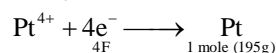


**PCM TEST- 2 MAIN
CHEMISTRY SOLUTION
R-3 BATCH**

31. PtCl_6^{2-} contains Pt^{4+} ion



195g of Pt will be deposited by $= 4 \times 96500\text{C}$

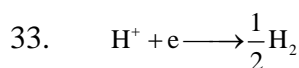
\therefore 0.5g of Pt will be deposited by

$$= \frac{4 \times 96500}{195} \times 0.5\text{C}$$

Now $Q = It$

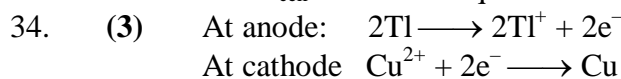
$$\therefore \text{Current (I)} = \frac{Q}{t} = \frac{4 \times 96500}{195} \times \frac{0.5}{5 \times 60 \times 60} = 0.055 \text{ ampere}$$

32. Calculate ΔG^0 for all these cells using the relation $E_{\text{cell}}^0 = E_{\text{cathode}}^0 - E_{\text{anode}}^0$. The cell in which E_{cell}^0 is more has more negative value of ΔG^0 .



$$E = E^0 - \frac{0.059}{1} \log \frac{[\text{P}_{\text{H}_2}]^{\frac{1}{2}}}{[\text{H}^+]} = 0 - \frac{0.059}{1} \log \frac{(100)^{\frac{1}{2}}}{1} = -0.059 \log 10 = -0.059\text{V}$$

Given that E_{cell}^0 at 308K is equal to that at 298K

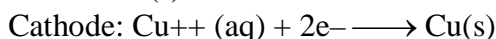
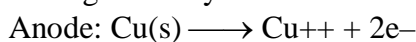


Using Nernst equation, we get

$$E_{\text{cell}} = E_{\text{cell}}^0 - \frac{2.303RT}{nF} \log \frac{[\text{Tl}^+]^2}{[\text{Cu}^{2+}]}$$

Thus, E_{cell} can be increased by decreasing $[\text{Tl}^+]$ or increasing $[\text{Cu}^{2+}]$ or by decreasing temperature.

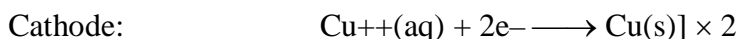
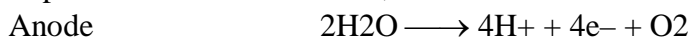
35. During electrolysis of CuSO_4 solution using Cu electrodes, the cell reaction is



The loss in weight of anode is equal to gain in weight of cathode.

\therefore (4)

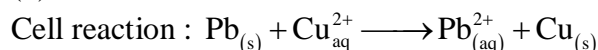
36. In presence of inert electrode, the cell reaction is



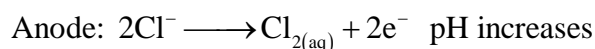
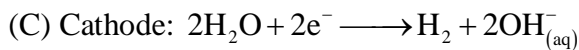
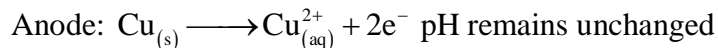
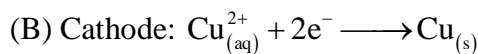
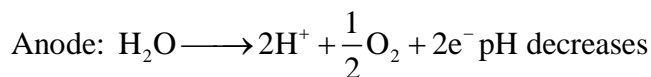
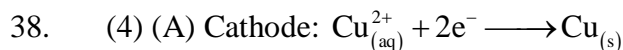
Due to increase in $[\text{H}^+]$, pH decreases.

\therefore (2)

37. (2)

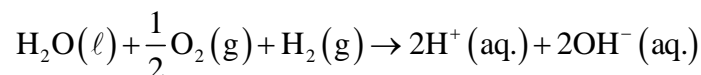
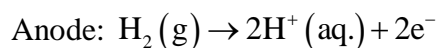
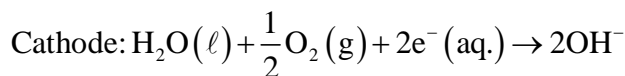


$$E_{\text{cell}} = E_{\text{cell}}^0 + \frac{0.0592}{2} \log \frac{[\text{Cu}^{2+}]}{[\text{Pb}^{2+}]}$$

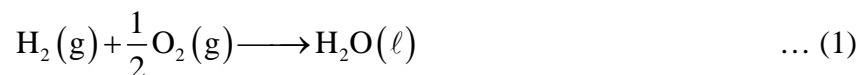


39. (1)

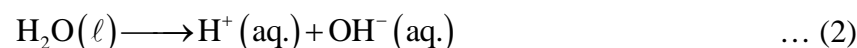
Cell reaction



Also we have



$\Delta_r G^0 = -237.20 \text{ kJ/mol}$



$\Delta G^0 = 80 \text{ kJ/mol}$

Hence for cell reaction [eq. (1) + eq. (2) × 2]

$\Delta G^0 = -237.20 + 80 \times 2$

$\Delta G^0 = -77.20 \text{ kJ/mol}$

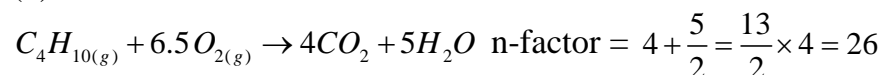
So, $E^0 = -\frac{\Delta G^0}{nF} = \frac{77200}{2 \times 96500} = 0.40 \text{ V}$

40. (4)

$$E_{\text{cell}} = -\frac{0.0591}{1} \log \frac{[\text{H}^+]_a}{[\text{H}^+]_c}$$

For E_{cell} to be highest $[\text{H}^+]_a$ should be lower and $[\text{H}^+]_c$ should be higher and that's why anode compartment should be more basic and cathodic compartment should be acidic

41. (4)



$\Delta G = -nFE_o$ $-nFE_o = -2746$ $E_o = \frac{2746 \times 10^3}{26 \times 96500} = 1.09 \text{ V}$

42. (4)

Number of moles of Na^+ reduced at cathode = $\frac{MV}{1000} = \frac{4 \times 500}{1000} = 2$

Mass of amalgam = Mass of 2 mol Na + Mass of 2 mol Hg

$$= 2 \times 23 + 2 \times 200 = 446 \text{g}$$

43. (2) $E = \frac{0.06}{2} \log \frac{1}{10} = -0.03 \text{ V.}$



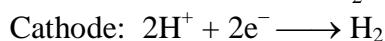
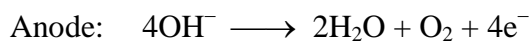
So, (ii) – (i)



so $E^0 = -\left(\frac{-0.772F}{1 \times F}\right) = 0.772 \text{ Volt.}$

45. (2) As the $[\text{H}^+]$ increases, the reduction potential of hydrogen electrode increases. The reduction potential of anode half-cell is greater than that of cathode half-cell. The emf of the given cell will be negative. Therefore, the cell reaction is **non-spontaneous**.

46. (4) During electrolysis of an aqueous solution of Na_2SO_4 , the following reactions occur at the electrodes.



Number of faradays needed to liberate 1 mole of $\text{O}_2 = 4$

$$\text{Charge} = \frac{2.4 \times 4}{22.4} F$$

Number of moles of H_2 liberated by 1 F of charge = $\frac{1}{2}$

$$\text{Number of moles of } \text{H}_2 = \frac{2.4 \times 4}{2 \times 22.4}$$

$$\text{Volume of } \text{H}_2 \text{ at STP} = \frac{2.4 \times 4 \times 22.4}{2 \times 22.4} = \mathbf{4.8 \text{ L.}}$$

47. (3) Number of Faraday = number of equivalents

$$\frac{2 \times 5 \times 60 \times 60}{96500} = n \times \frac{22.2}{177}$$

$$\therefore n = \mathbf{3.}$$

48. (4) The electrode potential of the half cell $\text{MnO}_4^-, \text{Mn}^{2+} | \text{Pt}$ is given by

$$E_{\text{MnO}_4^- | \text{Mn}^{2+}} = E_{\text{MnO}_4^- | \text{Mn}^{2+}}^0 - \frac{0.059}{5} \log \frac{[\text{Mn}^{2+}]}{[\text{MnO}_4^-][\text{H}^+]^8}$$

When the concentration of H^+ is reduced to half of its initial value the electrode potential is given by

$$E'_{\text{MnO}_4^- | \text{Mn}^{2+}} = E_{\text{MnO}_4^- | \text{Mn}^{2+}}^0 - \frac{0.059}{5} \log \frac{[\text{Mn}^{2+}] 2^8}{[\text{MnO}_4^-][\text{H}^+]^8}$$

The decrease in electrode potential is given by

$$E_{\text{MnO}_4^- | \text{Mn}^{2+}} - E'_{\text{MnO}_4^- | \text{Mn}^{2+}} = \frac{0.0591}{5} \log 2^8 = 0.02846 \text{ V.}$$

∴ Electrode potential decrease, by 28.46 mV.

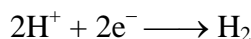
49. (3) When copper is deposited at cathode, oxygen gas is released at anode.

Equivalents of Copper = $0.1 \times 2 = 0.2$

Equivalents of Oxygen = 0.2

Volume of oxygen at NTP = $0.2 \times 5600 = 1120 \text{ ml}$.

50. (1) The half cell reaction for hydrogen half cell acting as cathode is



$$\therefore E_{\text{H}^+|\text{H}_2} = E_{\text{H}^+|\text{H}_2}^\circ - \frac{0.059}{2} \log \frac{P_{\text{H}_2}}{[\text{H}^+]^2}$$

$$E_{\text{H}^+|\text{H}_2} = -\frac{0.059}{2} \log \frac{1}{[\text{H}^+]^2}$$

Now, when the pressure of H_2 gas is changed to 100 atm without changing $[\text{H}^+]$, the reduction potential becomes

$$\therefore E'_{\text{H}^+|\text{H}_2} = -\frac{0.059}{2} \log \frac{100}{[\text{H}^+]^2}$$

$$\begin{aligned} \therefore \text{Change in reduction potential} &= E'_{\text{H}^+|\text{H}_2} - E_{\text{H}^+|\text{H}_2} = \frac{-0.059}{2} \log \left(\frac{100}{[\text{H}^+]^2} \times \frac{[\text{H}^+]^2}{1} \right) \\ &= -0.059 \text{ V} \end{aligned}$$

51. (3) Aqueous solution of LiCl contains Li^+ , H^+ , Cl^- and OH^- ions. H^+ would be preferentially reduced at cathode while at anode, OH^- ions would be oxidised, so H_2 and O_2 gases are liberated at cathode and anode respectively. As the current flows, H^+ ions around the cathode gets discharged to give H_2 gas, so the pH of the solution around the cathode increases.

52. (1) $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$
(from AgCl)

$$E_1 = E^\circ - \frac{0.0591}{1} \log \frac{1}{C_1}$$

Where C_1 is the concentration of Ag^+ in saturated AgCl solution.

$$E_2 = E^\circ - \frac{0.0591}{1} \log \frac{1}{C_2}$$

Where C_2 is the concentration of Ag^+ in saturated AgI solution.

$$E_1 - E_2 = 0.177 = +0.059 \log C_1/C_2$$

$$\log C_1/C_2 = 3$$

$$C_1/C_2 = 10^3.$$

- 53 Vapour pressure of pure acetone (p^0) = 185 torr
 Vapour pressure of solution (p) = 183 torr
 Then from Raoult's law,

$$\frac{p^0 - p}{p} = \frac{n_{\text{solute}}}{n_{\text{solvent}}}$$

$$\frac{185 - 183}{183} = \frac{1.2 / \text{MW}}{100 / 58}$$

$$\frac{2}{183} = \frac{1.2 \times 58}{100 \times \text{MW}}$$

$$\text{MW} = \frac{1.2 \times 58 \times 183}{200} = 63.68 \approx 64$$

54. (4)

$$r = k[A][B]^2$$

55. (4)

$$(I) \quad K_{f_1} = A_I e^{\frac{-E_{af_1}}{RT}} \quad K_{f_2} = A_{II} e^{\frac{-E_{af_2}}{RT}} \quad \frac{K_{f_1}}{K_{f_2}} = \frac{A_I}{A_{II}} e^{\frac{-(E_{af_1} - E_{af_2})}{RT}}$$

$$\therefore E_{af_1} - E_{af_2} \neq 0 \quad \therefore \frac{K_{f_1}}{K_{f_2}} \text{ is dependent on } T$$

$$(II) \quad K_{eq_1} = A_I e^{\frac{-\Delta H_1}{RT}} \quad K_{eq_2} = A_{II} e^{\frac{-\Delta H_2}{RT}}$$

$$K_{eq_1} \times K_{eq_2} = A_I A_{II} e^{\frac{-(\Delta H_1 + \Delta H_2)}{RT}}$$

$$\Delta H_1 = +600 \quad \Delta H_2 = -600 \quad \therefore \Delta H_1 + \Delta H_2 = 0$$

\therefore Independent of Temp.

$$(III) \quad K_{eq} = \frac{K_f}{K_b} = \frac{A e^{\frac{-E_{af}}{RT}}}{A e^{\frac{-E_{ab}}{RT}}} = e^{\frac{-(E_{af} - E_{ab})}{RT}}$$

$$\therefore E_{af} = 800 \quad E_{ab} = 200$$

$$\Rightarrow K_{eq} = e^{\frac{-(800 - 200)}{2 \times 300}} = e^{-1}$$

$$(IV) \quad K_{eq} = e^{\frac{-(E_{af} - E_{ab})}{RT}}$$

$$\therefore E_{af} = 200 \quad E_{ab} = 800$$

$$\Rightarrow e^{\frac{-(200 - 800)}{2 \times 300}} = e^1$$

Correct answer is II and IV . Which is not in the options

\therefore Ans : D

56. (3)

Lower boiling point implies positive deviation from Roults law.

57. (2)

$$[\text{Protein}] = \left(\frac{0.75\text{g}}{\text{Mol.wt}} \right) \times \frac{1000}{125\text{cm}^3} = \left(\frac{0.75 \times 8}{\text{mol.wt}} \right) \text{mol/L}$$

$$\text{Osmotic pressure} = \frac{2.6}{760} \text{atm} = \text{C.R.T} = \left(\frac{0.75 \times 8}{\text{Mol.Wt.}} \right) \times 0.0821 \times 277\text{K}$$

On solving, Mol.wt = 3.9×10^4 .

58 (2)

$$\Delta T_f = k_f \times m \times s$$

$$\Rightarrow 0.2046 = 1.86 \times 0.1 \times (1 + \alpha)$$

$$\Rightarrow (1 + \alpha) = \frac{0.2046}{0.186} = 1.1$$

$$\Rightarrow \alpha = 0.1$$

$$[\text{H}^+] = \text{C}\alpha = 0.1 \times 0.1 = 10^{-2}$$

$$\therefore \text{pH} = 2$$

59. (1)

60. (1)

$$t_{1/2} \propto \frac{1}{a^{n-1}}$$

$$\frac{3}{6} = \left(\frac{1}{0.5} \right)^{n-1}$$

$$\therefore n = 0$$