

## Short Answer Questions-II

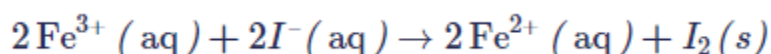
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### Short Answer Questions-II (PYQ)

Q.1. Answer the following questions

[CBSE (AI) 2017]

Q. The cell in which the following reaction occurs:



has  $E_{\text{cell}}^0 = 0.236 \text{ V}$  at 298 K. Calculate the standard Gibbs energy of the cell reaction.

(Given :  $1 \text{ F} = 96,500 \text{ C mol}^{-1}$ )

Ans.

Given,  $n = 2$ ,  $E_{\text{cell}}^0 = 0.236 \text{ V}$ ,  $\text{F} = 96,500 \text{ C mol}^{-1}$

$$\therefore \Delta G^{\circ} = -nFE_{\text{cell}}^{\circ}$$

$$= -2 \times 96500 \text{ C mol}^{-1} \times 0.236 \text{ V} = -45548 \text{ J mol}^{-1} = -45.548 \text{ kJ mol}^{-1}$$

Q. How many electrons flow through a metallic wire if a current of 0.5A is passed for 2 hours?

(Given :  $1 \text{ F} = 96,500 \text{ C mol}^{-1}$ )

[CBSE (AI) 2017]

Ans.  $Q = I \times t = 0.5 \text{ A} \times 2 \times 60 \times 60 \text{ s} = 3600 \text{ C}$

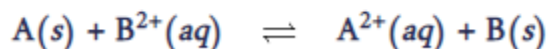
Number of electrons flowing through the wire on passing a charge of 96500 C =  $6.022 \times 10^{23}$

$\therefore$  Number of electrons flowing through the wire on passing a charge of 3600 C

$$= \frac{6.022 \times 10^{23} \times 3600 \text{ C}}{96500 \text{ C}} = 2.246 \times 10^{22} \text{ electrons}$$

**Q.2.**

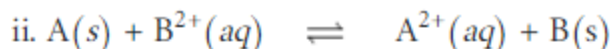
- i. Write two advantages of  $H_2 - O_2$  fuel cell over ordinary cell.
- ii. Equilibrium constant ( $K_c$ ) for the given cell reaction is 10. Calculate  $E_{\text{cell}}^o$ .



[CBSE (F) 2014]

**Ans. (i)** The two main advantages of  $H_2-O_2$  fuel cell over ordinary cell are as follows:

- It has high efficiency of 60%–70%.
- It does not cause any pollution.



Here,  $n = 2$ ,  $K_c = 10$

$$E_{\text{cell}}^o = \frac{0.059}{n} \times \log K_c$$

$$E_{\text{cell}}^o = \frac{0.059}{n} \times \log 10 = \frac{0.059}{2} \times 1 = 0.0295 \text{ V}$$

**Q.3.** Conductivity of  $2.5 \times 10^{-4} \text{ M}$  methanoic acid is  $5.25 \times 10^{-5} \text{ S cm}^{-1}$ . Calculate its molar conductivity and degree of dissociation.

Given:  $\lambda^o(H^+) = 349.5 \text{ S cm}^2 \text{ mol}^{-1}$  and  $\lambda^o(HCOO^-) = 50.5 \text{ S cm}^2 \text{ mol}^{-1}$ .

[CBSE Allahabad 2015]

**Ans.**

$$= \Lambda_m^c \frac{k \times 1000}{M}$$

$$\Lambda_m^c = \frac{5.25 \times 10^{-5} \text{ S cm}^{-1} \times 1000 \text{ cm}^3 \text{ L}^{-1}}{2.5 \times 10^{-4} \text{ mol L}^{-1}}$$

$$\Lambda_m^c = 210 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\begin{aligned}\Lambda_m^c(\text{HCOOH}) &= \Lambda_{\text{HCOO}^-}^o + \Lambda_{\text{H}^+}^o \\ &= 50.5 \text{ S cm}^2 \text{ mol}^{-1} + 349.5 \text{ S cm}^2 \text{ mol}^{-1} \\ &= 400 \text{ S cm}^2 \text{ mol}^{-1}\end{aligned}$$

$$\text{Degree of dissociation, } \alpha = \frac{\Lambda_m^c}{\Lambda_m^o} = \frac{210 \text{ S cm}^2 \text{ mol}^{-1}}{400 \text{ S cm}^2 \text{ mol}^{-1}} = 0.525$$

$$\text{or } \alpha = 52.5\%$$

**Q.4. Answer the following questions**

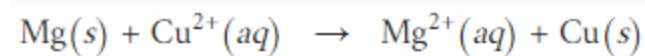
[CBSE (AI) 2014]

Calculate  $\Delta_r G^\circ$  for the reaction



$$\text{Given: } E_{\text{cell}}^\circ = +2.71 \text{ V, } 1 \text{ F} = 95600 \text{ C mol}^{-1}$$

**Ans.**



$$\Delta_r G^\circ = -nFE_{\text{cell}}^\circ$$

$$\text{Here, } n = 2, F = 96500 \text{ C mol}^{-1} \text{ and } E_{\text{cell}}^\circ = 2.71 \text{ V}$$

$$\begin{aligned}\therefore \Delta_r G^\circ &= -2 \times 96500 \text{ C mol}^{-1} \times 2.71 \text{ V} \\ &= -523030 \text{ J mol}^{-1} = -523.03 \text{ kJ mol}^{-1}\end{aligned}$$

**Q. Name the type of cell which was used in Apollo space programme for providing electrical power.**

**Ans.** H<sub>2</sub>—O<sub>2</sub> fuel cell

**Q.5. Conductivity of 0.00241 M acetic acid solution is  $7.896 \times 10^{-5} \text{ S cm}^{-1}$ . Calculate its molar conductivity in this solution. If  $\Lambda_m^o$  for acetic acid is  $390.5 \text{ S cm}^2 \text{ mol}^{-1}$ , what would be its dissociation constant?**

**[CBSE Delhi 2008]**

**Ans.**

$$c = 0.00241 \text{ M}, k = 7.896 \times 10^{-5} \text{ S cm}^{-1}, \Lambda_m^o = 390.5 \text{ S cm}^2 \text{ mol}^{-1}$$

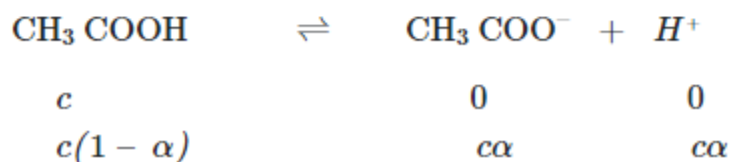
$$\Lambda_m = \frac{k \times 1000}{c}$$

Substituting the values

$$\Lambda_m = \frac{7.896 \times 10^{-5} \times 1000}{0.002} = 32.76 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\alpha = \frac{\Lambda_m^c}{\Lambda_m^o} = \frac{32.76}{390.5} = 0.084$$

$$\alpha = 8.4\%$$



$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]} = \frac{c\alpha \cdot c\alpha}{c(1 - \alpha)} = \frac{c\alpha^2}{1 - \alpha}$$

$$K_a = \frac{0.0241 (0.084)^2}{(1 - 0.084)} = 1.86 \times 10^{-5}$$

**Q.6.**

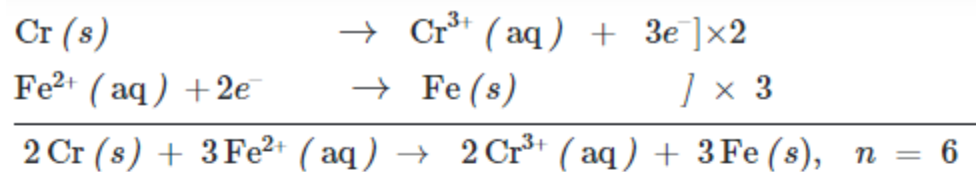
Calculate  $\Delta_r G^\circ$  and  $\log K_c$  for the following reaction at 298 K:



$$\text{Given: } E^\circ_{\text{cell}} = 0.30 \text{ V}$$

**[CBSE (F) 2016]**

Ans.



Here,  $n = 6$ ,  $E^\circ_{\text{cell}} = 0.30 \text{ V}$

Substituting the values in the expression,  $\log K_c = \frac{n}{0.059} E^\circ_{\text{cell}}$ , we get

$$\log K_c = \frac{6}{0.059} \times 0.30 \quad \text{or} \quad \log K_c = 30.5084$$

$$K_c = \text{Antilog (30.5084)} \quad \text{or} \quad K_c = 3.224 \times 10^{30}$$

$$\Delta G^\circ = - nFE^\circ_{\text{cell}}$$

$$= - 6 \times 96500 \times 0.30 = - 173700 \text{ J mol}^{-1}$$

$$\Delta G^\circ = - 173.7 \text{ kJ mol}^{-1}$$

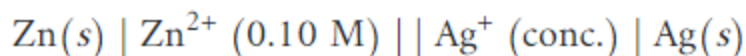
**Q.7.** One half-cell in a voltaic cell is constructed from a silver wire dipped in silver nitrate solution of unknown concentration. The other half-cell consists of a zinc electrode in a 0.10 M solution of  $\text{Zn}(\text{NO}_3)_2$ . A voltage of 1.48 V is measured for this cell. Use this information to calculate the concentration of silver ions in the solution.

$$(\text{Given: } E^\circ_{\text{Zn}^{2+}/\text{Zn}} = - 0.763 \text{ V}, E^\circ_{\text{Ag}^+/\text{Ag}} = + 0.80 \text{ V})$$

[CBSE (F) 2010]

Ans.

Electrochemical cell



$$\begin{aligned} E^\circ_{\text{cell}} &= E^\circ_R - E^\circ_L = E^\circ_{\text{Ag}^+/\text{Ag}} - E^\circ_{\text{Zn}^{2+}/\text{Zn}} \\ &= 0.80 \text{ V} - (-0.763) \text{ V} = 1.563 \text{ V} \end{aligned}$$

We know that,

$$\begin{aligned} E_{\text{cell}} &= E^\circ_{\text{cell}} - \frac{0.0591}{n} \log \frac{[\text{Zn}^{2+}]}{[\text{Ag}^+]^2} \\ 1.48 &= 1.563 - \frac{0.0591}{2} \log \frac{[0.10]}{[\text{Ag}^+]^2} \\ \log &= \frac{[0.10]}{[\text{Ag}^+]^2} = \frac{0.083}{0.02955} = 2.8087 \end{aligned}$$

Or  $\frac{[0.10]}{[\text{Ag}^+]^2} = \text{antilog } 2.8087 = 643.7$

$$[\text{Ag}^+]^2 = \frac{0.10}{643.7} = 1.553 \times 10^{-4}$$

$$[\text{Ag}^+] = 1.247 \times 10^{-2} \text{ M}$$

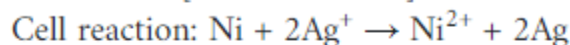
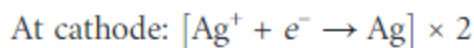
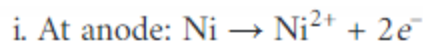
**Q.8.** A strip of nickel metal is placed in a 1-molar solution of  $\text{Ni}(\text{NO}_3)_2$  and a strip of silver metal is placed in a 1-molar solution of  $\text{AgNO}_3$ . An electrochemical cell is created when the two solutions are connected by a salt bridge and the two strips are connected by wires to a voltmeter.

- Write the balanced equation for the overall reaction occurring in the cell and calculate the cell potential.
- Calculate the cell potential,  $E$ , at  $25^\circ\text{C}$  for the cell if the initial concentration of  $\text{Ni}(\text{NO}_3)_2$  is 0.100 molar and the initial concentration of  $\text{AgNO}_3$  is 1.00 molar.

$$\left[ E_{\text{Ni}^{2+}/\text{Ni}}^{\circ} = -0.25 \text{ V}; E_{\text{Ag}^{+}/\text{Ag}}^{\circ} = 0.80 \text{ V}, \log 10^{-1} = -1 \right]$$

[CBSE (F) 2012]

Ans.



$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$$

$$= E_{\text{Ag}^{+}/\text{Ag}}^{\circ} - E_{\text{Ni}^{2+}/\text{Ni}}^{\circ} = 0.80 \text{ V} - (-0.25 \text{ V})$$

$$E_{\text{cell}}^{\circ} = 1.05 \text{ V}$$

ii.  $E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.059}{n} \log \frac{[\text{Ni}^{2+}]}{[\text{Ag}^{+}]^2}$

Here,  $n = 2$ ,  $E_{\text{cell}}^{\circ} = 1.05 \text{ V}$ ,  $[\text{Ni}^{2+}] = 0.1 \text{ M}$ ,  $[\text{Ag}^{+}] = 1.0 \text{ M}$

$$E_{\text{cell}} = 1.05 \text{ V} - \frac{0.059}{2} \log \frac{(0.1)}{(1)^2}$$

$$E_{\text{cell}} = 1.05 \text{ V} - 0.0295 \log 10^{-1}$$

$$= 1.05 + 0.0295 \text{ V} = 1.0795 \text{ V}$$

**Q.9. The electrical resistance of a column of 0.05 M NaOH solution of diameter 1 cm and length 50 cm is  $5.55 \times 10^3 \text{ ohm}$ . Calculate its resistivity, conductivity and molar conductivity.**

[CBSE (AI) 2012]

Ans.

$$A = \pi r^2 = 3.14 \times \left(\frac{1}{2} \text{ cm}\right)^2 = 0.785 \text{ cm}^2; l = 50 \text{ cm}$$

$$\text{Resistivity, } \rho = \frac{R \times A}{l} = \frac{5.55 \times 10^3 \text{ ohm} \times 0.785 \text{ cm}^2}{50 \text{ cm}} = 87.135 \text{ ohm cm}$$

$$\text{Conductivity, } k = \frac{1}{\rho} = \frac{1}{87.135 \text{ ohm cm}} = 0.01148 \text{ S cm}^{-1}$$

$$\text{Molar Conductivity, } \Lambda_m = \frac{k \times 1000}{M} = \frac{0.01148 \text{ S cm}^{-1} \times 1000 \text{ cm}^3 \text{ L}^{-1}}{0.05 \text{ mol L}^{-1}}$$

$$= 229.6 \text{ S cm}^2 \text{ mol}^{-1}$$

**Q.10. A voltaic cell is set up at 25°C with the following half cells:**

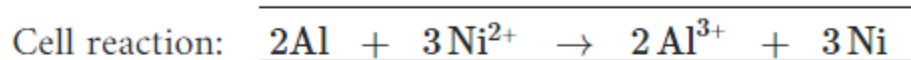
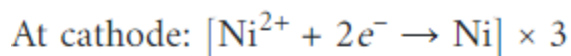
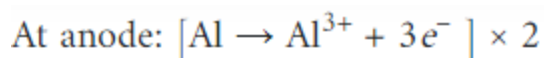
**Al/Al<sup>3+</sup> (0.001 M) and Ni/Ni<sup>2+</sup> (0.50 M)**

**Write an equation for the reaction that occurs when the cell generates an electric current and determine the cell potential.**

$$E_{\text{Ni}^{2+}/\text{Ni}}^{\circ} = -0.25 \text{ V}; E_{\text{Al}^{3+}/\text{Al}}^{\circ} = -1.66 \text{ V} (\log 8 \times 10^{-6} = -0.54)$$

**[CBSE (AI) 2012]**

**Ans.**



$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = E_{\text{Ni}^{2+}/\text{Ni}}^{\circ} - E_{\text{Al}^{3+}/\text{Al}}^{\circ}$$

$$= -0.25 \text{ V} - (-1.66 \text{ V}) = 1.41 \text{ V}$$

$$[\text{Al}^{3+}] = 1 \times 10^{-3} \text{ M}; [\text{Ni}^{2+}] = 0.5 \text{ M}; n = 6$$

Substituting the values in the Nernst equation,

$$E_{\text{cell}} = E_{\text{Ni}^{2+}/\text{Ni}}^{\circ} - E_{\text{Al}^{3+}/\text{Al}}^{\circ}$$

$$E_{\text{cell}} = 1.41 \text{ V} - \frac{0.059}{6} \log \frac{(10^{-3})^2}{(0.5)^3}$$

$$= 1.41 \text{ V} - \frac{0.059}{6} \log (8 \times 10^{-6}) = 1.41 \text{ V} - \frac{0.059}{6} (-0.54)$$

$$E_{\text{cell}} = 1.41 \text{ V} + 0.0053 \text{ V} = \mathbf{1.4153 \text{ V}}$$

## Short Answer Questions-II (OIQ)

**Q.1.** Chromium metal can be plated out from an acidic solution containing  $\text{CrO}_3$  according to the following equation:



Calculate (i) how many grams of chromium will be plated out by 24,000 coulombs and (ii) how long will it take to plate out 1.5 g of chromium by using 12.5 A current?

(At. mass of Cr = 52).

**Ans.**

- i.  $6 \times 96500$  coulomb deposit Cr = 1 mole = 52 g  
 $\therefore 24,000$  coulomb deposit Cr =  $\frac{52 \times 24000}{6 \times 96500}$  g = **2.1554 g**
- ii. 52 g of Cr is deposited by electricity =  $6 \times 96500$  C  
 $\therefore 1.5$  g require electricity =  $\times 1.5$  C = 16071.9 C  
 $\therefore$  Time for which the current is required to be passed =  $\frac{16071.9 \text{ C}}{12.5 \text{ A}} = \mathbf{1336 \text{ s.}}$

**Q.2.** Answer the following questions

**Q.** A current of 1.50 A was passed through an electrolytic cell containing  $\text{AgNO}_3$  solution with inert electrodes. The weight of Ag deposited was 1.50 g. How long did the current flow?

**Ans.**



Quantity of charge required to deposit 108 g of silver = 96500 C

$\therefore$  Quantity of charge required to deposit 1.50 g of silver =  $\frac{96500}{108} \times 1.50 = 1340.28 \text{ C}$

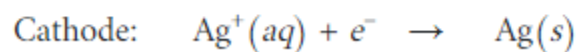
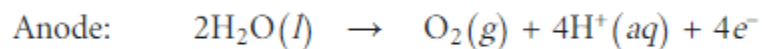
$$t = \frac{Q}{I}$$

$\therefore$  Time taken =  $\frac{1340.28}{1.50} = \mathbf{893.52 \text{ s}}$

**Q. Write the reactions taking place at the anode and cathode in the above cell.**

**Ans.**

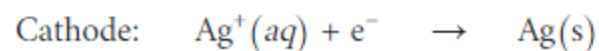
Inert electrodes



**Q. Give reactions taking place at the two electrodes if these are made up of Ag.**

**Ans.**

Ag electrodes

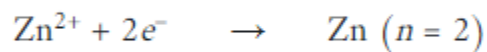


**Q.3. A zinc rod is dipped in 0.1 M solution of  $\text{ZnSO}_4$ . The salt is 95% dissociated at this dilution at 298 K. Calculate the electrode potential**

$$(E_{\text{Zn}^{2+}/\text{Zn}}^{\circ} = -0.76 \text{ V}).$$

**Ans.**

The electrode reaction written as reduction reaction is



Applying Nernst equation, we get

$$E_{\text{Zn}^{2+}/\text{Zn}} = E_{\text{Zn}^{2+}/\text{Zn}}^{\circ} - \frac{0.0591}{2} \log \frac{1}{[\text{Zn}^{2+}]}$$

As 0.1 M  $\text{ZnSO}_4$  solution is 95% dissociated, this means that in the solution,

$$[\text{Zn}^{2+}] = \frac{95}{100} \times 0.1 \text{ M} = 0.095 \text{ M}$$

$$E_{\text{Zn}^{2+}/\text{Zn}} = -0.76 - \frac{0.0591}{2} \log \frac{1}{0.095}$$

$$= -0.76 - 0.02955 (\log 1000 - \log 95)$$

$$= -0.76 - 0.02955 (3 - 1.9777) = -0.76 - 0.03021 = -0.79021 \text{ V}$$

**Q.4. The conductivity of  $0.001028 \text{ mol L}^{-1}$  acetic acid is  $4.95 \times 10^{-5} \text{ S cm}^{-1}$ .**

**Calculate its dissociation constant  $\Lambda_m^{\circ}$  if for acetic acid is  $390.5 \text{ S cm}^2 \text{ mol}^{-1}$ .**

**[HOTS]**

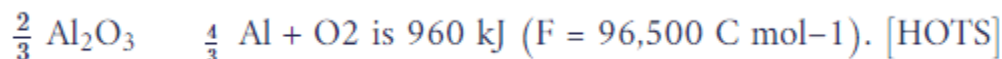
**Ans.**

$$\Lambda_m = \frac{k \times 1000}{c} = \frac{4.95 \times 10^{-5} \text{ S cm}^{-1}}{0.001028 \text{ mol L}^{-1}} \times \frac{1000 \text{ cm}^3}{\text{L}} = 48.15 \text{ S cm}^2 \text{ mol}^{-1}$$

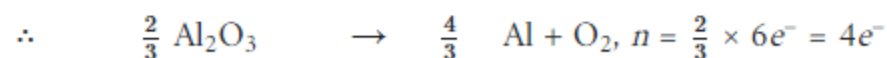
$$\alpha = \frac{\Lambda_m}{\Lambda_m^{\circ}} = \frac{48.15 \text{ S cm}^2 \text{ mol}^{-1}}{390.5 \text{ S cm}^2 \text{ mol}^{-1}} = 0.1233$$

$$K = \frac{c\alpha^2}{(1-\alpha)} = \frac{0.001028 \text{ mol L}^{-1} \times (0.1233)^2}{1 - 0.1233} = 1.78 \times 10^{-5} \text{ mol L}^{-1}$$

**Q.5. Estimate the minimum potential difference needed to reduce  $\text{Al}_2\text{O}_3$  at  $500^\circ\text{C}$ . The free energy change for the decomposition reaction**



**Ans.**



$$\Delta_r G = 960 \times 1000 = 960000 \text{ J}$$

$$\text{Now, } \Delta_r G = - nFE_{\text{cell}}$$

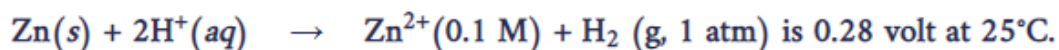
$$\Rightarrow E_{\text{cell}} = - \frac{\Delta_r G}{nF} = \frac{-960000}{4 \times 96500}$$

$$\Rightarrow E_{\text{cell}} = - 2.487 \text{ V}$$

$\therefore$  Minimum potential difference needed to reduce  $\text{Al}_2\text{O}_3$  is  $- 2.487 \text{ V}$ .

**Q.6.**

The emf of a cell corresponding to the reaction.

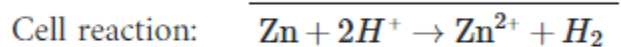
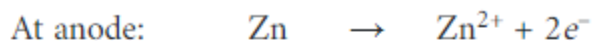


Write the half-cell reactions and calculate the pH of the solution at the hydrogen electrode.

$$E_{\text{Zn}^{2+}/\text{Zn}}^o = - 0.76 \text{ V}, E_{\text{H}^+/\text{H}_2}^o = 0 \text{ V} \quad [\text{HOTS}]$$

**Ans.**

Half-cell reactions:



$$E_{\text{cell}} = E_{\text{cell}}^o - \frac{0.0591}{n} \log \frac{[\text{Zn}^{2+}]}{[\text{H}^+]^2}$$

$$= \left( E_{\text{H}^+/\text{H}_2}^o - E_{\text{Zn}^{2+}/\text{Zn}}^o \right) - \frac{0.0591}{2} \log \frac{0.1}{[\text{H}^+]^2}$$

$$= [0 - (- 0.76)] - 0.02955 [\log 10^{-1} - 2 \log (\text{H}^+)]$$

$$0.28 = 0.76 - 0.02955 (-1 + 2 \text{ pH}) \quad [\because \text{pH} = -\log (\text{H}^+)]$$

$$2 \text{ pH} - 1 = 16.244$$

$$\text{pH} = 8.62$$

**Q.7.**

From the following molar conductivities at infinite dilution, calculate  $\Lambda_m^0$  for  $\text{NH}_4\text{OH}$ .

$$\Lambda_m^0 \text{ for } \text{Ba}(\text{OH})_2 = 457.6 \, \Omega^{-1} \text{ cm}^2 \text{ mol}^{-1}$$

$$\Lambda_m^0 \text{ for } \text{BaCl}_2 = 240.6 \, \Omega^{-1} \text{ cm}^2 \text{ mol}^{-1}$$

$$\Lambda_m^0 \text{ for } \text{NH}_4\text{Cl} = 129.8 \, \Omega^{-1} \text{ cm}^2 \text{ mol}^{-1}$$

**Ans.**

$$\begin{aligned} \Lambda_{m(\text{NH}_4\text{OH})}^o &= \lambda_{\text{NH}_4^+}^o + \lambda_{\text{OH}^-}^o \\ &= (\lambda_{\text{NH}_4^+}^o + \lambda_{\text{Cl}^-}^o) + \frac{1}{2}(\lambda_{\text{Ba}^{2+}}^o + 2\lambda_{\text{OH}^-}^o) - \frac{1}{2}(\lambda_{\text{Ba}^{2+}}^o + 2\lambda_{\text{Cl}^-}^o) \\ &= \Lambda_{m(\text{NH}_4\text{Cl})}^o + \frac{1}{2}[\Lambda_{m(\text{Ba}(\text{OH})_2)}^o] - \frac{1}{2}[\Lambda_{m(\text{BaCl}_2)}^o] \\ &= 129.8 + \frac{1}{2} \times 457.6 - \frac{1}{2} \times 240.6 = 238.3 \, \text{ohm}^{-1} \text{ cm}^2 \text{ mol}^{-1} \end{aligned}$$