MCQ 50 marks

1. **D. decreases**
Explanation: On raising the temperature, the vibration of cations in the metallic lattice increases and hinders the flow of free electrons, so the conductivity of the metal decreases.

2. **D.** $50\,\Omega^{-1}\text{cm}^2(\text{g-eqv})^{-1}$
Explanation: Number of equivalents of $Al_2(SO_4)_3$ per mole $= 2 \times 3 = 6$
$N = 0.01 \times 6 = 0.06\,N$
$\Lambda = \frac{1000\varkappa}{N} = \frac{1000 \times 3.0 \times 10^{-3}}{0.06} = 50\,\Omega^{-1}\text{cm}^2(\text{g-eqv})^{-1}$

3. **B.** Scm^{-1}
Explanation: $\varkappa = \frac{1}{A} \times \frac{1}{R}$
$\text{Unit} = \frac{\text{cm}}{\text{cm}^2} \times \Omega^{-1} = \Omega^{-1}\text{cm}^{-1} = \text{Scm}^{-1}$

4. **B. increases**
Explanation: Cathode: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$
Anode: $2Cl^- \rightarrow Cl_2 + 2e^-$
$NaOH$ accumulates in the solution, so the pH increases.

5. **A.** $2Cl^- \rightarrow Cl_2 + 2e^-$
Explanation: Oxidation, i.e. loss of electrons, occurs at the anode. In $2Cl^- \rightarrow Cl_2 + 2e^-$ electrons are lost, so it takes place at the anode.

6. **B.** $Ag^+ > Cu^{2+} > H^+ > Zn^{2+}$
Explanation: The lower a cation lies in the electrochemical series, the greater its tendency to be discharged at the cathode.
Order in the series (top to bottom):
Zn^{2+}, H^+, Cu^{2+}, Ag^+
Therefore, order of tendency: $Ag^+ > Cu^{2+} > H^+ > Zn^{2+}$

7. **C.** Cl_2
Explanation: The more positive the standard reduction potential, the stronger the oxidising agent. +1.36 V is the highest, so Cl_2 is the strongest oxidising agent.

8. **D. ions**
Explanation: Inside the cell, electricity is conducted by the movement of ions in the solution and the salt bridge, while electrons flow in the external wire.

9. **C. galvanic cell**
Explanation: In a galvanic cell a spontaneous redox reaction occurs and chemical energy is converted into electrical energy.

10. **B.** Cu
Explanation: A sacrificial anode must be more active than iron so that it gets oxidised itself and protects the iron. Cu is less active than iron, so it cannot be used.

11. **D.** 0.12 mol
Explanation: $n_{Cu} = \frac{3.81}{63.5} = 0.06\,\text{mol}$
$Cu^{2+} + 2e^- \rightarrow Cu$
Therefore, electrons $= 0.06 \times 2 = 0.12\,\text{mol}$

12. **C. When the concentration of** Cl^- **is much greater than that of** OH^-
Explanation: In a concentrated solution with $[Cl^-] \gg [OH^-]$, the effect of concentration makes the oxidation of Cl^- more favourable than that of OH^-, so Cl^- is oxidised first:
$2Cl^- \rightarrow Cl_2 + 2e^-$

13. **C.** 2.53×10^{22}
Explanation: $n = \frac{It}{eF} = \frac{3 \times 45 \times 60}{2 \times 96500} = 0.04197\,\text{mol}$
$N = 0.04197 \times 6.022 \times 10^{23} = 2.53 \times 10^{22}$

14. **A.** 4 g O^{2-}
Explanation: Moles of electrons $= \frac{W}{M} \times e$
O^{2-} : $\frac{4}{16} \times 2 = 0.50\,\text{mol}$
For comparison: F^- : 0.21, S^{2-} : 0.25, Mg^{2+} : 0.33, all lower. Therefore the greatest number is for O^{2-}.

15. **D.** 193000 C
Explanation: $Q = neF$
$Q = 0.5 \times 4 \times 96500 = 193000\,\text{C}$

16. **A.** 6 F
Explanation: $2Cl^- \rightarrow Cl_2 + 2e^-$
Producing one mole of Cl_2 requires 2 F.
$Q = 3 \times 2 = 6\,\text{F}$

17. **C.** 40.2 min
Explanation: $t = \frac{WeF}{MI} = \frac{2.16\times1\times96500}{108\times0.8} = 2412.5$ s
$t = \frac{2412.5}{60} = 40.2$ min

18. **A.** 12.43 g
Explanation: $2Cl^- \rightarrow Cl_2 + 2e^-$
1 F of charge produces 35.5 g of chlorine.
$W = 0.35 \times 35.5 = 12.43$ g

19. **A.** Li_xC_6
Explanation: In a lithium-ion battery, lithium intercalated in graphite (Li_xC_6) is used as the anode.

20. **A.** 0.5 F
Explanation: Each mole of ClO_4^- requires 2 mol of electrons, i.e. 2 F.
$Q = 0.25 \times 2 = 0.5$ F

21. **D.** 10.0 kg
Explanation: $W \propto It$
$\frac{W}{It} = \frac{1.0}{60\times1} = \frac{4.0}{120\times2} = \frac{8.0}{240\times2} = \frac{1}{60}$
$W = \frac{1}{60} \times 150 \times 4 = 10.0$ kg

22. **D. Cathode: iron key; anode: copper; electrolyte:** $CuSO_4$ **solution**
Explanation: The object to be plated (iron key) is the cathode and the plating metal (copper) is the anode; the electrolyte must contain ions of the plating metal (Cu^{2+}), so $CuSO_4$ solution is used.
At the cathode: $Cu^{2+} + 2e^- \rightarrow Cu$

23. **D.** 3.95 g
Explanation: $t = 3600 + 20 \times 60 = 4800$ s
$W = \frac{MIt}{eF} = \frac{63.5\times2.5\times4800}{2\times96500} = 3.95$ g

24. **B.** 3.02 g
Explanation: $Al^{3+} + 3e^- \rightarrow Al$
$W = \frac{MQ}{eF} = \frac{27\times32400}{3\times96500} = 3.02$ g

25. **D.** 4.08 g
Explanation: $Au^{3+} + 3e^- \rightarrow Au$
$W = \frac{MIt}{eF} = \frac{196.97\times2\times50\times60}{3\times96500} = 4.08$ g

26. **C.** 86.85 s
Explanation: At the anode: $Ag \rightarrow Ag^+ + e^-$
Condition for the start of precipitation:
$[Ag^+][Cl^-] = K_{sp}$
$[Ag^+] = \frac{1.8\times10^{-10}}{0.05} = 3.6 \times 10^{-9}$ M
$n_{Ag^+} = 3.6 \times 10^{-9} \times 0.5 = 1.8 \times 10^{-9}$ mol
$Q = nF = 1.8 \times 10^{-9} \times 96500 = 1.737 \times 10^{-4}$ C
$t = \frac{Q}{I} = \frac{1.737\times10^{-4}}{2\times10^{-6}} = 86.85$ s

27. **B.** 1.8 g-eqv
Explanation: By Faraday's law, 1 F of charge deposits 1 gram-equivalent of substance; the concentration is irrelevant here.
Therefore, 1.8 F → 1.8 g-eqv

28. **A. chemical equivalent**
Explanation: Faraday's second law: $W \propto E \Rightarrow \frac{W_1}{E_1} = \frac{W_2}{E_2}$
Here E = chemical equivalent = $\frac{M}{e}$

29. **D. in all of these**
Explanation: Extraction, purification and electroplating of metals are all electrolytic processes, and in each of them the amount of substance deposited or liberated obeys Faraday's law. Therefore it is applicable to all of these.

30. **C.** 25°C **temperature,** 1 M **concentration and** 1 atm **pressure**
Explanation: Standard conditions: temperature 25°C (298 K), ion concentration 1 M and gas pressure 1 atm.

31. **D. No current flows in the cell**
Explanation: When $E_{ext} = E_{cell}$ the cell is at equilibrium, so no current flows in the cell.

32. **A.** 0.312 V
Explanation: $E^0_{cell} = E^0_{Fe/Fe^{2+}} - E^0_{Sn/Sn^{2+}} = 0.44 - 0.14 = 0.30$ V
$E = E^0 - \frac{0.0591}{2} \log \frac{[Fe^{2+}]}{[Sn^{2+}]} = 0.30 - 0.02955 \log \frac{0.2}{0.5}$
$E = 0.30 + 0.0118 = 0.312$ V

33. **A.** $-2.2F - 4.606RT$
Explanation: $\Delta G = \Delta G^0 + RT \ln Q$
$Q = \frac{[Zn^{2+}]}{[Cu^{2+}]} = \frac{1}{100}$
$\Delta G^0 = -nF E^0 = -2 \times F \times 1.1 = -2.2F$
$RT \ln Q = RT \ln 10^{-2} = -2 \times 2.303RT = -4.606RT$
Therefore, $\Delta G = -2.2F - 4.606RT$

34. **A.** −0.148

Explanation: $2H^+ + 2e^- \rightarrow H_2$
$E_{red} = -\frac{0.0591}{2} \log \frac{1}{[H^+]^2} = -0.0591 \times pH$
$E_{red} = -0.0591 \times 2.5 = -0.148\ V$

35. **C. a straight line with positive slope (slope = 0.0295)**
Explanation: $E = E^0 + \frac{0.059}{2} \log[Zn^{2+}]$
This is a straight line of the form $y = c + mx$ with slope $m = +0.0295$ (positive).

36. **D.** $Fe^{2+} \longrightarrow Fe^{3+} + e^-$
Explanation: Oxidation (loss of electrons) occurs at the anode. In $Fe^{2+} \rightarrow Fe^{3+} + e^-$ an electron is lost, so it takes place at the anode.

37. **C. remain unchanged**
Explanation: E^0_{cell} is a constant for standard conditions and does not depend on concentration; only E_{cell} changes when the concentration changes. So E^0_{cell} remains unchanged.

38. **B.** Cu^{2+} **will be reduced by** Fe
Explanation: Possible reaction: $Fe + Cu^{2+} \rightarrow Fe^{2+} + Cu$
$E^0_{cell} = 0.34 - (-0.44) = +0.78\ V > 0$
The reaction is spontaneous; Fe acts as the reducing agent and Cu^{2+} is reduced.

39. **B.** $E^0 = \frac{RT}{nF} \ln K$
Explanation: $\Delta G^0 = -nFE^0$ and $\Delta G^0 = -RT \ln K$
$nFE^0 = RT \ln K$
Therefore, $E^0 = \frac{RT}{nF} \ln K$

40. **B.** $E_{cell} = E^0_{cell} - \frac{RT}{nF} \ln \frac{[Product]}{[Reactant]}$
Explanation: The Nernst equation:
$E_{cell} = E^0_{cell} - \frac{RT}{nF} \ln Q, \quad Q = \frac{[Product]}{[Reactant]}$

41. **C.** 0.0296 V
Explanation: This is a pressure cell, so $E^0_{cell} = 0$
Anode: $H_2(P_1) \rightarrow 2H^+ + 2e^-$
Cathode: $2H^+ + 2e^- \rightarrow H_2(P_2)$
$E = \frac{0.0591}{2} \log \frac{P_1}{P_2} = 0.02955 \log \frac{2}{0.2} = 0.0296\ V$

42. **B.** 0.0307 V
Explanation: $[Ag(NH_3)_2]^+ \rightarrow Ag^+ + 2NH_3, \quad E^0_1 = 0.0591 \log K_d = 0.0591 \times (-13) = -0.7683\ V$
$Ag^+ + e^- \rightarrow Ag, \quad E^0_2 = 0.799\ V$
Adding the two reactions (n = 1 in both):
$E^0 = E^0_1 + E^0_2 = -0.7683 + 0.799 = 0.0307\ V$

43. **C.** 2.88
Explanation: Reduction potential of the calomel electrode = +0.28 V; oxidation occurs at the H-electrode.
$emf = E_{ox(H_2)} + E_{red(cal)}$
$0.45 = E_{ox(H_2)} + 0.28 \Rightarrow E_{ox(H_2)} = 0.17\ V$
$E_{ox(H_2)} = 0.0591 \times pH$
$pH = \frac{0.17}{0.0591} = 2.88$

44. **A.** 1.482 V
Explanation: Here $n = 4$; $[H^+] = 10^{-4}$ M
$Q = \frac{[Fe^{2+}]^2}{P_{O_2}[H^+]^4} = \frac{(10^{-2})^2}{0.2 \times (10^{-4})^4} = 5 \times 10^{12}$
$E = E^0 - \frac{0.0591}{4} \log Q = 1.67 - \frac{0.0591}{4} \log(5 \times 10^{12})$
$E = 1.67 - 0.014775 \times 12.699 = 1.482\ V$

45. **C.** $E_{cell} = E^0_{cell} - \frac{2.303RT}{nF} \log \frac{[B^{n+}]}{[A^{n+}]}$
Explanation: $Q = \frac{[B^{n+}]}{[A^{n+}]}$ (activity of a solid = 1)
Therefore, $E_{cell} = E^0_{cell} - \frac{2.303RT}{nF} \log \frac{[B^{n+}]}{[A^{n+}]}$

46. **A.** $\ln\left(\frac{C_2}{C_1^2}\right)$
Explanation: $\Delta G = \Delta G^0 + RT \ln Q$
$Q = \frac{[Cu^{2+}]}{[Ag^+]^2} = \frac{C_2}{C_1^2}$
Therefore, ΔG depends on $\ln\left(\frac{C_2}{C_1^2}\right)$.

47. **D.** $E = E^0 - \frac{0.0611}{n} \log Q$
Explanation: $\frac{2.303RT}{F} = \frac{2.303 \times 8.314 \times 308}{96500} = 0.0611$
Therefore, $E = E^0 - \frac{0.0611}{n} \log Q$

48. **B. Calomel electrode**
Explanation: The calomel electrode has a well-defined potential and is easy to use, so it is used as a secondary reference electrode.

49. **A.** $2H_2(g) + 4OH^-(aq) \rightarrow 4H_2O + 4e^-$
Explanation: The electrolyte in an AFC is KOH solution; at the anode H_2 is oxidised in the alkaline medium:
$2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$

50. **B. Lithium battery**
Explanation: In a lithium battery, the anode is Li and the cathode is $Ag_2V_4O_{11}$:
$Ag_2V_4O_{11} + 7Li \rightarrow Li_7Ag_2V_4O_{11}$