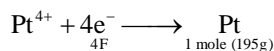


**PCM TEST- 4 MAIN  
CHEMISTRY SOLUTION  
R-1 & R-2 BATCH**

31.  $\text{PtCl}_6^{2-}$  contains  $\text{Pt}^{4+}$  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  $\Delta G^0$  for all these cells using the relation  $E_{\text{cell}}^0 = E_{\text{cathode}}^0 - E_{\text{anode}}^0$ . The cell in which  $E_{\text{cell}}^0$  is more has more negative value of  $\Delta G^0$ .

33.  $\text{H}^+ + e^- \longrightarrow \frac{1}{2}\text{H}_2$

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

34. (3) At anode:  $2\text{Tl} \longrightarrow 2\text{Tl}^+ + 2e^-$   
At cathode  $\text{Cu}^{2+} + 2e^- \longrightarrow \text{Cu}$

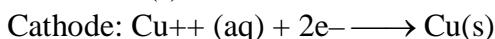
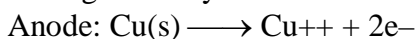


Using Nernst equation, we get

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

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

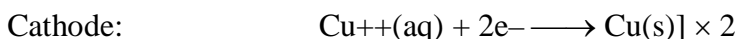
35. During electrolysis of  $\text{CuSO}_4$  solution using Cu electrodes, the cell reaction is



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.

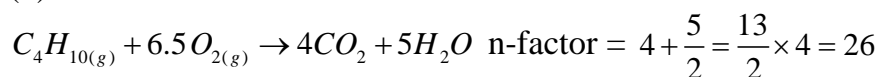
$\therefore$  (2)

37. (4)

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

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

38. (4)



$$\Delta G = -nFE_0 \quad -nFE_0 = -2746 \quad E_0 = \frac{2746 \times 10^3}{26 \times 96500} = 1.09V$$

39. (4)

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

$$\begin{aligned} \text{Mass of amalgam} &= \text{Mass of 2 mol Na} + \text{Mass of 2 mol Hg} \\ &= 2 \times 23 + 2 \times 200 = 446g \end{aligned}$$

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

41. (4) (i)  $Fe^{+2} + 2e^- \longrightarrow Fe$  so  $\Delta G_1^0 = -2 \times F(-0.44) = 0.88 F$

(ii)  $Fe^{+3} + 3e^- \longrightarrow Fe$   $\Delta G_2^0 = -3 \times F \times (-0.036) = 0.108 F$

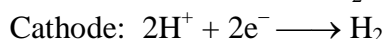
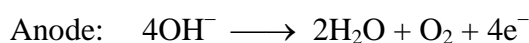
So, (ii) - (i)

$$\begin{aligned} Fe^{+3} + e^- \longrightarrow Fe^{+2} \quad \Delta G_3^0 &= .108 F - .88 F \\ &= -0.772F \end{aligned}$$

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

42. (2) As the  $[H^+]$  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**.

43. (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  $O_2 = 4$

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

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

$$\text{Number of moles of } H_2 = \frac{2.4 \times 4}{2 \times 22.4}$$

$$\text{Volume of } H_2 \text{ at STP} = \frac{2.4 \times 4 \times 22.4}{2 \times 22.4} = \mathbf{4.8 \text{ L.}}$$

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

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

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

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

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

When the concentration of  $\text{H}^+$  is reduced to half of its initial value the electrode potential is given by

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

The decrease in electrode potential is given by

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

$\therefore$  Electrode potential decrease, by 28.46 mV.

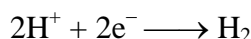
46. (3) When copper is deposited at cathode, oxygen gas is released at anode.

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

Equivalent of Oxygen = 0.2

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

47. (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  $\text{H}_2$  gas is changed to 100 atm without changing  $[\text{H}^+]$ , the reduction potential becomes

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

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

48. (3) Aqueous solution of  $\text{LiCl}$  contains  $\text{Li}^+$ ,  $\text{H}^+$ ,  $\text{Cl}^-$  and  $\text{OH}^-$  ions.  $\text{H}^+$  would be preferentially reduced at cathode while at anode,  $\text{OH}^-$  ions would be oxidised, so  $\text{H}_2$  and  $\text{O}_2$  gases are liberated at cathode and anode respectively. As the current flows,  $\text{H}^+$  ions around the cathode gets discharged to give  $\text{H}_2$  gas, so the pH of the solution around the cathode increases.

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

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

Where  $C_1$  is the concentration of  $\text{Ag}^+$  in saturated  $\text{AgCl}$  solution.

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

Where  $C_2$  is the concentration of  $\text{Ag}^+$  in saturated  $\text{AgI}$  solution.

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

$$\log C_1/C_2 = 3$$

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

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

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

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

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

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

51. (4)

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

52. (4)

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

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

$$(II) \quad Keq_1 = A_I e^{\frac{-\Delta H_1}{RT}} \quad Keq_2 = A_{II} e^{\frac{-\Delta H_2}{RT}}$$

$$Keq_1 \times Keq_2 = A_I A_{II} e^{\frac{-(\Delta H_1 + \Delta H_2)}{RT}}$$

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

$\therefore$  Independent of Temp.

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

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

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

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

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

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

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

∴ Ans : D

53. (4)

(a) Fact

(b) 'O' is in T-void coordination no : 4

∴ 'U' will be have coordination no.: 8

(c) 'O' atoms = 6 (HCP)

$$\text{'Al' atom} = 6 \times \frac{2}{3} = 4$$

∴ Al<sub>2</sub>O<sub>3</sub>

∴ (D) option is correct

54. (3)

Lower boiling point implies positive deviation from Rault's law.

55. (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 =  $3.9 \times 10^4$ .

56 (2)

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

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

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

$$\Rightarrow \alpha = 0.1$$

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

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

57.

(4)

$$4r_y = \sqrt{2}a$$

$$\Rightarrow 2r_y = \frac{\sqrt{2}}{4} a = \frac{\sqrt{2}}{2} (a)$$

58. (1)

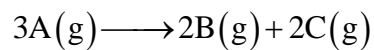
59. (1)

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

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

$\therefore n = 0$

**60. (A)**



$t = 0$	$P_0$	—	—
$t = 20 \text{ min}$	$P_0 - 3P$	$2P$	$2P$
$t = \infty$	—	$\frac{2}{3}P_0$	$\frac{2}{3}P_0$

Given at  $t = \infty$ ,  $P_\infty = 4 \text{ atm} = \frac{4P_0}{3}$

$\therefore P_0 = 3 \text{ atm}$

and at  $t = 20 \text{ min}$ ,  $P_{\text{total}} = 3.5 \text{ atm} = P_0 + P$

$\therefore 3.5 = 3 + P$

or  $P = 0.5 \text{ atm}$

$\Rightarrow 3P = 1.5 \text{ atm}$  and  $P_0 = 3 \text{ atm}$

$\therefore t = 20 \text{ min}$  is  $t_{1/2}$