanation: An unusually low E_1 (single valence electron) with a huge jump at $E_1 \rightarrow E_2$ — this is the classic Group 1 (alkali metal) signature.

9. **C.** $6s < 4f < 5d$
Explanation: (n + l): 6s = 6; 4f = 7; 5d = 7. 6s is lowest. 4f and 5d have equal (n + l) (7), so the smaller n (4f, n = 4) fills first. Order: $6s < 4f < 5d$.

10. **C. The filled** $4d^{10}$ **effectively screens** $5s^1$
Explanation: This is fully analogous to Cu's case ($3d^{10}4s^1$) — the same pattern appears throughout Group 11.

11. **A. After removing 5, removal starts from the inner** n = 3 **shell**
Explanation: This is analogous to Al's case (n = 3, 3 valence electrons, jump at E_4) — after removing 5 valence electrons at n = 4 (e.g. As), the jump will occur at E_6.

12. **C. Group 2,** +2
Explanation: A huge jump (8.4×) occurs at $E_2 \rightarrow E_3$ — 2 valence electrons, Group 2 (these are the actual values for Be).

13. **A.** Li = 520, Be = 899, B = 801, C = 1086
Explanation: Li ($1s^2 2s^1$) is lowest — the single $2s^1$ leaves easily: 520. Be($2s^2$ full) is high: 899. B($2s^2 2p^1$) is lower than Be since a new p^1 starts after breaking the filled $2s^2$: 801. C($2p^2$) is highest: 1086.

14. **B. Group 15**
Explanation: No huge jump through 5 successive values means at least 5 valence electrons — Group 15 (like P, where the real jump is at E_6).

15. **A.** Mg $\rightarrow$ Al **and** P $\rightarrow$ S
Explanation: Mg($3s^2$ full) $\rightarrow$ Al($3p^1$ new subshell) and P ($3p^3$ half-filled) $\rightarrow$ S($3p^4$) — these two points show dips analogous to Be $\rightarrow$ B and N $\rightarrow$ O in period 2.

16. **D. An ns^2np^4 configuration; the real jump is at E_7**
Explanation: An ns^2np^4 (Group 16) element may show minor fluctuations, but the main jump occurs at E_7, after all 6 valence electrons are removed.

17. **A. Partly correct — not isoelectronic, but $Na^+ = [Ne]$'s structure is hard to break**
Explanation: $Na^+ = [Ne]$ (10 electrons), $Mg = [Ne]3s^2$ (12 electrons) — so they are not isoelectronic, though it is true that Na's E_2 is higher.

18. **D. $P > S > Si$; stability of P 's half-filled $3p^3$**
Explanation: This is the period-3 version of N, O's pattern — P $(3p^3$, half-filled)'s E_1 exceeds its neighbour S$(3p^4)$'s, even though both exceed Si $(3p^2)$'s.

19. **D. The second has more**
Explanation: In the first set the jump is at $E_5 \rightarrow E_6$ (5 valence electrons, Group 15). In the second, the jump is at $E_6 \rightarrow E_7$ (6 valence electrons, Group 16). So the second has more valence electrons.

20. **C. $K^+ = [Ar]$ but $Ca^+ = [Ar]4s^1$ (easily removed)**
Explanation: K^+ has already reached a noble-gas structure, but Ca^+ still carries one extra 4s electron that comes off comparatively easily.

21. **A. N and Mn**
Explanation: The half-filled-subshell stability principle applies not just to the p-block but the d-block too — Mn $(3d^5)$'s electron affinity is also unusually low/positive compared to its neighbours.

22. **A. At least 4**
Explanation: No jump through the first 4 values means the element has at least 4 valence electrons (this data is actually Si's real values).

23. **D. With the smallest shell and least screening, attraction is greatest**
Explanation: He's only shell, $n = 1$, sits closest to the nucleus, and with no inner shell, screening is minimal — so Z_{eff} and E_1 are both the highest.

24. **D. About 0.85; partial screening**
Explanation: Each electron in the $(n - 1)$ shell provides 0.85 screening, yet once removal moves to the inner shell, Z_{eff} rises considerably, which is the main cause of the huge jump.

25. **D. Group 1**
Explanation: Since Group 1 shows a huge jump at $E_1 \rightarrow E_2$ (the ratio is typically 8-10× or more), a close value like $E_2/E_1 = 1.2$ is least likely for this group.

26. **A. Ca $(Z = 20)$**
Explanation: $l = 0, n = 4 \Rightarrow 4s$. A jump at $E_2 \rightarrow E_3$ means 2 valence electrons $(4s^2)$, i.e. Ca $(Z = 20)$.

27. **D. S > Se (normal downward-decreasing trend)**
Explanation: The exception occurs only for the small period-2 element (O). Since both S and Se are ordinary members, they follow the normal downward trend: S > Se.

28. **C. First exothermic, second endothermic**
Explanation: This principle is universal for any element including O, S, Cl — adding a second electron to a negative ion is always endothermic.

29. **C. Strongly positive (endothermic)**
Explanation: Like O^{2-}, forming S^{2-} has an energy-absorbing second step, but the overall ionic compound formation (aided by lattice energy, e.g. Na_2S) can still be exothermic overall.

30. **D. Higher than both B and N**
Explanation: X = C $(2p^2)$. By the normal trend C > B, and since N has its own exceptionally low affinity, C's affinity can even exceed N's.

31. **B. Two different properties, measured in different contexts**
Explanation: Attraction toward shared electrons in a bond (electronegativity) versus the net energy of adding an electron to an

isolated atom (electron affinity) — these are quantities measured in two different contexts.

32. **A. Group 17 (non-metal) exceeds Group 1**
Explanation: Non-metals readily accept electrons (high electron affinity), while metals readily give them up (low ionization energy, often low/negative electron affinity).

33. **A. By actual values, F > S (328 > 200)**
Explanation: Two opposing effects (group and period trends) partly cancel each other; but due to F's own small-size exception, the actual experimental values give $|E_a(F)| = 328 > |E_a(S)| = 200\,kJmol^{-1}$ — a diagonal comparison shouldn't rely on simple generalization; checking real data matters.

34. **D. It is O; small size causes greater electron repulsion**
Explanation: The period-2, Group 16 element O shows the same small-size-repulsion exception as F (S > O).

35. **B. Generally increases, but an exceptional dip also occurs**
Explanation: It generally increases (as Z_{eff} rises), but an exceptional dip appears right after a half-filled (ns^2np^3) or full (ns^2) subshell. This exception can occur in any period (as with Group 15 elements like N/P), not just period 2.

36. **C. $S^{2-} > Cl^- > K^+ > Ca^{2+}$**
Explanation: Proton counts: S(16) < Cl(17) < K(19) < Ca(20). Fewer protons means a larger radius. So S^{2-} is largest, Ca^{2+} is smallest.

37. **C. Not applicable; there is no 4f series between Sc and Y**
Explanation: The effect of lanthanide contraction appears only for 4d → 5d elements because the 4f series lies between them. There is no f-series between Sc (3d) and Y (4d) — so Y's radius will simply be larger than Sc's, as normally expected.

38. **C. Na^+'s conductivity is lower than K^+'s**
Explanation: Na^+, being smaller, has higher charge density, so it becomes more hydrated in aqueous solution and its effective size grows — so its conductivity is lower than K^+'s.

39. **A. B^{3+} and Si^{4+} have close ionic potential**
Explanation: Just like the diagonal pair Li/Mg, B/Si have similar charge density, so both tend to form covalent oxides and complex halides.

40. **C. Yes, for the same reason (a van der Waals radius is used)**
Explanation: This is a direct repeat of the same principle seen with Ne — every period's noble gas shows the same issue; Ar's van der Waals radius also has to be used, which is larger than Cl's covalent radius.

41. **B. The simple relationship doesn't hold — some H atoms bond directly to P**
Explanation: The simple relationship between oxidation number and acidity doesn't always hold — in H_3PO_3 and H_3PO_2, not all H atoms are in O − H bonds; some bond directly to P (P − H), which isn't ionizable.

42. **B. Both (d^3 and d^8); more electrons sit in t_{2g}**
Explanation: In d^3 and d^8 configurations, electrons sit mainly in the low-energy, ligand-avoiding t_{2g} orbitals, so effective nuclear attraction is felt more strongly, creating local minima in the ionic radius.

43. **B. Cu^+ is larger; Cu^{2+} is more stable via hydration**
Explanation: Although Cu^+'s filled $3d^{10}$ is structurally stable, Cu^{2+}'s higher hydration energy makes it more stable in aqueous solution.

44. **B. Significantly higher (more mass in about the same volume)**
Explanation: Because of lanthanide contraction, elements like Hf, Ta, W keep roughly the same radius as their period-5 counterparts but have nearly double the mass — so density is significantly higher (e.g. Os, Ir are the densest elements).

45. **A. Fe^{2+} (d^6); more deviation from the asymmetric configuration**

Explanation: Mn^{2+} ($3d^5$, half-filled) is perfectly symmetric and follows the expected smooth trend; Fe^{2+} ($3d^6$) is asymmetric (one paired electron in e_g) and shows slightly more crystal-field-driven deviation.

46. **C. The ionic radius applies, and it's smaller than the parent atom**
Explanation: Forming a cation removes electrons and often loses a whole shell, raising the Z/e ratio, so a cation is always smaller than its parent atom.

47. **D.** $PbO + 2HNO_3 \rightarrow Pb(NO_3)_2 + H_2O$ **and** $PbO + 2NaOH \rightarrow Na_2PbO_2 + H_2O$
Explanation: Like Al_2O_3, PbO is amphoteric — it forms salt + water with an acid, and a plumbite salt + water with a base.

48. **B. Yes, the general principle applies to every period**
Explanation: The underlying reason oxide character changes across a period applies equally to any period; the presence of the d-block only adds minor complications to the detailed trend.

49. **D.** SnO **is more basic,** SnO_2 **is more acidic**
Explanation: For the same element, the oxide with the higher oxidation number is always more acidic (due to higher charge density), just as Mn_2O_7 is strongly acidic but MnO is basic.

50. **B.** H_2SO_4 **is more acidic (S is smaller)**
Explanation: At equal oxidation number, the smaller central atom forms the stronger acid — this principle applies universally across every group.