Unit Practice Test

for Board Examination

Time allowed : 2 Hrs. *Maximum Marks : 35*

1. Give the units of conductivity and molar conductivity. (1)

2. Write the overall cell reaction for lead storage battery. (1)

3. Define Faraday's second law of electrolysis. (1)

4. What is the role of $ZnCl_2$ in a dry cell ? (1)

5. Out of copper and zinc vessels, which vessel would be suitable for storing 1M HCl? (1)

6. Predict the products of electrolysis in each of the following : (2)

(i) An aqueous solution of $AgNO_3$ with silver electrode.

(ii) A dilute solution of H_2SO_4 with platinum electrodes.

7. How much electricity is required in Coulombs to produce 40 *g* of Al from molten Al_2O_3? (2)

8. Conductivity of 0.00241 M acetic acid is 7.896×10^{-5} S cm^{-1}. Calculate its molar conductivity. If $\Lambda_m°$ for acetic acid is 390.5 S cm^2 mol^{-1}, what would be its dissociation constant? (2)

9. Write cell reactions which occur in lead storage battery

(i) When battery is in use

(ii) When the battery is on charging. (2)

10. Give two points of differences between emf and potential difference . (2)

11. Write the Nernst equation and calculate e.m.f. of the following cell at 298 K :

$$Sn(s) \mid Sn^{2+}(0.050M) \parallel H^+(0.020\ M) \mid H_2\ (1\ atm)\ Pt.$$ (3)

12. Explain the following :

(i) Electrical protection for preventing rusting of iron pipes in underground water.

(ii) Can you store copper sulphate solution in a zinc pot or not?

(iii) Effect of dilution on molar conductivity. (3)

13. Write the chemical equations for all the steps involved in the rusting of iron. Explain why does alkaline medium inhibits the rusting of iron. (3)

14. State Kohlrausch law of independent migration of ions. Mention one application of the law. (3)

15. List main differences between electrochemical cells and electrolytic cells ? (3)

16. *(a)* The cell in which the following reaction occurs :

$$2Fe^{3+}(aq) + 2I^-(aq) \longrightarrow 2Fe^{2+}(aq) + I_2$$

has $E°_{cell} = 0.236$ V at 298 K. Calculate the standard Gibbs energy and the equilibrium constant of the cell reaction.

(b) What are fuel cells ? Give one example.

(c) Give the units of cell constant. (5)

► To check your performance, see HINTS AND SOLUTIONS TO SOME QUESTIONS at the end of Part I of the book.

UNIT 3 : ELECTROCHEMISTRY

1. S cm^{-1}, S cm^2 mol^{-1}
2. $Pb(s) + PbO_2(s) + 2H_2SO_4\,(aq) \longrightarrow 2\,PbSO_4\,(s) + 2H_2O\,(l)$
3. When same quantity of electricity is passed through different electrolytic solutions connected in series, the weights of the substances produced at the electrodes are directly proportional to their chemical equivalent weights.
4. $ZnCl_2$ combines with NH_3 produced to form the complex $[Zn(NH_3)_2Cl_2]$, otherwise the pressure developed due to NH_3 would crack the seal of the cell.

7. $$Al^{3+} + 3e^- \longrightarrow Al$$

$$\text{27 g of Al require electricity} = 3 \times 96500\ \text{C}$$

$$\text{50 g of Al will require electricity} = \frac{3 \times 96500}{27} \times 50$$

$$= 5.36 \times 10^5\ \text{C}$$

8. $$\Lambda_m = \frac{\kappa \times 1000}{M} = \frac{7.896 \times 10^{-5}\,\text{S cm}^{-1} \times 1000\ \text{cm}^3\ \text{L}^{-1}}{0.00241\ \text{mol L}^{-1}}$$

$$= 32.76\ \text{S cm}^2\ \text{mol}^{-1}$$

$$\alpha = \frac{\Lambda_m^C}{\Lambda_m^\circ} = \frac{32.76}{390.5} = 8.39 \times 10^{-2}$$

$$K_a = \frac{C\alpha^2}{1-\alpha} = \frac{0.00241 \times (8.39 \times 10^{-2})^2}{(1 - 8.39 \times 10^{-2})}$$

$$= \mathbf{1.86 \times 10^{-5}}$$

11. The cell is :

$Sn(s)\ |\ Sn^{2+}(0.050M)\ ||\ H^+(0.020M)\ |\ H_2(1\ atm)\ |\ Pt$

The electrode reactions and cell reactions are :

$$Sn(s) \longrightarrow Sn^{2+}(aq) + 2e^- \quad \text{(At anode)}$$

$$2H^+(aq) + 2e^- \longrightarrow H_2(g) \quad \text{(At cathode)}$$

$$Sn\,(s) + 2H^+(aq) \longrightarrow Sn^{2+}(aq) + H_2(g) \quad \text{(Overall reaction)}$$

The reaction involves 2 moles of electrons, therefore, $n = 2$ and the Nernst equation is :

$$E = E^\ominus - \frac{0.059}{2} \log \frac{[Sn^{2+}]}{[H^+]^2}$$

$$E^\ominus_{cell} = E^\ominus(H^+\ |\ H_2) - E^\ominus(Sn^{2+}\ |\ Sn)$$

$$E^\ominus(H^+\ |\ H_2) = 0.0\ V,\ E^\ominus(Sn^{2+}\ |\ Sn) = -\,0.14\ V$$

$$\therefore \quad E^\ominus_{cell} = 0.00 - (-\,0.14) = 0.14\ V$$

$$[H^+] = 0.020\ M,\ [Sn^{2+}] = 0.050\ M$$

$$\therefore \quad E = 0.14 - \frac{0.059}{2} \log \frac{(0.050)}{(0.020)^2}$$

$$= 0.14 - 0.06 = \mathbf{0.08\ V.}$$

16 . (*a*)

$$E^\ominus_{cell} = 0.236\ V$$

$$\Delta G^\ominus = -\,nFE^\ominus$$

$$n = 2,\ F = 96500C$$

$$\Delta G^\ominus = 2 \times (96500\ C) \times (0.236\ V)$$

$$= -\,45548\ J \text{ or } = -\,45.55\ kJ$$

$$\Delta G^\ominus = -\,2.303\ RT \log K_c$$

or

$$\log K_c = -\frac{\Delta G}{2.303\ RT} = -\frac{-\,45.55}{2.303 \times 8.314 \times 10^{-3} \times 298} = 7.983$$

$$K_c = \text{antilog}\ (7.983) = \mathbf{9.62 \times 10^7}$$