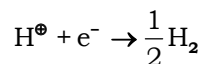


PART B - CHEMISTRY

31. b) decrease by 0.41V

$$\text{pH} = 0, [\text{H}^+] = 1\text{M}, E_{\text{electrode}}^{\circ} = 0$$

$$\text{pH} = 7, [\text{H}^+] = 10^{-7}\text{M}$$



$$E_{\text{el}} = E_{\text{el}}^{\circ} - \frac{0.0591}{1} \log \frac{1}{[\text{H}^+]}$$

$$= 0 - 0.0591 \times 7 = -0.41\text{V}$$

32. a) -0.62 V

$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$$

(both are reduction potential)

$$= 0.15 - 0.77$$

$$= -0.62\text{V}$$

33. c)
- H_2

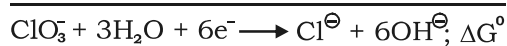
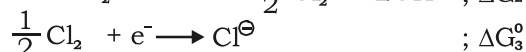
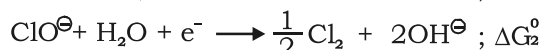
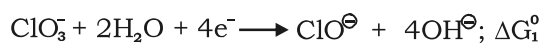
After deposition of Cu^{+2} ion from the H_2O molecule will be discharged at cathode as its SRP is higher than that of Fe^{+2} & Zn^{+2} .

34. b) nonspontaneous (
- $E_{\text{cell}} < 0$
-)

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{1} \log_{10} \frac{[\text{H}^+]_{\text{anode}}}{[\text{H}^+]_{\text{cathode}}}$$

$$= \frac{-0.0591}{1} \log_{10} \frac{10^{-2}}{10^{-3}} = (-) \text{ve}$$

35. b) 0.6 V



$$\Delta G^{\circ} = \Delta G_1^{\circ} + \Delta G_2^{\circ} + \Delta G_3^{\circ}$$

$$-6FE^{\circ} = -4F \times 0.54 - 1F \times 0.45 - 1F \times 1.07$$

$$E^{\circ} = \frac{3.68}{6} = 0.61\text{V}$$

36. d)
- $\text{Pt}, \text{H}_2 | 0.1\text{M CH}_3\text{COOK} || 0.1\text{M HCl} | \text{H}_2, \text{Pt}$

$$E_{\text{cell}} = \frac{-0.0591}{1} \log \frac{[\text{H}^+]_{\text{anode}}}{[\text{H}^+]_{\text{cathode}}}$$

($E_{\text{cell}}^{\circ} = 0$ for concentration cell) for E_{cell} to be highest $[\text{H}^+]_{\text{a}}$ should be lower and $[\text{H}^+]_{\text{c}}$ should be higher (anode compartment should be basic and cathodic compartment should be acidic)

37. c)
- $\text{Pt}(\text{H}_2) | \text{pH} = 1 || z_n^{+2} (1\text{M}) \text{Zn(s)}$

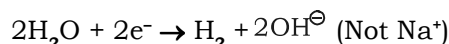
$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{[\text{Cu}^{+2}]}{[\text{Ag}^+]^2}$$

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{0.01}{(0.1)^2}$$

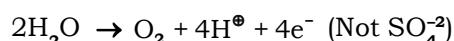
$$\Rightarrow E_{\text{cell}} = E_{\text{cell}}^{\circ}$$

Similarly for (b) and (d)

38. a)
- H_2, O_2
-
- at cathode



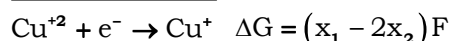
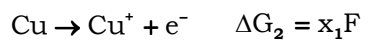
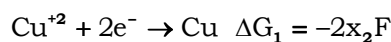
at anode



39. b) reaction quotient Q decrease,
- E_{cell}
- increase

With time, Q increases, E_{cell} decreases at equilibrium $Q = K_{\text{eq}}$ & $E_{\text{cell}} = 0$

40. d)
- $2x_2 - x_1$



$$E = \frac{\Delta G}{-nF} = \frac{(x_1 - 2x_2)}{-F}$$

$$E = 2x_2 - x_1$$

41. a) $k_{sp}(\text{AgCl}) > k_{sp}(\text{AgBr}) > k_{sp}(\text{AgI}) > k_{sp}(\text{Ag}_2\text{S})$

$$E_{\text{Ag}|\text{AgCl}|\text{Cl}^\ominus}^0 = E_{\text{Ag}^+|\text{Ag}}^0 + \frac{0.059}{1} \log k_{sp}$$

Other equations can be written similarly

42. d) **Pt | Fe⁺³, Fe⁺², metal - metalion electrode**

Pt | Fe⁺³, Fe⁺² is a redox electrode or inert electrode. Metal - metal ion electrode : Zn⁺² | Zn, Cu⁺² | Cu , etc.

43. c) **4950**

$$\frac{100 \times 99}{2} = 4950$$

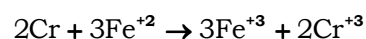
44. d) **anode is (-) ve in galvanic cell**

45. d) **above all**

all statement are true.

46. d) **0.26 V**

$$E_{\text{cell}} = E_{\text{cell}}^0 - \frac{0.0591}{6} \log \frac{[\text{Cr}^{+3}]^2}{[\text{Fe}^{+2}]^3}$$



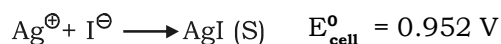
$$E_{\text{cell}}^0 = 0.30 \text{ V}$$

$$E_{\text{cell}} = 0.30 - \frac{0.06}{6} \log \frac{(0.1)^2}{(0.01)^3}$$

$$= 0.30 - 0.04$$

$$= 0.26 \text{ V}$$

47. d) **- 16.13**



$$E_{\text{cell}} = E_{\text{cell}}^0 - \frac{0.0591}{1} \log \frac{1}{[\text{Ag}^+][\text{I}^\ominus]}$$

at equilibrium, $E_{\text{cell}} = 0$

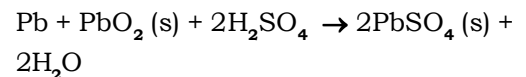
$$E_{\text{cell}}^0 = \frac{0.059}{1} \log \frac{1}{k_{sp}}$$

$$E_{\text{cell}}^0 = - \frac{0.0591}{1} \log k_{sp}$$

$$0.952 = - 0.059 \log k_{sp}$$

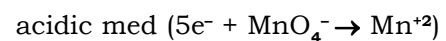
$$\log k_{sp} = - \frac{0.952}{0.059} = -16.13$$

48. b) **decreasing**



H₂SO₄ is consumed during discharge

49. d) **5**



(unbalanced)

5mol el are required to reduce one mole MnO₄⁻ change on 5 mole el = 5 Faradays

50. b) **84 g**

51. c) **0.1 M**

$$\Lambda_m (\text{molar conductivity}) = \frac{K (\text{sp. conductivity})}{\text{conc. (molarity)}}$$

$$\text{molarity} = \frac{0.0129 \Omega^{-1} \text{cm}^{-1}}{129 \Omega^{-1} \text{cm}^2 \text{mol}^{-1}}$$

$$= 10^{-4} \text{ mol cm}^{-3}$$

$$= 10^{-1} \text{ mol dm}^{-3}$$

52. a) **B > C > A**

B has highest tendency to get oxidised so it has highest reducing power.

53. d) **10¹⁰**

$$E^0 = + \frac{0.0591}{2} \log k_{eq}$$

$$\frac{0.295 \times 2}{0.0591} = \log k_{eq}$$

$$k_{eq} = 10^{10}$$

54. a) **128 S cm² mol⁻¹**

$$\Lambda_{\text{NaBr}}^0 = \Lambda_{\text{NaCl}}^0 + \Lambda_{\text{KBr}}^0 - \Lambda_{\text{KCl}}^0$$

$$= 126 + 152 - 150$$

$$= 128$$

55. **c)** 5×10^{-4}
 sp conductance = conductance \times cell constant
 cell constant = 1.4×50
 sp conductance = conductance \times cell constant

$$= \frac{1}{280} \times 1.4 \times 50$$

$$= 0.25 \Omega^{-1} \text{ cm}^{-1}$$
 molar conductance ($\text{S m}^2 \text{ mol}^{-1}$) =

$$\frac{\text{specific conductance} (\text{Sm}^{-1})}{\text{conductance} (\text{mol} / \text{dm}^3) \times 10^3 \text{ dm}^3 / \text{m}^3}$$

$$= \frac{0.25 \text{ Sm}^{-1}}{0.5 \times 10^3}$$

$$= 5 \times 10^{-4} \text{ Sm}^2 \text{ mol}^{-1}$$
56. **a)** $\lambda_c = \lambda_\infty - \mathbf{B}\sqrt{\mathbf{C}}$
 For strong electrolytes, this is the experimental result
57. **c)** **6 : 3 : 2**
 $\text{Ag}^+ + e^- \rightarrow \text{Ag}$ (3 mol Ag)
 $\text{Cu}^{+2} + 2e^- \rightarrow \text{Cu}$ (3/2 mol Cu)
 $\text{Ag}^{+3} + 3e^- \rightarrow \text{Au}$ (1 mol Au)
 28, 29, 30
58. **a)** **7.77 min**
 90gm Hg has 10 gm Na
 $\therefore 10\text{gm Hg} = \frac{10}{90} \times 10 = \frac{10}{9} \text{ gm Na}$
 $\therefore \text{weight of Na} = \frac{\mathbf{M}}{\mathbf{n}} \times \frac{\mathbf{i} \times \mathbf{t}}{96500}$

$$\frac{10}{9} = \frac{23}{1} \times \frac{10 \times t}{96500} [\because \text{Na}^+ + e^- \rightarrow \text{Na}]$$

 $\therefore t = \frac{10 \times 96500}{9 \times 10 \times 23}$
 7.77 min
59. **a)** **39**
 The solubility of AgBr in presence of 10^{-7} molar AgNO_3 is $3 \times 10^{-7} \text{M}$
 Therefore $[\text{Br}^-] = 3 \times 10^{-4} \text{ m}^3$, $[\text{Ag}^+] = 4 \times 10^{-4} \text{ m}^3$ and $[\text{NO}_3^-] = 10^{-4} \text{ m}^3$
 Therefore $k_{\text{total}} = k_{\text{Br}^-} + k_{\text{Ag}^+} + k_{\text{NO}_3^-}$
 $= 39 \text{ Sm}^{-1}$

60. **b)** $\frac{-\mathbf{E}_1 - \mathbf{E}_2}{\mathbf{0.118}}$
 $\text{H}^+ + e^- \Rightarrow \frac{1}{2} \text{H}_2 (\text{g})$

$$E_1 = 0 - 0.0591 \log \frac{1}{(\text{H}^+)_1}$$

$$E_1 = 0 - 0.0591 \log (\text{H}^+)_1$$

$$= - .0591 \text{ pH}_1$$

$$\text{pH}_1 = \text{pK}_a + \log \frac{\text{Salt}}{\text{Acid}}$$

$$\text{pH}_1 = \text{pK}_a + \log \frac{\mathbf{a}}{\mathbf{b}} \quad \dots(\text{i})$$

$$\text{pH}_2 = \text{pK}_a + \log \frac{\mathbf{b}}{\mathbf{a}}$$

$$\text{pH}_2 = \text{pK}_a + \log \frac{\mathbf{a}}{\mathbf{b}} \quad \dots(\text{ii})$$
 Add(i) and (ii)

$$\text{pH}_1 + \text{pH}_2 = 2\text{pK}_a$$

$$2\text{pK}_a = \frac{E_1}{0.0591} - \frac{E_2}{0.0591}$$

$$\Rightarrow \text{pK}_a = - \left[\frac{E_1 + E_2}{0.118} \right]$$