

2 Marks Questions

3. Define electrochemical cell. What happens when applied external potential becomes greater than E_{cell}° of electrochemical cell?

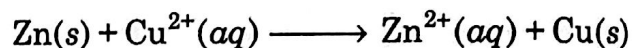
4. Using the E° values of X and Y predict which is better for coating the surface of iron to prevent rust and why?

Given, $E_{\text{Fe}^{2+}/\text{Fe}}^{\circ} = -0.44 \text{ V}$

$$E_{\text{X}^{2+}/\text{X}}^{\circ} = -2.36 \text{ V}$$

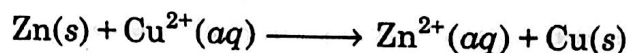
$$E_{\text{Y}^{2+}/\text{Y}}^{\circ} = -0.14 \text{ V}$$

5. The standard electrode potential (E°) for Daniell cell is + 1.1 V. Calculate the ΔG° for the reaction



(1F = 96500 C mol⁻¹).

Or The standard electrode potential for Daniell cell is 1.1 V. Calculate the standard Gibbs energy for the cell reaction. ($F = 96500 \text{ C mol}^{-1}$)



6. A zinc rod is dipped in 0.1 M solution of ZnSO₄. The salt is 95% dissociated at this dilution at 298 K. Calculate the electrode potential. [$E_{\text{Zn}^{2+}/\text{Zn}}^{\circ} = -0.76 \text{ V}$]

7. Given that, the standard electrode potentials (E°) of metals are

$$\text{K}^+ / \text{K} = -2.93 \text{ V}, \quad \text{Ag}^+ / \text{Ag} = 0.80 \text{ V}$$

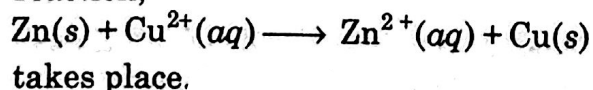
$$\text{Cu}^{2+} / \text{Cu} = 0.34 \text{ V}, \quad \text{Mg}^{2+} / \text{Mg} = -2.37 \text{ V}$$

$$\text{Cr}^{3+} / \text{Cr} = -0.74 \text{ V}, \quad \text{Fe}^{2+} / \text{Fe} = -0.44 \text{ V}$$

Arrange these metals in an increasing order of their reducing power.

1 Mark Questions

1. Represent the galvanic cell in which the reaction,



2. What is the necessity to use a salt bridge in a Galvanic cell?

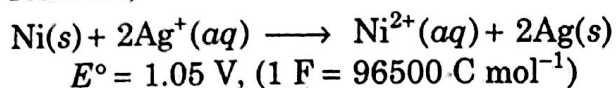
Test 1

Test 1

3 Marks Questions

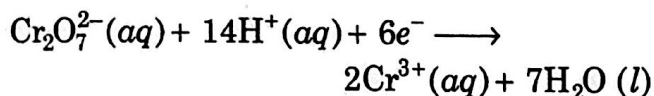
8. Calculate $\Delta_r G^\circ$ and $\log K_C$ for the following reaction :
- $$\text{Cd}^{2+}(\text{aq}) + \text{Zn}(\text{s}) \longrightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cd}(\text{s})$$
- Given : $E_{\text{Cd}^{2+}/\text{Cd}}^\circ = -0.403 \text{ V}$
 $E_{\text{Zn}^{2+}/\text{Zn}}^\circ = -0.763 \text{ V}$
9. Consider the following reaction :
- $$\text{Cu}(\text{s}) + 2\text{Ag}^+(\text{aq}) \longrightarrow 2 \text{Ag}(\text{s}) + \text{Cu}^{2+}(\text{aq})$$
- (i) Depict the galvanic cell in which the given reaction takes place.
(ii) Give the direction of flow of current.
(iii) Write the half-cell reactions taking place at cathode and anode.
10. Calculate the emf of the following cell at 298 K.
- $$2\text{Cr}(\text{s}) + 3\text{Fe}^{2+}(0.1\text{M}) \longrightarrow 2\text{Cr}^{3+}(0.01\text{M}) + 3\text{Fe}(\text{s})$$
- Given,
 $E_{(\text{Cr}^{3+}/\text{Cr})}^\circ = -0.74 \text{ V}$, $E_{(\text{Fe}^{2+}/\text{Fe})}^\circ = -0.44 \text{ V}$
11. Calculate the emf of the following cell at 25°C.
- $$\text{Fe} | \text{Fe}^{2+}(0.001\text{M}) || \text{H}^+(0.01 \text{ M}) | \text{H}_2(\text{g}) (1\text{bar}) | \text{Pt} (\text{s})$$
- $$E_{(\text{Fe}^{2+}/\text{Fe})}^\circ = -0.44 \text{ V}; E_{(\text{H}^+/\text{H}_2)}^\circ = 0.00 \text{ V}$$
12. Calculate the emf of the following cell at 25°C.
- $$\text{Zn} | \text{Zn}^{2+} (0.001 \text{ M}) || \text{H}^+ (0.01 \text{ M}) | \text{H}_2(\text{g}) (1 \text{ bar}) | \text{Pt}(\text{s})$$
- $$E_{(\text{Zn}^{2+}/\text{Zn})}^\circ = -0.76 \text{ V}, E_{\text{H}^+/\text{H}_2}^\circ = 0.00 \text{ V}$$
13. Calculate the emf of the following cell at 298 K.
- $$\text{Fe}(\text{s}) | \text{Fe}^{2+}(0.001 \text{ M}) || \text{H}^+(1 \text{ M}) | \text{H}_2(\text{g}) (1 \text{ bar}) | \text{Pt} (\text{s})$$
- [Given, $E_{\text{cell}}^\circ = +0.44 \text{ V}$]
14. Calculate the emf of the following cell at 25°C.
- $$\text{Ag} (\text{s}) | \text{Ag}^+(10^{-3} \text{ M}) || \text{Cu}^{2+}(10^{-1} \text{ M}) | \text{Cu} (\text{s})$$
- [Given, $E_{\text{cell}}^\circ = +0.46 \text{ V}$ and $\log 10^n = n$]
15. 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 one molar solution of 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.
- (i) Write the balanced equations for the overall reaction occurring in the cell and calculate the cell potential.
(ii) Calculate the cell potential, E , at 25°C for the cell, if the initial concentration of $\text{Ni}(\text{NO}_3)_2$ is 0.100 molar and the initial concentration of AgNO_3 is 1.00 molar.
- $$[E_{\text{Ni}^{2+}/\text{Ni}}^\circ = -0.25 \text{ V}, E_{\text{Ag}^+/\text{Ag}}^\circ = 0.80 \text{ V}, \log 10^{-1} = -1]$$
16. In the button cell, widely used in watches, the following reaction takes place,
- $$\text{Zn} (\text{s}) + \text{Ag}_2\text{O}(\text{s}) + \text{H}_2\text{O}(\text{l}) \longrightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Ag} (\text{s}) + 2\text{OH}^-(\text{aq})$$
- Determine E° and ΔG° for the reaction.
[Given, $E_{\text{Ag}^+/\text{Ag}}^\circ = +0.80 \text{ V}$,
 $E_{\text{Zn}^{2+}/\text{Zn}}^\circ = -0.76 \text{ V}$]
17. A voltaic cell is set up at 25°C with the following half-cells.
- $$\text{Al} | \text{Al}^{3+}(0.001\text{M}) \text{ and } \text{Ni} | \text{Ni}^{2+} (0.50 \text{ M})$$
- Write an equation for the reaction that occurs when the cell generates an electric current and determine the cell potential.
- $$E_{\text{Ni}^{2+}/\text{Ni}}^\circ = -0.25 \text{ V}$$
- and $E_{\text{Al}^{3+}/\text{Al}}^\circ = -1.66 \text{ V}$
 $(\log 8 \times 10^{-6} = -5.0969)$

18. Determine the values of equilibrium constant (K_C) and ΔG° for the following reaction,



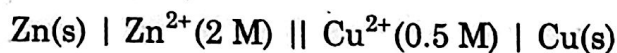
19. Calculate the potential for half-cell containing 0.10 M $\text{K}_2\text{Cr}_2\text{O}_7$ (aq), 0.20 M Cr^{3+} (aq) and 1.0×10^{-4} M H^+ (aq).

The half-cell reaction is



and the standard electrode potential is given as $E^\circ = 1.33 \text{ V}$.

20. For the cell,



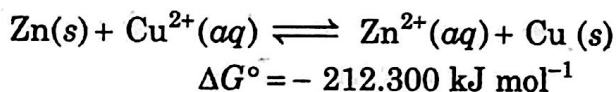
(i) Write the equation for each half reaction.

(ii) Calculate the cell potential at 25°C.

$$[\text{Given}, E^\circ_{\text{Zn}^{2+}/\text{Zn}} = -0.76 \text{ V},$$

$$E^\circ_{\text{Cu}^{2+}/\text{Cu}} = +0.34 \text{ V}]$$

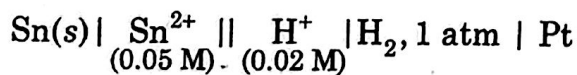
21. Calculate the equilibrium constant, K , for the reaction at 298 K.



$$[\text{Given}, E^\circ_{\text{Zn}^{2+}/\text{Zn}} = -0.76 \text{ V},$$

$$E^\circ_{\text{Cu}^{2+}/\text{Cu}} = +0.34 \text{ V}]$$

22. Write the Nernst equation and compute the emf of the following cell at 298 K.



$$[E^\circ_{\text{Sn}^{2+}/\text{Sn}} = -0.144 \text{ V}]$$

23. A copper-silver cell is set up. The copper ion concentration is 0.10 M. The concentration of silver ion is not known. The cell potential when measured was 0.422 V. Determine the concentration of silver ions in the cell.

$$[\text{Given}, E^\circ_{\text{Ag}^+/\text{Ag}} = +0.80 \text{ V},$$

$$E^\circ_{\text{Cu}^{2+}/\text{Cu}} = +0.34 \text{ V}]$$

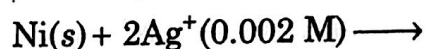
24. One half-cell in a voltaic cell is constructed from a silver wire dipped in silver nitrate solution of unknown concentration. The other half-cell consists of a zinc electrode in 1.0M solution of $\text{Zn}(\text{NO}_3)_2$. A voltage of 1.48 V is measured for this cell.

Use this information to calculate the concentration of silver nitrate solution.

$$[\text{Given}, E^\circ_{\text{Zn}^{2+}/\text{Zn}} = -0.763 \text{ V},$$

$$E^\circ_{\text{Ag}^+/\text{Ag}} = +0.80 \text{ V}]$$

25. Calculate the emf of the cell in which the following reaction takes place,



[Given that, $E^\circ_{\text{cell}} = 1.05 \text{ V}$].

26. A voltaic cell is set-up at 25°C with the following half-cells.

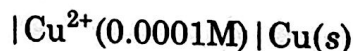
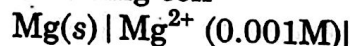


What would be the voltage of this cell?

[Given, $E^\circ_{\text{cell}} = 0.46 \text{ V}$]

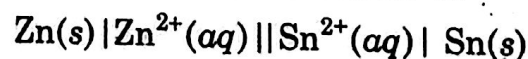
5 Marks Questions

27. Calculate the emf and ΔG for the following cell



$$E^\circ_{(\text{Mg}^{2+}/\text{Mg})} = -2.37\text{V}, E^\circ_{(\text{Cu}^{2+}/\text{Cu})} = 0.34\text{V}$$

28. Calculate the $\Delta_r G^\circ$ and emf (E) that can be obtained from the following cell under the standard conditions at 25°C.



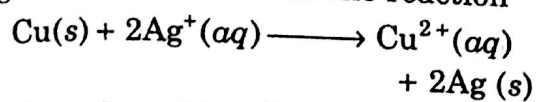
$$[\text{Given}, E^\circ_{\text{Zn}^{2+}/\text{Zn}} = -0.76 \text{ V};$$

$$E^\circ_{\text{Sn}^{2+}/\text{Sn}} = -0.14 \text{ V}$$

and

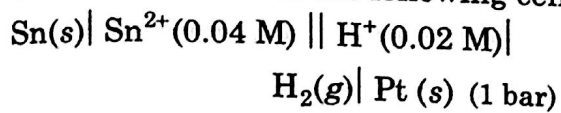
$$\text{F} = 96500 \text{ C mol}^{-1}]$$

29. (i) Write the formulation for the galvanic cell in which the reaction



takes place. Identify the cathode and the anode reactions in it.

- (ii) Write the Nernst equation and calculate the emf of the following cell.



[Given $E^\circ_{\text{Sn}^{2+}/\text{Sn}} = -0.14 \text{ V}$]

8. The molar conductivity of a 1.5 M solution of an electrolyte is found to be $138.9 \text{ S cm}^2 \text{ mol}^{-1}$. Calculate the conductivity of this solution.

9. State and explain Kohlrausch's law of independent migration of ions. Write an expression for the molar conductivity of acetic acid at infinite dilution according to Kohlrausch's law.

10. Define conductivity and molar conductivity for the solution of an electrolyte. Discuss their variation with concentration.

1 Mark Questions

1. Express the relation among the conductivity of a solution in the cell, the cell constant and the resistance of solution in the cell.
2. What is meant by limiting molar conductivity?

2 Marks Questions

3. Explain with a graph, the variation of molar conductivity of a strong electrolyte with dilution.
4. Calculate the degree of dissociation (α) of acetic acid if its molar conductivity (Λ_m) is $39.05 \text{ S cm}^2 \text{ mol}^{-1}$.
Given: $\lambda^\circ(\text{H}^+) = 349.6 \text{ S cm}^2 \text{ mol}^{-1}$
and $\lambda^\circ(\text{CH}_3\text{COO}^-) = 40.9 \text{ S cm}^2 \text{ mol}^{-1}$

5. State Kohlrausch's law of independent migration of ions. Why does the conductivity of a solution decrease with dilution?

6. The conductivity of 0.001 M acetic acid is $4 \times 10^{-5} \text{ S/cm}$. Calculate the dissociation constant of acetic acid, if molar conductivity at infinite dilution for acetic acid is $390 \text{ S cm}^2 / \text{mol}$.

7. Express the relation among cell constant, resistance of the solution in the cell and conductivity of the solution. How is molar conductivity of a solution related to its conductivity?

3 Marks Questions

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

12. The resistance of 0.01 M NaCl solution at 25°C is 200Ω . The cell constant of the conductivity cell used is unity. Calculate the molar conductivity of the solution.

13. The value of Λ_m° of $\text{Al}_2(\text{SO}_4)_3$ is $858 \text{ S cm}^2 \text{ mol}^{-1}$, while $\lambda_m^\circ(\text{SO}_4^{2-})$ is $160 \text{ S cm}^2 \text{ mol}^{-1}$ calculate the limiting ionic conductivity of Al^{3+} .

14. When a certain conductance cell was filled with 0.1 M KCl, it has a resistance of 85 ohms at 25°C . When the same cell was filled with an aqueous solution of 0.052 M unknown electrolyte, the resistance was 96 ohms. Calculate the molar conductance of the electrolyte at this concentration.

[Specific conductance of 0.1 M KCl = $1.29 \times 10^{-2} \text{ ohm}^{-1} \text{ cm}^{-1}$]

16. The electrical resistance of a column of 0.05M NaOH solution of diameter 1 cm and length 50 cm is $5.55 \times 10^3 \Omega$. Calculate its resistivity, conductivity and molar conductivity.

16. Calculate the degree of dissociation of acetic acid at 298 K, given that

$$\Lambda_m(\text{CH}_3\text{COOH}) = 11.7 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\Lambda_m^\circ(\text{CH}_3\text{COO}^-) = 40.9 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\Lambda_m^\circ(\text{H}^+) = 349.1 \text{ S cm}^2 \text{ mol}^{-1}$$

17. Calculate Λ_m° for acetic acid,

Given that,

$$\Lambda_m^\circ(\text{HCl}) = 426 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\Lambda_m^\circ(\text{NaCl}) = 126 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\Lambda_m^\circ(\text{CH}_3\text{COONa}) = 91 \text{ S cm}^2 \text{ mol}^{-1}$$

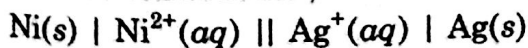
5 Marks Questions

18. (i) The conductivity of 0.001 mol L⁻¹ solution of CH₃COOH is $3.905 \times 10^{-5} \text{ S cm}^{-1}$. Calculate its molar conductivity and degree of dissociation (α). Given, $\lambda^\circ(\text{H}^+) = 349.6 \text{ S cm}^2 \text{ mol}^{-1}$ and $\lambda^\circ(\text{CH}_3\text{COO}^-) = 40.9 \text{ S cm}^2 \text{ mol}^{-1}$.

(ii) Define electrochemical cell. What happens if external potential applied becomes greater than E_{cell}° of electrochemical cell?

19. (i) Define the term degree of dissociation. Write an expression that relates the molar conductivity of a weak electrolyte to its degree of dissociation.

(ii) For the cell reaction,



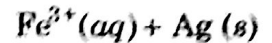
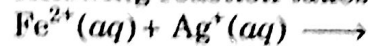
Calculate the equilibrium constant at 25°C.

How much maximum work would be obtained by operation of this cell?

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

20. (i) Define conductivity and molar conductivity for the solution of an electrolyte. Discuss their variation with concentration.

(ii) Calculate the standard cell potential of the galvanic cell in which the following reaction takes place:

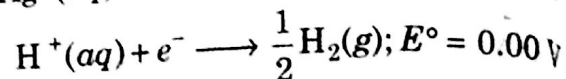
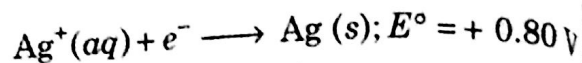


Calculate the $\Delta_r G^\circ$ and equilibrium constant of the reaction also.

$$E_{\text{Ag}^+/\text{Ag}}^\circ = 0.80\text{V}; E_{\text{Fe}^{3+}/\text{Fe}^{2+}}^\circ = 0.77\text{V}$$

21. (i) Define molar conductivity of a solution and explain how molar conductivity changes with change in concentration of solution for a weak and a strong electrolyte.

(ii) The resistance of a conductivity cell containing 0.001 M KCl solution at 298 K is 1500 Ω . What is the cell constant if the conductivity of 0.001 M KCl solution at 298 K is $0.146 \times 10^{-3} \text{ S cm}^{-1}$?



On the basis of their standard reduction electrode potential (E°) values, which reaction is feasible at the cathode and why?

2 Marks Questions

- A current of 1.50 A was passed through an electrolytic cell containing AgNO_3 solution with inert electrodes. The weight of silver deposited was 1.50 g. How long did the current flow? (Molar mass of Ag = 108 g mol^{-1} , $1F = 96500 \text{ C mol}^{-1}$).
- The conductivity of a 0.01 M solution of acetic acid at 298 K is $1.65 \times 10^{-4} \text{ S cm}^{-1}$. Calculate molar conductivity (Λ_m) of the solution.
- Write the name of the cell which is generally used in hearing aids. Write the reactions taking place at the anode and the cathode of this cell.
- Write the name of the cell which is generally used in transistors. Write the reactions taking place at the anode and the cathode of this cell.
- From the given cells:
Lead storage cell, Mercury cell, Fuel cell and Dry cell
Answer the following:
 - Which cell is used in hearing aids?
 - Which cell was used in Apollo Space Programme?
 - Which cell is used in automobiles and inverters?
 - Which cell does not have long life?
- Calculate the time to deposit 1.27 g of copper at cathode when a current of 2 A was passed through the solution of CuSO_4 . (Molar mass of Cu = 63.5 g mol^{-1} , $1F = 96500 \text{ C mol}^{-1}$)
 - Following reactions occur at cathode during the electrolysis of aqueous silver chloride solution :

$$\text{Ag}^+(aq) + e^- \longrightarrow \text{Ag}(s); E^\circ = + 0.80 \text{ V}$$

$$\text{H}^+(aq) + e^- \longrightarrow \frac{1}{2} \text{H}_2(g); E^\circ = 0.00 \text{ V}$$

On the basis of their standard reduction electrode potential (E°) values, which reaction is feasible at the cathode and why?
 - Define limiting molar conductivity. Why conductivity of an electrolyte solution decreases with the decrease in concentration?
- Following reactions occur at cathode during the electrolysis of aqueous copper (II) chloride solution.

$$\text{Cu}^{2+}(aq) + 2e^- \longrightarrow \text{Cu}(s), E^\circ = + 0.34 \text{ V}$$

$$\text{H}^+(aq) + e^- \longrightarrow \frac{1}{2} \text{H}_2(g), E^\circ = 0.00 \text{ V}$$

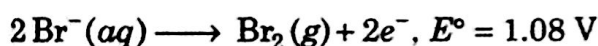
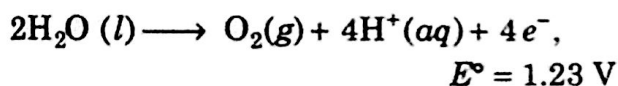
On the basis of their standard reduction electrode potential (E°) values, which reaction is feasible at the cathode and why?
 - State Kohlrausch's law of independent migration of ions. Write its one application.
- Set-up Nernst equation for the standard dry cell. Using this equation show that the voltage of a dry cell decreases with use.
- A solution of $\text{Ni}(\text{NO}_3)_2$ is electrolysed between platinum electrodes using a current of 5.0 A for 20 min. What mass of nickel will be deposited at the cathode? (Given : Atomic mass of Ni = 58.7 g mol^{-1} , $1F = 96500 \text{ C mol}^{-1}$)
- What are fuel cells? Explain the electrode reactions involved in the working of $\text{H}_2\text{—O}_2$ fuel cell.
- How much electricity in terms of Faradays is required to produce 20 g of calcium from molten CaCl_2 ?
- Silver is uniformly electrodeposited on a metallic vessel of surface area of 900 cm^2 by passing a current of 0.5 A for 2 h. Calculate the thickness of silver

deposited. Given, the density of silver is 10.5 g cm^{-3} and atomic mass of $\text{Ag} = 108 \text{ amu}$.

14. What type of a battery is the lead storage battery? Write the anode and the cathode reactions and the overall reaction occurring in a lead storage battery when current is drawn from it.
5. The chemistry of corrosion of iron is essentially an electrochemical phenomenon. Explain the reactions occurring during the corrosion of iron in the atmosphere.

16. How many moles of mercury will be produced by electrolysing $10 \text{ M Hg(NO}_3)_2$ solution with a current of 2.00 A for three hours?
 $[\text{Hg(NO}_3)_2 = 200.6 \text{ g mol}^{-1}]$

7. In the electrolysis of aqueous sodium bromide, there are two possible anodic reactions.



Which reaction occurs at anode and why?

18. Mention the reactions occurring at
 (i) anode
 (ii) cathode, during working of a mercury cell.

Why does the voltage of a mercury cell remain constant during its operation?

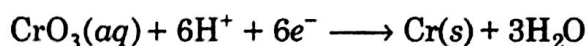
19. Give an example of a fuel cell and write the cathode and anode reactions for it.
20. Write the overall reaction that occurs during use (discharging) of

nickel-cadmium cell. Is it a primary or a secondary cell? Mention its one merit over the lead storage cell.

21. Write the anode and cathode reactions occurring in a commonly used mercury cells. How is the overall reaction represented?
22. What is a nickel-cadmium cell? State its one merit and one demerit over lead storage cell. Write the overall reaction that occurs during discharging of this cell.

3 Marks Questions

23. Chromium metal is electroplated using an acidic solution containing CrO_3 according to the following equation :



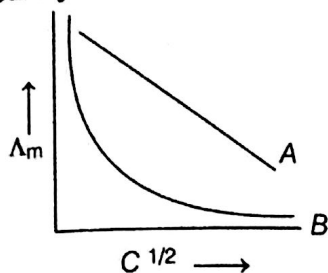
Calculate how many grams of chromium will be electroplated by $24,000$ coulombs. How long will it take to electroplate 1.5 g chromium using 12.5 A current? [Atomic mass of $\text{Cr} = 52 \text{ g mol}^{-1}$, $1 \text{ F} = 965000 \text{ C mol}^{-1}$]

24. (i) The cell in which the following reaction occurs :
- $$2\text{Fe}^{3+}(aq) + 2\text{I}^-(aq) \longrightarrow 2\text{Fe}^{2+}(aq) + \text{I}_2(s)$$
- has $E^\circ_{\text{cell}} = 0.236 \text{ V}$ at 298 K . Calculate the standard Gibbs energy of the cell reaction. (Given, $1 \text{ F} = 96500 \text{ C mol}^{-1}$)
- (ii) How many electrons flow through a metallic wire if a current of 0.5 A is passed for 2 h ? (given, $1 \text{ F} = 96500 \text{ C mol}^{-1}$)

5 Marks Questions

25. (i) Calculate the mass of Ag deposited at cathode when a current of 2A was passed through a solution of AgNO_3 for 15 min.
(Given : Molar mass of Ag = 108 g mol^{-1} , $1F = 96500 \text{ C mol}^{-1}$).
- (ii) Define fuel cell.
26. (i) Calculate ΔG° for the reaction,
 $\text{Mg}(s) + \text{Cu}^{2+}(aq) \longrightarrow \text{Mg}^{2+}(aq) + \text{Cu}(s)$
(Given,
 $E^\circ_{\text{cell}} = + 2.71 \text{ V}$, $1F = 96500 \text{ C mol}^{-1}$)
- (ii) Name the type of cell which was used in Apollo space programme for providing electrical power.
27. (i) Write two advantages of $\text{H}_2\text{—O}_2$ fuel cell over ordinary cell.
- (ii) Equilibrium constant (K_C) for the given cell reaction is 10. Calculate E°_{cell} .
$$\text{A}(s) + \text{B}^{2+}(aq) \rightleftharpoons \text{A}^{2+}(aq) + \text{B}(s)$$
28. What type of battery is lead storage battery? Write the anode and cathode reactions, and the overall cell reaction occurring in the operation of a lead storage battery.
29. An aqueous solution of copper sulphate, CuSO_4 was electrolysed between platinum electrodes using a current of 0.1287 A for 50 min.
[Given, atomic mass of Cu = 63.5 g mol^{-1}]
- (i) Write the cathodic reaction.
- (ii) Calculate
- (a) Electric charge passed during electrolysis.
- (b) Mass of copper deposited at the cathode.
[Given, $1F = 96500 \text{ C mol}^{-1}$]
30. (i) Write the cell reaction and calculate the emf of the following cell at 298 K
 $\text{Sn}(s) | \text{Sn}^{2+}(0.004 \text{ M}) || \text{H}^+(0.020 \text{ M}) | \text{H}_2(g) (1 \text{ bar}) | \text{Pt}(s)$
(Give: $E^\circ_{\text{Sn}^{2+}/\text{Sn}} = - 0.14 \text{ V}$)
- (ii) Give reasons:
- (a) On the basis of E° values, Cl_2 gas should be liberated at anode, but it is Cl_2 gas which is liberated in the electrolysis of aqueous NaCl.
- (b) Conductivity of CH_3COOH decreases on dilution.
31. (i) For the reaction
 $2\text{AgCl}(s) + \text{H}_2(g) (1 \text{ atm}) \longrightarrow 2\text{Ag}(s) + 2\text{H}^+(0.1 \text{ M}) + 2\text{Cl}^-(0.1 \text{ M})$
 $\Delta G^\circ = - 43600 \text{ J at } 25^\circ \text{ C}$
Calculate the emf of the cell.
 $[\log 10^{-n} = -n]$
- (ii) Define fuel cell and write its two advantages.
32. E°_{cell} for the given redox reaction is 2.71 V
 $\text{Mg}(s) + \text{Cu}^{2+}_{(0.001 \text{ M})} \longrightarrow \text{Mg}^{2+}_{(0.01 \text{ M})} + \text{Cu}(s)$
Calculate E_{cell} for the reaction, write the direction of flow of current when an external opposite potential applied is
- (i) less than 2.71 V
- (ii) greater than 2.71 V
33. (a) A steady current of 2 amperes was passed through two electrolytic cells X and Y connected in series containing electrolytes FeSO_4 and ZnSO_4 until 2.8 g of Fe deposited at the cathode of cell X. How long did the current flow? Calculate the mass of Zn deposited at the cathode of cell Y.
(Molar mass : Fe = 56 g mol^{-1}
Zn = 65.3 g mol^{-1} , $1F = 96500 \text{ C mol}^{-1}$)

- (b) In the plot of molar conductivity (Λ_m) vs square root of concentration ($C^{1/2}$), following curves are obtained for two electrolytes A and B.



Answer the following

- (i) Predict the nature of electrolytes A and B.
- (ii) What happens on extrapolation of Λ_m to concentration approaching zero for electrolytes A and B?
34. (i) Calculate E°_{cell} for the following reaction at 298 K:
 $2\text{Al}(s) + 3\text{Cu}^{2+}(0.01\text{ M}) \longrightarrow 2\text{Al}^{3+}(0.01\text{ M}) + 3\text{Cu}(s)$
 Given, $E_{\text{cell}} = 1.98\text{ V}$
- (ii) Using the E° values of A and B, predict which is better for coating the surface of iron [$E^\circ(\text{Fe}^{2+}/\text{Fe}) = -0.44\text{ V}$] to prevent corrosion and why?
 Given : $E^\circ(\text{A}^{2+}/\text{A}) = -2.37\text{ V}$,
 $E^\circ(\text{B}^{2+}/\text{B}) = -0.14\text{ V}$

35. (i) Define the following terms:

- (a) Molar conductivity (Λ_m)
 (b) Secondary batteries
 (c) Fuel cell

- (ii) State the following laws:

- (a) Faraday's first law of electrolysis
 (b) Kohlrausch's law of independent migration of ions

36. (i) Predict the products of electrolysis in each of the following :
- (a) An aqueous solution of AgNO_3 with platinum electrodes.
 (b) An aqueous solution of H_2SO_4 with platinum electrodes.

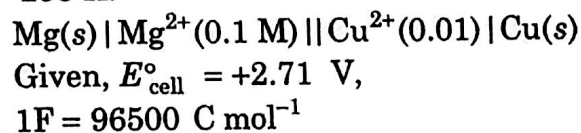
- (ii) Estimate the minimum potential difference needed to reduce Al_2O_3 at 500°C . The Gibbs energy change for the decomposition reaction
 $\frac{2}{3}\text{Al}_2\text{O}_3 \longrightarrow \frac{4}{3}\text{Al} + \text{O}_2$ is -960 kJ .
 ($F = 96500\text{ C mol}^{-1}$)

37. (i) Define the following terms :

- (a) Limiting molar conductivity
 (b) Fuel cell

- (ii) Resistance of a conductivity cell filled with 0.1 mol L^{-1} KCl solution is $100\ \Omega$. If the resistance of the same cell when filled with 0.02 mol L^{-1} KCl solution is $520\ \Omega$, calculate the conductivity and molar conductivity of 0.02 mol L^{-1} KCl solution. The conductivity of 0.1 mol L^{-1} KCl solution is $1.29 \times 10^{-2}\ \Omega^{-1}\text{ cm}^{-1}$.

38. (i) State Faraday's first law of electrolysis. How much charge in terms of Faraday's is required for the reduction of 1 mole of Cu^{2+} to Cu.
- (ii) Calculate emf of the following cell at 298 K.



Objective Questions

(For Complete Chapter)

1 Mark Questions

- Oxidation potential of unimoles of calomel is
(a) + 0.25 V (b) 0.00 V (c) + 0.287 V (d) - 0.28 V
- Given, $E_{\text{Cr}^{3+}/\text{Cr}}^{\circ} = -0.72 \text{ V}$, $E_{\text{Fe}^{2+}/\text{Fe}}^{\circ} = -0.42 \text{ V}$. The potential for the cell
 $\text{Cr} / \text{Cr}^{3+} (0.1 \text{ M}) \parallel \text{Fe}^{2+} (0.01 \text{ M}) / \text{Fe}$ is
(a) - 0.339 V (b) - 0.26 V (c) 0.26 V (d) 0.3 V
- The standard emf for the given cell reaction, $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Cu} + \text{Zn}^{2+}$ is 1.10 V at 25°C. The emf for the cell reaction, when 0.1 M Cu^{2+} and 0.1 M Zn^{2+} solutions are used at 25°C is
(a) 1.10 V (b) - 1.10 V (c) 2.20 V (d) - 2.20 V

- If the E_{cell}° for a given reaction has a negative value, then which of the following gives the correct relationships for the values of ΔG° and K_{eq} ?
(a) $\Delta G^{\circ} > 0; K_{\text{eq}} < 1$ (b) $\Delta G^{\circ} > 0; K_{\text{eq}} > 1$
(c) $\Delta G^{\circ} < 0; K_{\text{eq}} > 1$ (d) $\Delta G^{\circ} < 0; K_{\text{eq}} < 1$
- How is electrical conductance of a conductor related with length and area of cross-section of the conductor?
(a) $G = l \cdot a \cdot \kappa^{-1}$ (b) $G = \kappa \cdot l \cdot a^{-1}$
(c) $G = \kappa \cdot a \cdot l^{-1}$ (d) $G = \kappa \cdot l \cdot a^{-2}$
- Limiting molar conductivity of NH_4OH i.e. $\Lambda_{\text{m}}^{\circ}(\text{NH}_4\text{OH})$ is equal to
(a) $\Lambda_{\text{m}}^{\circ}(\text{NH}_4\text{Cl}) + \Lambda_{\text{m}}^{\circ}(\text{NaCl}) + \Lambda_{\text{m}}^{\circ}(\text{NaOH})$
(b) $\Lambda_{\text{m}}^{\circ}(\text{NaOH}) + \Lambda_{\text{m}}^{\circ}(\text{Na}_4\text{Cl}) - \Lambda_{\text{m}}^{\circ}(\text{NH}_4\text{Cl})$
(c) $\Lambda_{\text{m}}^{\circ}(\text{NH}_4\text{OH}) + \Lambda_{\text{m}}^{\circ}(\text{NH}_4\text{Cl}) - \Lambda_{\text{m}}^{\circ}(\text{HCl})$
(d) $\Lambda_{\text{m}}^{\circ}(\text{NH}_4\text{Cl}) + \Lambda_{\text{m}}^{\circ}(\text{NaOH}) - \Lambda_{\text{m}}^{\circ}(\text{NaCl})$
- An increase in equivalent conductance of a strong electrolyte with dilution is mainly due to
(a) increase in number of ions
(b) increase in ionic mobility of ions
(c) 100% ionisation of electrolyte at normal dilution
(d) increase in both i.e. number of ions and ionic mobility of ions
- The unit of conductivity is
(a) ohm^{-1} (b) $\text{ohm}^{-1} \text{cm}^{-1}$
(c) $\text{ohm}^{-2} \text{cm}^2 \text{equiv}^{-1}$ (d) $\text{ohm}^{-1} \text{cm}^2$
- The correct expression in SI system relating the equivalent conductance (Λ_c), specific conductance (κ) and equivalent concentration (C) is
[where, C is the number of gram equivalents in one litre of the solution].
(a) $\Lambda_c = \frac{\kappa}{C}$ (b) $\Lambda_c = \frac{\kappa \times 1000}{C}$
(c) $\Lambda_c = \frac{\kappa \times 10^{-3}}{C}$ (d) $\Lambda_c = \frac{\kappa \times 10^{-6}}{C}$
- The conductivity of 0.20 M KCl solution at 298 K is 0.0248 S cm^{-1} . What will be its molar conductivity?
(a) 124 S cm^2 (b) 124 cm^{-1}
(c) $124 \text{ ohm}^{-1} \text{cm}^2 \text{equiv}^{-1}$ (d) $124 \text{ S cm}^2 \text{mol}^{-1}$

11. Which of the following is not a non-electrolyte?
 (a) Acetic acid (b) Glucose (c) Ethanol (d) Urea
12. Pure water does not conduct electricity because it is
 (a) basic (b) almost unionised
 (c) decomposed easily (d) acidic
13. 1 C electricity deposits
 (a) 10.8 g of Ag (b) 96500 g of Ag
 (c) electrochemical equivalent of Ag
 (d) half of electrochemical equivalent of Ag
14. When iron is rusted, it is
 (a) reduced (b) decomposed
 (c) oxidised (d) changed in the fine powder

Explanations

1. (d) Calomel is used as a secondary reference electrode. Since, Hg is less reactive than H, its standard reduction potential is positive. Hence, its standard oxidation potential must be negative, i.e. -0.28 V. [As standard reduction potential = - standard oxidation potential]

2. (d) $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = -0.42 - (-0.72)$
 $= -0.42 + 0.72 = 0.3$ V

3. (a) $E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.059}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$
 $= 1.10 - \frac{0.059}{2} \log \frac{0.1}{0.1}, E_{\text{cell}} = 1.10$ V

4. (a) The standard free energy (ΔG°) is related to standard electrode potential (E°) and equilibrium constant as

$$\Delta G^{\circ} = -nFE^{\circ} \text{ and } \Delta G^{\circ} = -RT \log_e K_{\text{eq}}$$

Thus, if E_{cell}° is -ve, then ΔG° is +ve, thus, $\Delta G^{\circ} > 0$.

Hence, K_{eq} would be negative and less than 1, $K_{\text{eq}} < 1$.

5. (c) Specific conductance, $\kappa = G \times \frac{l}{a}$

where, G = electrical conductance or $G = \kappa \times \frac{a}{l}$

thus, electrical conductance is given by,

$$G = \kappa \cdot a \cdot l^{-1}$$

6. (d) According to Kohlrausch's law, Limiting molar conductivity of NH_4OH ,

$$\Lambda_{\text{m}}^{\circ}(\text{NH}_4\text{OH}) = \Lambda_{\text{m}}^{\circ}(\text{NH}_4\text{Cl}) + \Lambda_{\text{m}}^{\circ}(\text{NaOH}) - \Lambda_{\text{m}}^{\circ}(\text{NaCl})$$

7. (b) In strong electrolyte, number of ions remains constant so, equivalent conductance increases due to increase in ionic mobility.

8. (b) The unit of conductivity (specific conductivity) is $\text{ohm}^{-1} \text{cm}^{-1}$ or S cm^{-1} .

9. (d) Molar conductivity, $\Lambda_{\text{m}} = \frac{\kappa \times 1000}{C}$

Given, $\kappa = 0.0248 \text{ S cm}^{-1}$, $C = 0.20 \text{ M}$

$$\therefore \Lambda_{\text{m}} = \frac{0.0248 \times 1000}{0.20} = 124 \text{ S cm}^2 \text{ mol}^{-1}$$

10. (b) Equivalent conductance is related to specific conductance as, $\Lambda_{\text{c}} = \frac{\kappa \times 1000}{C}$

where, C is the number of gram equivalents of solute dissolved in one litre of the solution.

11. (a) The substance whose aqueous solution allows the passage of electric current and is chemically decomposed, is termed as electrolyte. Electrolytic substances are classified as strong or weak, according to acid-base concept. Acetic acid is a weak electrolyte while glucose, ethanol and urea are non-electrolyte.

12. (b) Pure water does not contains any mineral ions thus, it does not conduct electricity.

13. (c) From Faraday's first law, $\text{Mass}(w) = z \cdot Q$
 when, $Q = 1 \text{ C}$
 $w = z = \text{electrochemical equivalent.}$

14. (c) When iron is rusted, it is oxidised to hydrated ferric oxide.

