MCQ 50 marks

1. **D.** $8.33 \times 10^{-3}\ \Omega^{-1}\text{cm}^{-1}$

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

$\varkappa = \frac{0.9}{1.8 \times 60} = 8.33 \times 10^{-3}\ \Omega^{-1}\text{cm}^{-1}$

2. **A.** $HNO_3 > H_2SO_4 > H_3PO_4$

Explanation: $\Lambda = \frac{1000\varkappa}{N}$

When $\varkappa$ is the same, $\Lambda \propto \frac{1}{N}$

$N_{HNO_3} = 0.02$, $N_{H_2SO_4} = 0.04$, $N_{H_3PO_4} = 0.06$

Therefore, $HNO_3 > H_2SO_4 > H_3PO_4$

3. **B. Mercury**

Explanation: Order of metallic conductivity: $Ag > Cu > Au \gg Hg$

Therefore, mercury has the lowest electrical conductivity.

4. **C. In the solid state**

Explanation: In solid NaCl the ions are held fixed in the crystal lattice and cannot move freely, so it does not conduct electricity in the solid state.

5. **A.** Ag **at the anode is oxidised to** Ag^+

Explanation: The active Ag anode itself is oxidised: $Ag \rightarrow Ag^+ + e^-$

6. **C. Sodium amalgam (Na-Hg)**

Explanation: H_2 has a high overpotential on a mercury cathode, so Na^+ is discharged: $Na^+ + e^- + Hg \rightarrow Na\text{-}Hg$

7. **D.** 2 : 1

Explanation: Cathode: $4H^+ + 4e^- \rightarrow 2H_2$

Anode: $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

For the same charge, $H_2 : O_2 = 2 : 1$

The mole ratio of gases equals the volume ratio, so the ratio is 2 : 1.

8. **C.** $Z > X > Y$

Explanation: The more negative the reduction potential, the stronger the reducing agent.

$-2.90 < -0.40 < +0.80$

Therefore, $Z > X > Y$

9. **A.** Pt

Explanation: Pt does not take part in the electrolysis, so it is used as an inert electrode.

10. **B.** I^-

Explanation: Among the given ions, I^- is oxidised most easily, because $E^0_{I_2/I^-} = 0.54\ V$ is the lowest of the corresponding reduction potentials.

At the anode: $2I^- \rightarrow I_2 + 2e^-$

11. **D. i and iii**

Explanation: Activity series: Al > Zn > Fe > H > Cu

Zn lies above Fe and H but below Al.

Therefore (i) and (iii) are correct.

12. **B.** H_2, I_2 **and** KOH

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

At the anode: $2I^- \rightarrow I_2 + 2e^-$

K^+ and OH^- remain in the solution, giving KOH. Therefore, H_2, I_2 and KOH are produced.

13. **D.** 3.57 A

Explanation: $2I^- \rightarrow I_2 + 2e^-$

$n_{I_2} = \frac{12.7}{254} = 0.05\ \text{mol}$

$Q = neF = 0.05 \times 2 \times 96500 = 9650\ C$

$I = \frac{Q}{t} = \frac{9650}{45 \times 60} = 3.57\ A$

14. **D.** 6.47 g

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

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

$W = \frac{52 \times 2.5 \times 4 \times 3600}{3 \times 96500} = 6.47\ g$

15. **D.** 1.73 g

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

$W = \frac{63.5 \times 3.5 \times 25 \times 60}{2 \times 96500} = 1.73\ g$

16. **A.** 27

Explanation: $W = \frac{MIt}{eF} \Rightarrow M = \frac{WeF}{It}$

$M = \frac{0.2015 \times 3 \times 96500}{1.2 \times 30 \times 60} = 27$

17. **C.** 5.6 L

Explanation: $2H^+ + 2e^- \rightarrow H_2$

Here, 2 F produces 22.4 L of H_2

Therefore, 0.5 F produces $\frac{22.4 \times 0.5}{2} = 5.6\ L$

18. **C.** 167.9 kg
Explanation: $Al^{3+} + 3e^- \rightarrow Al$
$W = \frac{MIt}{eF} = \frac{27\times5\times10^4\times10\times3600}{3\times96500} = 167876\,g$
$W = 167.9\,kg$

19. **B.** 1.40×10^{19}
Explanation: $Ca^{2+} + 2e^- \rightarrow Ca$
$n = \frac{It}{eF} = \frac{50\times10^{-3}\times90}{2\times96500} = 2.332 \times 10^{-5}$ mol
Therefore, number of atoms = $2.332 \times 10^{-5} \times 6.022 \times 10^{23} = 1.40 \times 10^{19}$

20. **B.** $3.04 \times 10^{-4}\ gC^{-1}$
Explanation: $Z = \frac{M}{eF} = \frac{58.7}{2\times96500} = 3.04 \times 10^{-4}\ gC^{-1}$

21. **A.** 1.42 g
Explanation: $W = \frac{MIt}{eF}$
$W = \frac{65.37\times1.75\times40\times60}{2\times96500} = 1.42\,g$

22. **D.** 0.504 F
Explanation: $Q = It = 4.5 \times 3 \times 3600 = 48600\,C$
$Q = \frac{48600}{96500}\,F = 0.504\,F$

23. **D.** 2.35 A
Explanation: $n = \frac{PV}{RT} = \frac{1.5\times0.6}{0.0821\times300} = 0.03654$ mol
$Q = neF = 0.03654 \times 2 \times 96500 = 7052\,C$
$I = \frac{Q}{t} = \frac{7052}{50\times60} = 2.35\,A$

24. **A.** K
Explanation: 1 F deposits one gram-equivalent ($\frac{M}{e}$) of metal.
$K = \frac{39.1}{1} = 39.1\,g,\ Ca = \frac{40}{2} = 20\,g,\ Al = \frac{27}{3} = 9\,g,\ Zn = \frac{65.4}{2} = 32.7\,g$
Therefore, K is deposited in the greatest amount.

25. **A.** 1.5m
Explanation: $W \propto It$
$\frac{W_2}{m} = \frac{2\times9}{3\times4} = \frac{3}{2}$
$W_2 = 1.5m$

26. **C. Anode** 10.49 g**, cathode** 13.51 g
Explanation: $W = \frac{MIt}{eF} = \frac{108\times1.5\times15\times60}{1\times96500} = 1.51\,g$
At the anode, $Ag \rightarrow Ag^+ + e^-$; mass decreases: $12.0 - 1.51 = 10.49\,g$
At the cathode, $Ag^+ + e^- \rightarrow Ag$; mass increases: $12.0 + 1.51 = 13.51\,g$

27. **C.** $\frac{1}{2}$ mol **of X will be deposited**
Explanation: $X^{2+} + 2e^- \rightarrow X$
Since e = 2, 1 F deposits $\frac{1}{2}$ mol of X.

28. **B.** 1.5×10^{18}
Explanation: $Q = It = 0.5 \times 10^{-3} \times 8 \times 60 = 0.24\,C$
$N = \frac{Q}{e} = \frac{0.24}{1.6\times10^{-19}} = 1.5 \times 10^{18}$

29. **B.** 6 F
Explanation: $Al^{3+} + 3e^- \rightarrow Al$
Here, 27 g of Al needs 3 F
Therefore, 54 g of Al needs $\frac{3\times54}{27} = 6\,F$

30. **A.** 193000 C
Explanation: Each Cu^{2+} ion carries 2 unit positive charges.
$Q = 2F = 2 \times 96500 = 193000\,C$

31. **C.** 2
Explanation: $Q = It = 0.0402 \times 2 \times 3600 = 289.4\,C$
$W = \frac{MQ}{eF} \Rightarrow e = \frac{MQ}{WF}$
$e = \frac{58.7\times289.4}{0.0880\times96500} = 2$

32. **B.** 13.99
Explanation: At the cathode: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$
$n_{OH^-} = \frac{It}{F} = \frac{1.5\times180}{96500} = 2.798 \times 10^{-3}$ mol
$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$
$n_{H_2SO_4} = \frac{2.798\times10^{-3}}{2} = 1.399 \times 10^{-3}$ mol
$V = \frac{1.399\times10^{-3}}{0.1}\,L = 13.99\,mL$

33. **B.** 0.672 g
Explanation: $Q = It = 0.8 \times 2.5 \times 3600 = 7200\,C$
Decomposing one mole of H_2O transfers 2 mol of electrons: $H_2O \rightarrow H_2 + \frac{1}{2}O_2$
$n_{H_2O} = \frac{7200}{2\times96500} = 0.0373$ mol
$W = 0.0373 \times 18 = 0.672\,g$

34. **A.** 1.5 M
Explanation: At the anode: $Cu \rightarrow Cu^{2+} + 2e^-$
At the cathode: $Cu^{2+} + 2e^- \rightarrow Cu$
The amount of Cu^{2+} entering the solution at the anode equals the amount deposited at the cathode, so the concentration remains unchanged at 1.5 M.

35. **D.** 1.08 V

Explanation: $E^0_{cell} = E^0_{cathode} - E^0_{anode}$
$E^0_{cell} = 0.80 - (-0.28) = 1.08\ V$

36. **D. It becomes zero**
Explanation: The salt bridge completes the internal circuit of the cell. On removing it, the circuit is broken and ionic movement stops, so the cell potential becomes zero.

37. **D.** Ni
Explanation: In a galvanic cell the anode is the negative terminal. Here Ni is oxidised ($Ni \rightarrow Ni^{2+} + 2e^-$), so Ni is the negative terminal.

38. **C.** $\frac{G^*}{R}$
Explanation: $\varkappa = \frac{1}{A} \times \frac{1}{R}$
Cell constant $G^* = \frac{1}{A}$
Therefore, $\varkappa = \frac{G^*}{R}$

39. **A.** $0.2613\ V$
Explanation: $2Cr + 3Fe^{2+} \rightarrow 2Cr^{3+} + 3Fe$
Here, $n = 6$
$E^0 = -0.44 - (-0.72) = 0.28\ V$
$Q = \frac{[Cr^{3+}]^2}{[Fe^{2+}]^3} = \frac{(0.1)^2}{(0.05)^3} = 80$
$E = E^0 - \frac{0.0591}{6} \log Q = 0.28 - \frac{0.0591}{6} \log 80$
$E = 0.28 - 0.00985 \times 1.903 = 0.2613\ V$

40. **B.** $0.371\ V$
Explanation: $Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag, \quad n = 2$
$E^0 = 0.80 - 0.34 = 0.46\ V$
$E = E^0 - \frac{0.0591}{2} \log \frac{[Cu^{2+}]}{[Ag^+]^2} = 0.46 - 0.02955 \log \frac{0.1}{(0.01)^2}$
$E = 0.46 - 0.02955 \times 3 = 0.371\ V$

41. **B.** 10^6
Explanation: $\log K = \frac{nE^0}{0.0591} = \frac{1 \times 0.3546}{0.0591} = 6$
$K = 10^6$

42. **C.** $Cu \mid Cu^{2+}(1.0\ M) \parallel Ag^+(1.0\ M) \mid Ag$
Explanation: For every cell $n = 2$ and $\Delta G^0 = -nFE^0$, so the lowest E^0 gives the highest (least negative) ΔG^0.
$E^0_{Cu|Ag} = 0.80 - 0.337 = 0.463\ V$
E^0 of the other cells: $1.097\ V$, $1.56\ V$, $3.16\ V$, all higher.

43. **C.** $-384.07\ kJ$
Explanation: $E^0_{cell} = 1.23 - (-0.76) = 1.99\ V$
Here, $n = 2$
$\Delta G^0 = -nFE^0 = -2 \times 96500 \times 1.99 = -384070\ J = -384.07\ kJ$

44. **B. (i), (ii) and (iii)**
Explanation: $E^0 = 0.80 - (-0.76) = 1.56\ V$
$E_{cell} = E^0 - \frac{0.0591}{2} \log \frac{[Zn^{2+}]}{[Ag^+]^2} = 1.56 - 0.02955 \log \frac{10^{-6}}{(0.01)^2}$
$E_{cell} = 1.56 + 0.0591 = 1.619\ V$
$\Delta G^0 = -nFE^0 = -2 \times 96500 \times 1.56 = -3.011 \times 10^5\ J$
Therefore, (i), (ii) and (iii) are all correct.

45. **A.** Zn^{2+} **will not be reduced**
Explanation: Possible reaction: $Cu + Zn^{2+} \rightarrow Cu^{2+} + Zn$
$E^0_{cell} = -0.76 - 0.34 = -1.10\ V < 0$
The reaction is not spontaneous, so Zn^{2+} will not be reduced.

46. **D.** $0.0517\ V$
Explanation: $\alpha = \frac{\Lambda}{\Lambda^\infty} = \frac{5.2}{390} = 0.01333$
$[H^+]_{HA} = 0.1 \times 0.01333 = 1.333 \times 10^{-3}\ M$
This is a concentration cell, so $E^0_{cell} = 0$
$E = 0.0591 \log \frac{[H^+]_{cathode}}{[H^+]_{anode}} = 0.0591 \log \frac{0.01}{1.333 \times 10^{-3}}$
$E = 0.0591 \times 0.875 = 0.0517\ V$

47. **C.** $2.5\ mol$
Explanation: Energy required to oxidise one mole of M^{2+}: $\Delta G^0 = -nFE^0 = -2 \times 96500 \times (-0.50) = +96500\ J$
Energy produced $= 241.25\ kJ = 241250\ J$
$n_{M^{2+}} = \frac{241250}{96500} = 2.5\ mol$

48. **A.** -18.27
Explanation: $2A + B^{2n+} \rightarrow 2A^{n+} + B, \quad Q = \frac{(3)^2}{1} = 9$
$E = 0 \Rightarrow \Delta G = 0 \Rightarrow \Delta G^0 = -RT \ln Q$
$\Delta H^0 = 2\Delta G^0$
$\Delta S^0 = \frac{\Delta H^0 - \Delta G^0}{T} = \frac{\Delta G^0}{T} = -R \ln 9$
$\Delta S^0 = -8.314 \times 2.197 = -18.27\ JK^{-1}mol^{-1}$

49. **B. decreases**
Explanation: Discharge reaction: $Pb + PbO_2 + 2H_2SO_4 \rightarrow 2PbSO_4 + 2H_2O$
H_2SO_4 is consumed and water is produced, so the specific gravity decreases.

50. **C.** CO_2

Explanation: Anode reaction: $2CH_3OH + 2H_2O \rightarrow 2CO_2 + 12H^+ + 12e^-$
Therefore, CO_2 gas is produced.