SOLUTION TEST-B

1. Dilute H_2SO_4 liberates O_2 at anode

   aq. $CuCl_2 \rightarrow Cl_2$ at anode

   Conc. $H_2SO_4 \rightarrow S_2O_8^{2-}$ at anode

   molten $NaCl \rightarrow Cl_2$ at anode

   aq. $AgNO_3 \rightarrow O_2$ at anode.

   The correct answer is (iv)

2. $Pb + SO_4^{2-} \rightarrow PbSO_4 + 2e^-$ Anode

   $PbO_2 + 4H^+ + SO_4^{2-} + 2e^- \rightarrow PbSO_4 + 2H_2O$ Cathode

   $Pb + PbO_2 + 2H_2SO_4 \rightarrow 2PbSO_4 + 2H_2O$

   Number of moles of H_2SO_4 consumed = Number of moles of electrons transferred or Number of faradays.

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

   Number of coulombs = 0.1 × 96500 = 9650 C

   The correct answer is (ii)

3. Anode: $H_2(g) \rightarrow 2H^+_{(aq)} + 2e^-$

   Cathode: $Cl_2(g) + 2e^- \rightarrow 2Cl^-_{(aq)}$

   Net reaction: $H_2(g) + Cl_2(g) \rightarrow 2H^+_{(aq)} + 2Cl^-_{(aq)}$

   The correct answer is (iii)

4. $$Zn + 2\overset{+4}{Mn}O_2 + 2NH_4^+ \longrightarrow Zn^{2+} + 2NH_3 + 2\overset{+3}{Mn}O(OH)$$

   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)

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

   Therefore, $E^o_{Ag^+/Ag} = E^o_{2Ag^+/2Ag}$

   The correct answer is (i)

6. 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.

7. (i) $\Delta_r G° = -nFE^o_{cell}$

   (ii) $\log K = \frac{nE°}{0.059}$

8. $2H^+_{aq} + 2e^- \longrightarrow H_2(g)$

$$E_{H^+/H_2} = E^o_{H^+/H_2} - \frac{0.059}{2}\log\frac{1}{[H^+]^2}$$

$$E_{H^+/H_2} = 0 - \frac{0.059}{2}\log[H^+]^2$$

$$E_{H^+/H_2} = -\frac{0.059}{2}\times 2(-\log[H^+]) = -0.059\times pH = -0.059\times 10$$

$$E_{H^+/H_2} = -0.59\,V$$

9. (i) When C → 0 (at infinite dilution) $\alpha \to 1$, 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 $E^o_{Cu^{2+}/Cu}$ is greater than $E^o_{H^+/H_2}$

10. (i)

$$\underset{\substack{(-)\\ \text{anode}}}{Zn(s)}\,|\,Zn^{2+}_{aq}\,||\,Ag^+_{aq}\,|\,\underset{\substack{(+)\\ \text{cathode}}}{Ag(s)}$$

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

   (iii) $w_{max} = -nFE°$

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

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

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

$$H_2(g) \longrightarrow 2H^+_{(aq)} + 2e^-$$

   (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 → 0 is known as the limiting value of the molar conductance.

12. I = 100 A, w_B = 1.08 g

   (i) m = ZIt

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

   t = 9.65 s

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

$$\frac{1.08}{108/1} = \frac{w_{O_2}}{MW_{O_{2/4}}} = n_{O_2} \times 4$$

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

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

13. Primary cells are those cells which cannot be recharged while secondary cells are rechargeable

Zinc-mercuric oxide Cell.

Anode: $Zn + 2OH^- \rightarrow ZnO + H_2O\ \ +2e^-$

Cathode: $HgO + H_2O + 2e^- \rightarrow Hg + 2OH^-$

Net reaction: $Zn + HgO \rightarrow\ ZnO + Hg$

14. $E^o_{cell} = E^o_{Cl_2/Cl^-} - E^o_{H^+/H_2}$

$$E^o_{cell} = +1.36\,V$$

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

$$Q = (0.02)^2(0.02)^2\ \ = (2\times10^{-2} \times 2\times10^{-2})^2$$

$$Q = (4\times10^{-4})^2$$

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

$$E_{cell} + 1.36 - \frac{0.059}{2}\log(4\times10^{-4})^2$$

$$= 1.36 - 0.059\log 4\times10^{-4}\ \ = 1.36 - 0.059\,[\log 4 - 4]\ \ = 1.36 - 0.059\,[0.602 - 4]$$

$$= 1.36 + 0.059 \times 3.398 \qquad E_{cell} = +1.560\,V$$

15. (a) $\wedge = \frac{10^3 \kappa}{m} = \frac{10^3 \times 1.63\times10^{-4}}{0.01}$

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

$$\wedge^o = \wedge^o_{CH_3COONa} + \wedge^o_{HCl} - \wedge^o_{NaCl}$$

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

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

% dissociation of ethanoic acid is 4.2%

(b) At anode $2SO_4^{2-} \xrightarrow{\quad} \underset{\text{per sulphate ion}}{S_2O_8^{2-}} + 2e^-$

At cathode $2H^+ + 2e^- \longrightarrow H_2$

(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) × (conductance)