MCQ 50 marks

1. **B.** $171.92\,\Omega^{-1}\text{cm}^2(\text{g-eqv})^{-1}$

Explanation: $\varkappa = \frac{1}{A} \times \frac{1}{R}$

Here, $l = 1.5\,\text{cm},\ A = 2.5\,\text{cm}^2,\ R = 6.5\,\Omega$

$\varkappa = \frac{1.5}{2.5\times6.5} = 0.09231\,\Omega^{-1}\text{cm}^{-1}$

Normality, $N = \frac{40}{74.5} = 0.5369$

$\Lambda = \frac{1000\varkappa}{N} = \frac{1000\times0.09231}{0.5369}$

Therefore, $\Lambda = 171.92\,\Omega^{-1}\text{cm}^2(\text{g-eqv})^{-1}$

2. **C.** $149.8\,\Omega^{-1}\text{cm}^2\text{mol}^{-1}$

Explanation: By Kohlrausch's law, $\Lambda_m^\infty = \lambda_+^\infty + \lambda_-^\infty$

Therefore, $\Lambda_m^\infty = 73.5 + 76.3 = 149.8\,\Omega^{-1}\text{cm}^2\text{mol}^{-1}$

3. **C.** CH_3COOH

Explanation: CH_3COOH is a weak electrolyte; its degree of dissociation rises rapidly on dilution, so its equivalent conductivity increases at the greatest rate.

4. **C.** $KNO_3 > NaNO_3 > LiNO_3$

Explanation: In water, the smaller the ion, the greater its hydration and the lower its mobility. Order of hydrated size: $Li^+ > Na^+ > K^+$

Therefore, the order of conductivity is: $KNO_3 > NaNO_3 > LiNO_3$

5. **A.** H_2 **and** O_2

Explanation: At the cathode, H^+ is discharged in preference to K^+: $2H^+ + 2e^- \rightarrow H_2$

At the anode, OH^- is oxidised in preference to SO_4^{2-}: $4OH^- \rightarrow 2H_2O + O_2 + 4e^-$

Therefore, H_2 is formed at the cathode and O_2 at the anode.

6. **B. i, ii and iv**

Explanation: At the cathode: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$

At the anode (concentrated solution): $2Cl^- \rightarrow Cl_2 + 2e^-$

K^+ and OH^- remain in the solution, giving KOH. Therefore, (i), (ii) and (iv) are produced.

7. **C.** KCl

Explanation: The cation and anion of a salt-bridge electrolyte must have nearly equal mobilities and must not react with the half-cell solutions. The K^+ and Cl^- ions of KCl have nearly equal mobilities, so it can be used.

8. **B.** $Ag(s)$

Explanation: The reduction potential of Ag ($+0.80\,\text{V}$) is higher than that of Cu ($+0.34\,\text{V}$), i.e. Ag lies below Cu in the activity series. So Ag cannot reduce Cu^{2+}, and no solid copper is obtained.

9. **D.** $54\,\text{g}$

Explanation: $Al^{3+} + 3e^- \rightarrow Al$

Here, $3\,\text{F}$ deposits $27\,\text{g}$ of Al

Therefore, $6\,\text{F}$ will deposit $\frac{27\times6}{3} = 54\,\text{g}$

10. **B.** $78\,\text{g}$

Explanation: $Cr^{3+} + 3e^- \rightarrow Cr$

$W = \frac{M\times Q}{eF}$

Here, $e = 3$, $Q = 4.5\,\text{F}$

$W = \frac{52\times4.5\text{F}}{3\text{F}} = 78\,\text{g}$

11. **C.** $1.617\,\text{g}$

Explanation: $W = \frac{MIt}{eF}$

Here, $M = 52,\ I = 6\,\text{A},\ t = 25\times60\,\text{s},\ e = 3$

$W = \frac{52\times6\times25\times60}{3\times96500} = 1.617\,\text{g}$

12. **B.** $0.806\,\text{g}$

Explanation: $W = \frac{MIt}{eF}$

Here, $M = 108,\ I = 0.3\,\text{A},\ t = 40\times60\,\text{s},\ e = 1$

$W = \frac{108\times0.3\times40\times60}{1\times96500} = 0.806\,\text{g}$

13. **C. Tk. 2204**

Explanation: $W = \frac{MIt}{eF} = \frac{196.97\times6\times12\times60}{3\times96500} = 2.939\,\text{g}$

Total cost of gold $= \frac{2.939\times60000}{10} \approx 17635$

Therefore, per chain $= \frac{17635}{8} \approx 2204$

14. **A.** 2.340×10^{21}

Explanation: $n = \frac{It}{eF}$

$n = \frac{0.25\times3000}{2\times96500} = 3.886\times10^{-3}\,\text{mol}$

Therefore, number of atoms $= 3.886\times10^{-3}\times6.022\times10^{23} = 2.340\times10^{21}$

15. **B.** 7.107 g
Explanation: $W = \frac{MIt}{eF}$
$W = \frac{63.5\times4\times90\times60}{2\times96500} = 7.107\,g$

16. **B.** 0.7788 V
Explanation: $Zn \rightarrow Zn^{2+} + 2e^-$
$E = E^0 - \frac{RT}{nF}\ln[Zn^{2+}]$
$E = 0.758 - \frac{8.314\times300}{2\times96500}\ln 0.2$
$E = 0.758 + 0.01292 \times 1.609 = 0.7788\,V$

17. **B.** 1.204×10^{24}
Explanation: $Al^{3+} + 3e^- \rightarrow Al$
$n_{Al} = \frac{18}{27} = 0.6667\,mol$
Therefore, number of electrons = $0.6667 \times 3 \times 6.022 \times 10^{23} = 1.204 \times 10^{24}$

18. **D.** 144750 C
Explanation: $Ca^{2+} + 2e^- \rightarrow Ca$
$Q = \frac{WeF}{M} = \frac{30\times2\times96500}{40} = 144750\,C$

19. **B.** $3.30 \times 10^{-7}\,kgC^{-1}$
Explanation: $W = ZIt \Rightarrow Z = \frac{W}{It}$
Here, $W = 0.198\,g = 0.198 \times 10^{-3}\,kg,\ I = 0.25\,A,\ t = 40 \times 60\,s$
$Z = \frac{0.198\times10^{-3}}{0.25\times40\times60} = 3.30 \times 10^{-7}\,kgC^{-1}$

20. **A.** 0.4029 g
Explanation: $W \propto It$
$\frac{W_2}{W_1} = \frac{I_2t_2}{I_1t_1}$
$W_2 = \frac{0.2417\times0.15\times40}{0.30\times12} = 0.4029\,g$

21. **A.** 2316 s
Explanation: $n_{Fe^{3+}} = \frac{120}{1000} \times 0.15 \times 2 = 0.036\,mol$
$Fe^{3+} + e^- \rightarrow Fe^{2+}$
$t = \frac{neF}{I} = \frac{0.036\times1\times96500}{1.5} = 2316\,s$

22. **B.** 64.33 min
Explanation: $t = \frac{WeF}{MI}$
$t = \frac{3.27\times2\times96500}{65.4\times2.5} = 3860\,s$
$t = \frac{3860}{60} = 64.33\,min$

23. **D.** 35.90 min
Explanation: $n_{H_2} = \frac{500}{22400} = 0.02232\,mol$
$2H^+ + 2e^- \rightarrow H_2$
$t = \frac{neF}{I} = \frac{0.02232\times2\times96500}{2} = 2154\,s$
$t = \frac{2154}{60} = 35.90\,min$

24. **D.** 482.5 s
Explanation: $n_{H_2} = \frac{67.2}{22.4} = 3\,mol$
$H_2 \rightarrow 2H^+ + 2e^-$
$t = \frac{neF}{I} = \frac{3\times2\times96500}{1200} = 482.5\,s$

25. **A.** 0.635 g
Explanation: Faraday's second law: $\frac{W_{Ag}}{E_{Ag}} = \frac{W_{Cu}}{E_{Cu}}$
Here, $E_{Ag} = 108,\ E_{Cu} = \frac{63.5}{2} = 31.75$
$W_{Cu} = 2.16 \times \frac{31.75}{108} = 0.635\,g$

26. **A.** 1.31
Explanation: Here, $t = 6 \times 60 + 30 = 390\,s$
$Q = 3 \times 390 \times 0.80 = 936\,C$
$n_{e^-} = \frac{936}{96500} = 9.70 \times 10^{-3}\,mol$
At the anode: $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$
Therefore, moles of H^+ formed = 9.70×10^{-3}
$[H^+] = \frac{9.70\times10^{-3}}{0.2} = 0.0485\,M$
$pH = -\log(0.0485) = 1.31$

27. **C.** 1.05 V
Explanation: $E^0_{cell} = E^0_{cathode} - E^0_{anode}$
$E^0_{cell} = 0.80 - (-0.25) = 1.05\,V$

28. **C.** 1.218 V
Explanation: $E = E^0 - \frac{0.0591}{n}\log\frac{[Zn^{2+}]}{[Cu^{2+}]}$
Here, $n = 2$
$E = 1.1 - \frac{0.0591}{2}\log\frac{10^{-6}}{10^{-2}}$
$E = 1.1 + 0.02955 \times 4 = 1.218\,V$

29. **A.** −0.251 V
Explanation: $E = E^0_{cell} - \frac{0.0591}{2}\log\frac{[Ni^{2+}]}{[Cu^{2+}]}$
$0.65 = E^0_{cell} - 0.02955\log\frac{0.001}{0.1} = E^0_{cell} + 0.0591$
$E^0_{cell} = 0.5909\,V$
$E^0_{cell} = E^0_{Cu^{2+}/Cu} - E^0_{Ni^{2+}/Ni}$
$E^0_{Ni^{2+}/Ni} = 0.34 - 0.5909 = -0.251\,V$

30. **D.** $E_1 > E_4 > E_3 > E_2$
Explanation: $E = E^0 - \frac{0.0591}{2}\log Q,\quad Q = \frac{[Zn^{2+}]}{[Ag^+]^2}$
The smaller the value of Q, the larger the value of E.
$Q_1 = \frac{0.01}{1^2} = 0.01,\quad Q_2 = \frac{1}{(0.01)^2} = 10^4$
$Q_3 = 1,\quad Q_4 = \frac{0.001}{(0.1)^2} = 0.1$
$Q_1 < Q_4 < Q_3 < Q_2$
Therefore, $E_1 > E_4 > E_3 > E_2$

31. **A.** $E_1 > E_3 > E_2$

Explanation: $E = E^0 - \frac{0.0591}{2}\log Q, \quad Q = \frac{[Zn^{2+}]}{[Cu^{2+}]}$
$Q_1 = \frac{0.01}{1} = 0.01, \quad Q_2 = \frac{1}{0.1} = 10, \quad Q_3 = 1$
$Q_1 < Q_3 < Q_2$
Therefore, $E_1 > E_3 > E_2$

32. **D.** 1.182 V
Explanation: $E^0 = \frac{0.0591}{n}\log K$
Here, $n = 2, \ K = 10^{40}$
$E^0 = \frac{0.0591}{2} \times 40 = 1.182\ V$

33. **D.** 1.581 V
Explanation: $E = E^0 - \frac{0.0591}{2}\log\frac{[Zn^{2+}]}{[Ag^+]^2}$
$E = 1.56 - 0.02955\log\frac{0.05}{(0.5)^2}$
$E = 1.56 + 0.02955 \times 0.699 = 1.581\ V$

34. **C.** 0.1076 V
Explanation: $Pb + 2H^+ \rightarrow Pb^{2+} + H_2$
$E^0_{cell} = 0 - (-0.14) = 0.14\ V$
$E = E^0 - \frac{0.0591}{2}\log\frac{[Pb^{2+}]P_{H_2}}{[H^+]^2}$
$E = 0.14 - 0.02955\log\frac{0.5\times1}{(0.2)^2}$
$E = 0.14 - 0.02955 \times 1.097 = 0.1076\ V$

35. **D.** $E^0_{cell} = +0.59\ V$**; the solution cannot be kept in a nickel vessel**
Explanation: Possible reaction: $Ni + Cu^{2+} \rightarrow Ni^{2+} + Cu$
$E^0_{cell} = 0.34 - (-0.25) = +0.59\ V$
Since $E^0_{cell} > 0$, the reaction is spontaneous and the Ni vessel will corrode. Therefore, the solution cannot be kept in a nickel vessel.

36. **A.** 0.8096 V
Explanation: $E^0_{cell} = E^0_{Fe/Fe^{2+}} - E^0_{Cu/Cu^{2+}} = 0.44 - (-0.34) = 0.78\ V$
$E = E^0 - \frac{0.0591}{2}\log\frac{[Fe^{2+}]}{[Cu^{2+}]}$
$E = 0.78 - 0.02955\log\frac{0.05}{0.5} = 0.78 + 0.02955 = 0.8096\ V$

37. **C.** 0.1919 V
Explanation: $Sn \rightarrow Sn^{2+} + 2e^-$
$E = E^0 - \frac{RT}{nF}\ln[Sn^{2+}]$
$E = 0.14 - \frac{8.314\times308}{2\times96500}\ln 0.02$
$E = 0.14 + 0.01327 \times 3.912 = 0.1919\ V$

38. **D.** 0.1360 V
Explanation: $H_2 \rightarrow 2H^+ + 2e^-$
$E_{ox} = -\frac{0.0591}{2}\log[H^+]^2 = -0.0591\log[H^+]$
$E_{ox} = -0.0591\log(0.005) = 0.0591 \times 2.301 = 0.1360\ V$

39. **C.** 1.527 V
Explanation: $2Al + 3Sn^{2+} \rightarrow 2Al^{3+} + 3Sn$
Here, $n = 6$
$E^0 = -0.14 - (-1.66) = 1.52\ V$
$Q = \frac{[Al^{3+}]^2}{[Sn^{2+}]^3} = \frac{(0.2)^2}{(0.6)^3} = 0.1852$
$E = E^0 - \frac{0.0591}{6}\log Q = 1.52 - \frac{0.0591}{6}\log 0.1852$
$E = 1.52 + 0.0072 = 1.527\ V$

40. **C.** 0.1182 V
Explanation: This is a concentration cell, so $E^0_{cell} = 0$
$E = \frac{0.0591}{n}\log\frac{C_{cathode}}{C_{anode}}$
$E = 0.0591\log\frac{0.1}{0.001} = 0.0591 \times 2 = 0.1182\ V$

41. **B.** $-231.6\ kJmol^{-1}$
Explanation: $3Mn + 2Cr^{3+} \rightarrow 3Mn^{2+} + 2Cr$
Here, $n = 6$
$\Delta G = -nFE$
$\Delta G = -6 \times 96500 \times 0.40 = -231600\ Jmol^{-1} = -231.6\ kJmol^{-1}$

42. **A.** $E^0_3 = 2E^0_1 - E^0_2$
Explanation: Potentials cannot be added directly; the free energies $\Delta G^0 = -nFE^0$ are added.
$Cu^{2+} + 2e^- \rightarrow Cu, \quad \Delta G^0_1 = -2FE^0_1$
$Cu^+ + e^- \rightarrow Cu, \quad \Delta G^0_2 = -FE^0_2$
$Cu^{2+} + e^- \rightarrow Cu^+, \quad \Delta G^0_3 = \Delta G^0_1 - \Delta G^0_2$
$-FE^0_3 = -2FE^0_1 + FE^0_2$
Therefore, $E^0_3 = 2E^0_1 - E^0_2$

43. **B.** −202.65 kJ
Explanation: $E^0_{cell} = 0.80 - (-0.25) = 1.05\ V$
Here, $n = 2$
$\Delta G^0 = -nFE^0 = -2 \times 96500 \times 1.05 = -202650\ J = -202.65\ kJ$

44. **D.** 9.25×10^{19}
Explanation: $\log K = \frac{nE^0}{0.0591}$
$\log K = \frac{2\times0.59}{0.0591} = 19.966$
$K = 10^{19.966} = 9.25 \times 10^{19}$

45. **C.** 4.62×10^{-13}
Explanation: Required reaction: $AgBr \rightarrow Ag^+ + Br^-$
It is the sum of $(AgBr + e^- \rightarrow Ag + Br^-)$ and

$(Ag \rightarrow Ag^+ + e^-)$.
$E^0 = +0.071 + (-0.800) = -0.729\,V$
$\log K_{sp} = \frac{nE^0}{0.0591} = \frac{1\times(-0.729)}{0.0591} = -12.335$
$K_{sp} = 4.62 \times 10^{-13}$

46. **A. positive,** < 1**, negative**
Explanation: $\Delta G^0 = -nF\,E^0 = -RT\,\ln K$
For a non-spontaneous reaction, $\Delta G^0 > 0$
Therefore, $E^0 < 0$ and $\ln K < 0$, i.e. $K < 1$.

47. **A.** $P_{H_2} = 2\,atm$ **and** $[H^+] = 1\,M$
Explanation: $E_{red} = -\frac{0.0591}{2} \log \frac{P_{H_2}}{[H^+]^2}$
For $E_{red} < 0$, we need $\frac{P_{H_2}}{[H^+]^2} > 1$.
For $P_{H_2} = 2\,atm$ and $[H^+] = 1\,M$, $\frac{P_{H_2}}{[H^+]^2} = 2 > 1$
, so E_{red} is negative.

48. **D. Decreasing** $[Ni^{2+}]$
Explanation: $E = E^0 - \frac{0.0591}{2} \log \frac{[Ni^{2+}]}{[Cu^{2+}]}$
Decreasing $[Ni^{2+}]$ lowers $\frac{[Ni^{2+}]}{[Cu^{2+}]}$, so E increases.

49. **D.** $1.092\,V$
Explanation: Each O_2 molecule takes up 4 electrons, so $n = 5 \times 4 = 20$
$\Delta G^0 = -nF\,E^0$
$E^0 = \frac{2108\times1000}{20\times96500} = 1.092\,V$

50. **B.** PbO_2
Explanation: During charging, oxidation occurs at the positive plate: $PbSO_4 + 2H_2O \rightarrow PbO_2 + 4H^+ + SO_4^{2-} + 2e^-$
Therefore, PbO_2 is formed.