

SOLUTION TEST-B

- Dilute H_2SO_4 liberates O_2 at anode
 aq. $\text{CuCl}_2 \rightarrow \text{Cl}_2$ at anode
 Conc. $\text{H}_2\text{SO}_4 \rightarrow \text{S}_2\text{O}_8^{2-}$ at anode
 molten $\text{NaCl} \rightarrow \text{Cl}_2$ at anode
 aq. $\text{AgNO}_3 \rightarrow \text{O}_2$ at anode.

The correct answer is (iv)

- $\text{Pb} + \text{SO}_4^{2-} \rightarrow \text{PbSO}_4 + 2\text{e}^-$ Anode
 $\text{PbO}_2 + 4\text{H}^+ + \text{SO}_4^{2-} + 2\text{e}^- \rightarrow \text{PbSO}_4 + 2\text{H}_2\text{O}$ Cathode
 $\text{Pb} + \text{PbO}_2 + 2\text{H}_2\text{SO}_4 \rightarrow 2\text{PbSO}_4 + 2\text{H}_2\text{O}$
 Number of moles of H_2SO_4 consumed = Number of moles of electrons transferred or Number of faradays.

$$\text{Number of faradays} = \frac{9.8}{98} = 0.1 \text{ F}$$

$$\text{Number of coulombs} = 0.1 \times 96500 = 9650 \text{ C}$$

The correct answer is (ii)

- Anode: $\text{H}_2(\text{g}) \rightarrow 2\text{H}^+_{(\text{aq})} + 2\text{e}^-$
 Cathode: $\text{Cl}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{Cl}^-_{(\text{aq})}$
 Net reaction: $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{H}^+_{(\text{aq})} + 2\text{Cl}^-_{(\text{aq})}$

The correct answer is (iii)

- $\text{Zn}^{+4} + 2\text{MnO}_2 + 2\text{NH}_4^+ \longrightarrow \text{Zn}^{2+} + 2\text{NH}_3 + 2\text{MnO}(\text{OH})^{+3}$

The oxidation state of Mn changes from +4 and +3 when dry cell is in use and when it act as a galvanic cell.

The correct answer is (ii)

- E° is an intensive quantity and is independent of the size of the system or mass independent.

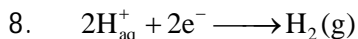
$$\text{Therefore, } E^\circ_{\text{Ag}^+/\text{Ag}} = E^\circ_{2\text{Ag}^+/2\text{Ag}}$$

The correct answer is (i)

- Faraday's second law of electrolysis states that the same number of grams equivalents are produced when same quantity of electricity is passed through different cells connected in series.

- (i) $\Delta_r G^\circ = -nFE^\circ_{\text{cell}}$

$$\text{(ii) } \log K = \frac{nE^\circ}{0.059}$$



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

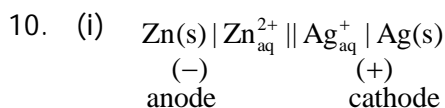
$$E_{\text{H}^+/\text{H}_2} = 0 - \frac{0.059}{2} \log [\text{H}^+]^2$$

$$E_{\text{H}^+/\text{H}_2} = -\frac{0.059}{2} \times 2(-\log[\text{H}^+]) = -0.059 \times \text{pH} = -0.059 \times 10$$

$$E_{\text{H}^+/\text{H}_2} = -0.59 \text{ V}$$

9. (i) When $C \rightarrow 0$ (at infinite dilution) $\alpha \rightarrow 1$, the electrolyte is completely dissociated, ions are infinitely apart and each ion moves independently of its co-ion. So, there is a steep rise in molar conductance.

(ii) This is because at cathode, reduction takes place and $E_{\text{Cu}^{2+}/\text{Cu}}^\circ$ is greater than $E_{\text{H}^+/\text{H}_2}^\circ$



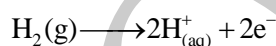
(ii) The direction of flow of electrons is from anode to cathode.

(iii) $w_{\text{max}} = -nFE^\circ$

$$w_{\text{max}} = -2 \times 96500 \times 15.6$$

$$w_{\text{max}} = -301080 \text{ J}$$

11. (i) Negative electrode means anode and at anode, oxidation takes place.



(ii) In cathodic protection, we protect iron from rusting by connecting it to more electropositive metal (Mg, Al). The more electropositive metal act as a anode and gets oxidised, Fe acts as a cathode.

In galvanization, it is a coating of iron with zinc. Zinc gets oxidised and form a protective layer of ZnO over iron surface.

(iii) Molar conductance value at infinite dilution when $C \rightarrow 0$ is known as the limiting value of the molar conductance.

12. $I = 100 \text{ A}, \quad w_{\text{B}} = 1.08 \text{ g}$

(i) $m = Zit$

$$1.08 = \frac{108}{1 \times 96500} \times 100 \times t$$

$$t = 9.65 \text{ s}$$

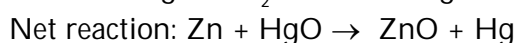
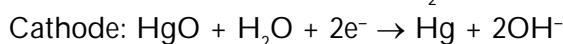
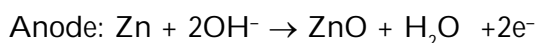
(ii) Using Faraday's second law $\frac{W_{Ag}}{EW_{Ag}} = \frac{W_{O_2}}{EW_{O_2}}$

$$\frac{1.08}{108/1} = \frac{W_{O_2}}{MW_{O_2/4}} = n_{O_2} \times 4$$

$$n_{O_2} = \frac{1.08}{108 \times 4} = 2.5 \times 10^{-3} \text{ mol}$$

$$\text{Volume of } O_2 \text{ at STP} = 2.5 \times 10^{-3} \times 22.4 = 0.056 \text{ L}$$

13. Primary cells are those cells which cannot be recharged while secondary cells are rechargeable
Zinc-mercuric oxide Cell.



14. $E_{cell}^{\circ} = E_{Cl_2/Cl^-}^{\circ} - E_{H^+/H_2}^{\circ}$

$$E_{cell}^{\circ} = +1.36 \text{ V}$$

$$Q = [H^+]^2 [Cl^-]^2$$

$$Q = (0.02)^2 (0.02)^2 = (2 \times 10^{-2} \times 2 \times 10^{-2})^2$$

$$Q = (4 \times 10^{-4})^2$$

$$E_{cell} = E_{cell}^{\circ} = \frac{0.059}{n} \log Q$$

$$E_{cell} + 1.36 - \frac{0.059}{2} \log(4 \times 10^{-4})^2$$

$$= 1.36 - 0.059 \log 4 \times 10^{-4} = 1.36 - 0.059 [\log 4 - 4] = 1.36 - 0.059 [0.602 - 4]$$

$$= 1.36 + 0.059 \times 3.398 \quad E_{cell} = +1.560 \text{ V}$$

15. (a) $\wedge = \frac{10^3 \kappa}{m} = \frac{10^3 \times 1.63 \times 10^{-4}}{0.01}$

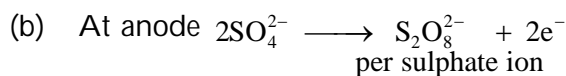
$$\wedge = 16.3 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\wedge^{\circ} = \wedge_{CH_3COONa}^{\circ} + \wedge_{HCl}^{\circ} - \wedge_{NaCl}^{\circ}$$

$$\wedge^{\circ} = 391 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\alpha = \frac{\wedge}{\wedge^{\circ}} = \frac{16.3}{391} = 0.042$$

% dissociation of ethanoic acid is 4.2%



- (c) Conductivity is the measure of the ease with which the current can be passed through a given electrolytic solution. The relation between conductivity and conductance is
conductivity = (cell constant) \times (conductance)