

# Electrochemistry

## Question1

Electrolysis of aqueous copper (II) sulphate between Pt electrodes gives ' X ' at anode and ' Y ' at cathode. X and Y are respectively.

TG EAPCET 2025 (Online) 2nd May Evening Shift

Options:

A.

Cu, O<sub>2</sub>

B.

O<sub>2</sub>, Cu

C.

SO<sub>2</sub>, H<sub>2</sub>

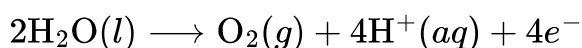
D.

O<sub>2</sub>, H<sub>2</sub>

**Answer: B**

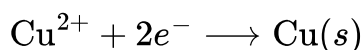
**Solution:**

Reaction at anode is,



This produce gas = O<sub>2</sub>(X)

Reaction at cathode



Metal deposited = Cu(Y)



## Question2

At 298 K, if emf of the cell corresponding to the reaction  $\text{Zn}(s) + 2\text{H}^+(aq) \longrightarrow \text{Zn}^{2+}(0.01\text{M}) + \text{H}_2(g)(1 \text{ atm})$  is 0.28 V, then the pH of the solution at the hydrogen electrode is

$$\left( \frac{2.303RT}{F} = 0.06 \text{ V} \right), \left( E_{\text{Zn}^{2+}/\text{Zn}}^{\circ} = -0.76 \text{ V} \right)$$

### TG EAPCET 2025 (Online) 2nd May Morning Shift

Options:

A.

8

B.

7

C.

9

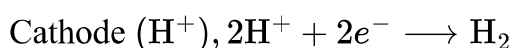
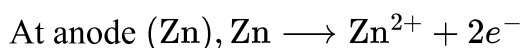
D.

10

**Answer: C**

### Solution:

The cell reaction is,



Using Nernst equation,

$$E_{\text{cell}} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} - \frac{2.303RT}{F} \log Q$$

$$0.28 = +0.76 - \frac{0.06}{2} \log \left( \frac{0.01}{[\text{H}^+]^2} \right)$$



Solving this,

$$[\text{H}^+] = 10^{-9}$$

$$\text{pH} = -\log [\text{H}^+] = 9$$

---

### Question3

**0.592 g of copper is deposited in 60 minutes by passing**

**0.5 A current through a solution of copper (II) sulphate. The electrochemical equivalent of copper (II) (in  $\text{gC}^{-1}$ ) is**

(  $F = 96500\text{Cmol}^{-1}$  )

**TG EAPCET 2024 (Online) 11th May Morning Shift**

**Options:**

A.  $3.3 \times 10^{-3}$

B.  $3.3 \times 10^{-4}$

C.  $6.6 \times 10^{-3}$

D.  $6.6 \times 10^{-4}$

**Answer: B**

**Solution:**

Given, Mass of copper = 0.592 g, time = 60 minutes, current = 0.5 A Total charge passes through solution,

$$Q = I \times t$$

$$Q = 0.5 \times 60 \times 60 = 1800\text{C}$$

Number of moles of electrons transferred.

$$n = \frac{Q}{F} = \frac{1800\text{C}}{96500\text{C/mol}}$$

$$\Rightarrow n = 0.01865 \text{ mol}$$



$\text{Cu}^{2+}$  requires 2 moles of electron to deposit 1 mole of copper

$$\text{Moles of copper} = \frac{n}{2} = \frac{0.01865}{2}$$

$$= 0.009325 \text{ mol}$$

Mass of copper deposited

$$= 0.009325 \times 63.55$$

$$= 0.592 \text{ g}$$

Electrochemical equivalent is given by

$$\frac{\text{Mass of copper deposited}}{\text{Charged passed}} = \frac{0.592}{180}$$

$$= 3.3 \times 10^{-4} \text{ g/C}$$

---

## Question4

The standard electrode potentials  $E^\circ$  (V) for  $\text{Li}^+/\text{Li}$ ,  $\text{Na}^+/\text{Na}$  respectively are

**TG EAPCET 2024 (Online) 10th May Evening Shift**

**Options:**

A.  $-3.04, -2.714$

B.  $-2.714, -3.04$

C.  $-3.04, -3.04$

D.  $-2.714, -2.714$

**Answer: A**

**Solution:**

The standard electrode potentials for the  $\text{Li}^+/\text{Li}$  and  $\text{Na}^+/\text{Na}$  half-cells are as follows:

$$\text{Li}^+/\text{Li}: E^\circ = -3.04 \text{ V}$$

$$\text{Na}^+/\text{Na}: E^\circ = -2.714 \text{ V}$$

Therefore, the correct option is:

**Option A**



-3.04, -2.714

These values indicate the tendency of lithium and sodium to lose electrons and form cations, with lithium being more reactive in this context due to its more negative electrode potential.

---

## Question 5

Two statements are given below.

**Statement I : Molten NaCl is electrolysed using Pt electrodes. Cl<sub>2</sub> is liberated at anode.**

**Statement II : Aqueous CuSO<sub>4</sub> is electrolysed using Pt electrodes. O<sub>2</sub> is liberated at cathode.**

The correct answer is

### TG EAPCET 2024 (Online) 10th May Evening Shift

Options:

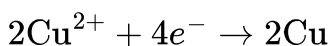
- A. Both statements I and II are correct.
- B. Both statements I and II are not correct.
- C. Statement I is correct but statement II is not correct.
- D. Statement I is not correct but statement II is correct.

**Answer: C**

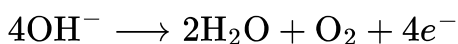
### Solution:

Statement I is correct but statement II is incorrect. Correct form of Statement II is

At cathode (Reduction take place)



At anode (oxidation)



Hence, oxygen is liberated at anode not cathode.

---



## Question6

The molar conductivity of 0.02 M solution of an electrolyte is  $124 \times 10^{-4} \text{ S m}^2 \text{ mol}^{-1}$ . What is the resistance of same solution (in ohms), kept in a cell of cell constant  $129 \text{ m}^{-1}$  ?

**TG EAPCET 2024 (Online) 10th May Morning Shift**

**Options:**

A. 390

B. 260

C. 130

D. 520

**Answer: D**

**Solution:**

Given,

Molar conductivity

$$= 124 \times 10^{-4} \text{ S m}^2 / \text{mol}$$

Molarity = 0.02M,

Cell constant =  $129 \text{ m}^{-1}$

Molar conductivity is given by

$$\Lambda_m = \frac{1000 \times \kappa}{M}$$

where  $\kappa$  is conductivity

$$K = \frac{\Lambda_m \times M}{1000}$$

Substitute the values in Eq. (i),

$$\kappa = \frac{124 \times 10^{-4} \times 0.02}{1000} \Rightarrow \kappa = 0.248 \dots$$

Conductivity,  $\kappa$  is given by



$$K = \frac{\text{Cell constant}}{\text{Resistance}}$$

$$\text{or Resistance} = \frac{\text{Cell constant}}{K}$$

$$R = \frac{129}{0.248} \text{ or } R = 520\Omega$$

---

## Question 7

If the degree of dissociation of formic acid is 11.0%, the molar conductivity of 0.02 M solution of it is (Given,  $\lambda^\circ(\text{H}^+) = 349.6 \text{ S cm}^2 \text{ mol}^{-1}$   $\lambda^\circ(\text{HCOO}^-) = 54.6 \text{ S cm}^2 \text{ mol}^{-1}$ )

**TG EAPCET 2024 (Online) 9th May Evening Shift**

**Options:**

A.  $44.46 \text{ S m}^2 \text{ mol}^{-1}$

B.  $44.46 \text{ S cm}^2 \text{ mol}^{-1}$

C.  $22.23 \text{ S m}^2 \text{ mol}^{-1}$

D.  $22.23 \text{ S cm}^2 \text{ mol}^{-1}$

**Answer: B**

**Solution:**

Using Kohlrausch's law we get, (for formic acid)

$$\lambda^\circ_{(\text{HCOOH})} = \lambda^\circ_{(\text{H}^+)} + \lambda^\circ_{(\text{HCOO}^-)}$$

$$\lambda^\circ_{(\text{HCOOH})} = 349.6 + 54.6$$

$$\lambda^\circ_{(\text{HCOOH})} = 404.2 \text{ S cm}^2/\text{mol}$$

Degree of dissociation is given by

$$\alpha = \frac{\lambda_{(\text{HCOOH})}}{\lambda^\circ_{(\text{HCOOH})}}$$

[where,  $\alpha = 0.11$  (as 11% given)]



$\lambda_{(\text{HCOOH})} =$  Molar conductivity of HCOOH

$$\lambda_{(\text{HCOOH})} = \alpha \times \lambda_{(\text{HCOOH})}^{\circ}$$

$$= 0.11 \times 404 : 2$$

$$= 44.46 \text{ S cm}^2/\text{mol}$$

---

## Question8

Identify the correct statements from the following

(A) At 298 K, the potential of hydrogen electrode placed in a solution of pH = 10, is -0.59 V

(B) The limiting molar conductivity of  $\text{Ca}^{2+}$  and  $\text{Cl}^{-}$  is 119 and  $76 \text{ S cm}^2 \text{ mol}^{-1}$  respectively. The limiting molar conductivity of  $\text{CaCl}_2$  is  $195 \text{ S cm}^2 \text{ mol}^{-1}$

(C) The correct relationship between  $K_C$  and  $E_{\text{cell}}^{\ominus}$  is

$$E_{\text{cell}}^{\ominus} = \frac{2303RT}{nF} \log K_C$$

**TG EAPCET 2024 (Online) 9th May Morning Shift**

Options:

A. A, B, C

B. A, B only

C. A, C only

D. B, C only

**Answer: C**

**Solution:**

Correct Statements

To evaluate the given statements, we analyze each one:

**Statement (A):** The potential of the hydrogen electrode in a solution of  $\text{pH} = 10$  at 298 K.

The potential of a hydrogen electrode can be calculated using the Nernst equation for the hydrogen electrode:

$$E = E^\ominus - \frac{0.0591}{n} \log[\text{H}^+]$$

At standard conditions  $E^\ominus = 0$ , and  $[\text{H}^+]$  is calculated from the pH:

$$\text{pH} = 10 \implies [\text{H}^+] = 10^{-10}$$

Substitute into the Nernst equation:

$$E = -0.0591 \times \log(10^{-10}) = -0.59 \text{ V}$$

Statement A is correct.

---

**Statement (B):** Limiting molar conductivity of  $\text{Ca}^{2+}$  and  $\text{Cl}^-$  are 119 and 76  $\text{S cm}^2 \text{ mol}^{-1}$  respectively.

The limiting molar conductivity of  $\text{CaCl}_2$  is the sum of the molar conductivities of  $\text{Ca}^{2+}$  and two  $\text{Cl}^-$  ions:

$$\lambda_{\text{CaCl}_2}^\circ = \lambda_{\text{Ca}^{2+}}^\circ + 2\lambda_{\text{Cl}^-}^\circ$$

Substitute the given values:

$$\lambda_{\text{CaCl}_2}^\circ = 119 + 2(76) = 119 + 152 = 271 \text{ S cm}^2 \text{ mol}^{-1}$$

The statement given as 195  $\text{S cm}^2 \text{ mol}^{-1}$  is incorrect.

---

**Statement (C):** Relationship between  $K_C$  and  $E_{\text{cell}}^\ominus$ .

The relationship between the standard cell potential and the equilibrium constant is given by the Nernst equation:

$$E_{\text{cell}}^\ominus = \frac{RT}{nF} \ln K_C$$

Converting natural log to base-10 log:

$$E_{\text{cell}}^\ominus = \frac{2.303RT}{nF} \log K_C$$

The statement is correctly stated.

## Conclusion

Based on the analysis, the correct statements are A and C.

**Option C: A, C only** is the correct answer.

---



## Question9

At 300 K , the conductivity of  $0.01 \text{ mol dm}^{-3}$  aqueous solution of acetic acid is  $19.5 \times 10^{-5} \text{ mhocm}^{-1}$  and limiting molar conductivity of acetic acid at the same temperature is  $390 \text{ mhocm}^2 \text{ mol}^{-1}$ . The degree of dissociation of acetic acid is

TS EAMCET 2023 (Online) 12th May Evening Shift

Options:

A.  $5.0 \times 10^{-5}$

B.  $5.0 \times 10^{-2}$

C.  $2.5 \times 10^{-5}$

D.  $7.5 \times 10^{-2}$

**Answer: B**

**Solution:**

To determine the degree of dissociation of acetic acid, follow these steps:

**Concentration of Acetic Acid:**

$$C = 0.01 \text{ mol dm}^{-3} = 10 \text{ mol m}^{-3}$$

**Conductivity ( $\kappa$ ):**

$$\kappa = 19.5 \times 10^{-5} \text{ S cm}^{-1} = 19.5 \times 10^{-3} \text{ S m}^{-1}$$

**Limiting Molar Conductivity ( $\lambda_m^\circ$ ):**

$$\lambda_m^\circ = 390 \text{ S cm}^2 \text{ mol}^{-1} = 390 \times 10^{-4} \text{ S m}^2 \text{ mol}^{-1}$$

**Calculate Molar Conductivity ( $\lambda_m$ ):**

$$\lambda_m = \frac{\kappa}{C} = \frac{19.5 \times 10^{-3}}{10} = 19.5 \times 10^{-4} \text{ S m}^2 \text{ mol}^{-1}$$

**Degree of Dissociation ( $\alpha$ ):**

$$\alpha = \frac{\lambda_m}{\lambda_m^\circ} = \frac{19.5 \times 10^{-4}}{390 \times 10^{-4}} = 0.05 = 5 \times 10^{-2}$$

Therefore, the degree of dissociation of acetic acid is  $5 \times 10^{-2}$ .

-----



## Question 10

The electrode potential of chlorine electrode is maximum, when the concentration of chloride ion in the solution (in  $\text{molL}^{-1}$ ) is  $X$ . What is the value of  $X$  ?

TS EAMCET 2023 (Online) 12th May Morning Shift

Options:

A.  $2.5 \times 10^{-3}$

B.  $7.5 \times 10^{-3}$

C.  $7.5 \times 10^{-2}$

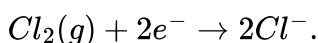
D.  $2.5 \times 10^{-2}$

Answer: A

Solution:

Let's analyze the problem step by step.

The chlorine electrode involves the half-reaction



Its potential at  $25^\circ\text{C}$  is given by the Nernst equation

$$E = E^\circ - \frac{0.059}{2} \log \frac{[\text{Cl}^-]^2}{P_{\text{Cl}_2}},$$

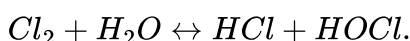
where  $P_{\text{Cl}_2}$  is the partial pressure of chlorine gas (in atm) and  $[\text{Cl}^-]$  is the activity (or concentration) of the chloride ion.

Note the following:

For fixed  $P_{\text{Cl}_2}$  (usually taken as 1 atm), the electrode potential  $E$  is a decreasing function of the logarithm of  $[\text{Cl}^-]^2$ .

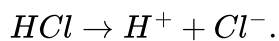
Therefore, to get the maximum electrode potential, the ratio  $\frac{[\text{Cl}^-]^2}{P_{\text{Cl}_2}}$  must be as small as possible—that is,  $[\text{Cl}^-]$  must be minimized.

In practice the “chlorine electrode” is prepared in a solution with no added chloride. However, when chlorine gas is bubbled into pure water, it undergoes the following hydrolysis:



Since HCl is a strong acid, it dissociates completely:





In the absence of any added chloride, the only source of  $[Cl^-]$  in the solution is from this hydrolysis. Experiments have shown that under these conditions (i.e. when the solution is free of extrinsic chloride) the concentration of chloride ion attains a minimum value.

It turns out that this minimum (and hence the maximum electrode potential) corresponds to a chloride concentration of approximately

$$2.5 \times 10^{-3} \text{ mol L}^{-1}.$$

Thus, among the options given, the maximum potential is achieved when

$$X = 2.5 \times 10^{-3} \text{ mol L}^{-1}.$$

So, the correct answer is Option A.

---

