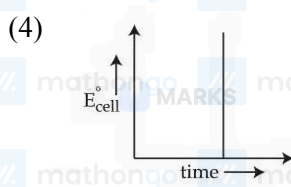
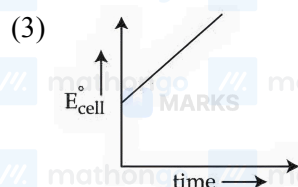
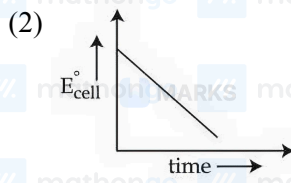
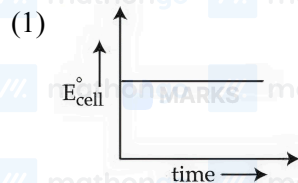
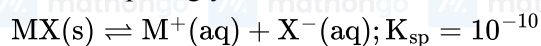


**Q1. JEE Main 2026 (21 January Shift 2)**

For a closed circuit Daniell cell, which of the following plots is the accurate one at a given temperature?

**Q2. JEE Main 2026 (21 January Shift 2)**

MX is a sparingly soluble salt that follows the given solubility equilibrium at 298 K.



If the standard reduction potential for  $\text{M}^+(\text{aq}) \xrightarrow{+e^-} \text{M(s)}$  is  $(E_{\text{M}^+/\text{M}}^\ominus) = 0.79 \text{ V}$ , then the value of the standard reduction potential for the metal/metal insoluble salt electrode  $E_{\text{X}^-/\text{MX(s)}/\text{M}}^\ominus$  is \_\_\_\_\_ mV. (nearest integer)

[Given :  $\frac{2.303RT}{F} = 0.059 \text{ V}$ ]

**Q3. JEE Main 2026 (23 January Shift 1)**

In the given electrochemical cell,  $\text{Ag(s)}|\text{AgCl(s)}|\text{FeCl}_2(\text{aq}), \text{FeCl}_3(\text{aq})|\text{Pt(s)}$  at 298 K, the cell potential ( $E_{\text{cell}}$ ) will increase when :

- A. Concentration of  $\text{Fe}^{2+}$  is increased.
- B. Concentration of  $\text{Fe}^{3+}$  is decreased.
- C. Concentration of  $\text{Fe}^{2+}$  is decreased.
- D. Concentration of  $\text{Fe}^{3+}$  is increased.
- E. Concentration of  $\text{Cl}^-$  is increased.

Choose the correct answer from the options given below :

- (1) B Only
- (2) A and E Only
- (3) C, D and E Only
- (4) A and B Only

**Q4. JEE Main 2026 (28 January Shift 2)**

For strong electrolyte  $\Lambda_m$  increases slowly with dilution and can be represented by the equation  $\Lambda_m = \Lambda_m^\circ - Ac^{1/2}$

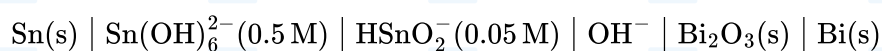
Molar conductivity values of the solutions of strong electrolyte AB at 18°C are given below :

|                                                    |      |      |      |      |
|----------------------------------------------------|------|------|------|------|
| c [mol L <sup>-1</sup> ]                           | 0.04 | 0.09 | 0.16 | 0.25 |
| $\Lambda_m$ [S cm <sup>2</sup> mol <sup>-1</sup> ] | 96.1 | 95.7 | 95.3 | 94.9 |

The value of constant A based on the above data [ in Scm<sup>2</sup> mol<sup>-1</sup>/(mol/L)<sup>1/2</sup> ] unit is \_\_\_\_.

**Q5. JEE Main 2026 (28 January Shift 2)**

A volume of  $x$  mL of 5 M NaHCO<sub>3</sub> solution was mixed with 10 mL of 2 M H<sub>2</sub>CO<sub>3</sub> solution to make an electrolytic buffer. If the same buffer was used in the following electrochemical cell to record a cell potential of 235.3 mV, then the value of  $x$  = \_\_\_\_ mL (nearest integer).



Consider up to one place of decimal for intermediate calculations.

Given:  $E^\circ_{[\text{Sn(OH)}_6]^{2-}/\text{HSnO}_2^-} = -0.90 \text{ V}$

$$E^\circ_{\text{Bi}_2\text{O}_3/\text{Bi}} = -0.44 \text{ V}$$

$$\text{p}K_a(\text{H}_2\text{CO}_3) = 6.11$$

$$\frac{2.303 RT}{F} = 0.059 \text{ V}$$

$$\text{Antilog}(1.29) = 19.5$$

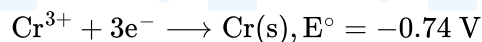
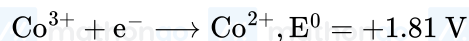
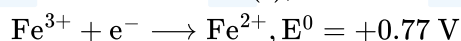
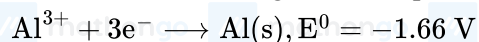
**Q6. JEE Main 2026 (24 January Shift 2)**

Molar conductivity of a weak acid HQ of concentration 0.18 M was found to be 1/30 of the molar conductivity of another weak acid HZ with concentration of 0.02 M. If  $\lambda^\circ \text{Q}^-$  happened to be equal with  $\lambda^\circ \text{Z}^-$ , then the difference of the  $\text{p}K_a$  values of the two weak acids ( $\text{p}K_a(\text{HQ}) - \text{p}K_a(\text{HZ})$ ) is \_\_\_\_ (Nearest integer).

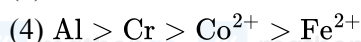
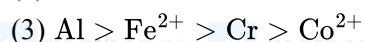
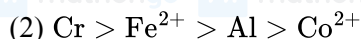
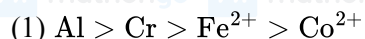
[Given: degree of dissociation ( $\alpha$ )  $\ll 1$  for both weak acids,  $\lambda^\circ$  : limiting molar conductivity of ions]

**Q7. JEE Main 2026 (22 January Shift 2)**

Consider the following reduction processes :



The tendency to act as reducing agent decreases in the order :

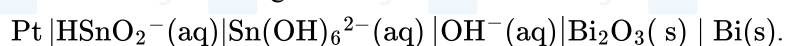
**Q8. JEE Main 2026 (22 January Shift 2)**

Consider the following electrochemical cell :  $\text{Pt} | \text{O}_2(\text{g})(1\text{bar}) | \text{HCl}(\text{aq}) || \text{M}^{2+}(\text{aq}, 1.0\text{M}) | \text{M(s)}$  The pH above which, oxygen gas would start to evolve at anode is \_\_\_\_ (nearest integer).

[ Given:  $E^{\circ}_{\text{M}^{2+}/\text{M}} = 0.994 \text{ V}$   
 $E^{\circ}_{\text{O}_2/\text{H}_2\text{O}} = 1.23 \text{ V}$  } standard reduction potential ]

**Q9. JEE Main 2026 (22 January Shift 1)**

Consider the following electrochemical cell at 298 K

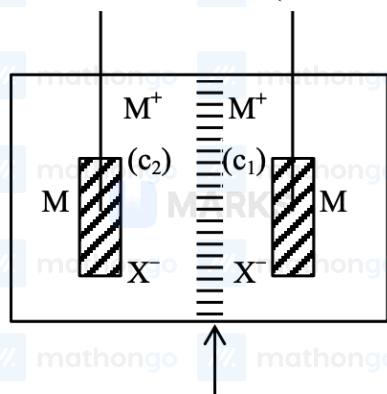


If the reaction quotient at a given time is  $10^6$ , then the cell EMF ( $E_{\text{cell}}$ ) is \_\_\_\_  $\times 10^{-1} \text{ V}$  (Nearest integer).

Given the standard half-cell reduction potential as



## Q10. JEE Main 2026 (23 January Shift 2)



Consider the above electrochemical cell where a metal electrode (M) is undergoing redox reaction by forming  $M^+$  ( $M \rightarrow M^+ + e^-$ ). The cation  $M^+$  is present in two different concentrations  $c_1$  and  $c_2$  as shown above. Which of the following statement is correct for generating a positive cell potential?

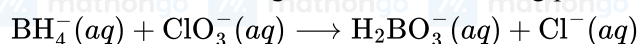
- (1) If  $c_1$  is present at cathode, then  $c_1 < c_2$ .
- (2) If  $c_1$  is present at anode, then  $c_1 > c_2$ .
- (3) If  $c_1$  is present at cathode, then  $c_1 > c_2$ .
- (4) If  $c_1$  is present at anode, then  $c_1 = c_2$ .

## Q11. JEE Main 2026 (21 January Shift 1)

The pH and conductance of a weak acid (HX) was found to be 5 and  $4 \times 10^{-5}$  S, respectively. The conductance was measured under standard condition using a cell where the electrode plates having a surface area of  $1 \text{ cm}^2$  were at a distance of 15 cm apart. The value of the limiting molar conductivity is \_\_\_\_\_  $\text{Sm}^2 \text{ mol}^{-1}$ . (nearest integer) (Given : degree of dissociation of the weak acid ( $\alpha$ )  $\ll 1$ )

## Q12. JEE Main 2026 (28 January Shift 1)

Consider the following redox reaction taking place in acidic medium



If the Nernst equation for the above balanced reaction is

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{RT}{nF} \ln Q,$$

then the value of  $n$  is \_\_\_\_\_.

(Nearest integer)

**Q13. JEE Main 2026 (24 January Shift 1)**

Electricity is passed through an acidic solution of  $\text{Cu}^{2+}$  till all the  $\text{Cu}^{2+}$  was exhausted, leading to the deposition of 300 mg of Cu metal. However, a current of 600 mA was continued to pass through the same solution for another 28 minutes by keeping the total volume of the solution fixed at 200 mL. The total volume of oxygen evolved at STP during the entire process is \_\_\_\_\_ mL. (Nearest integer)

[Given:  $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$   $E_{\text{red}}^{\circ} = +0.34 \text{ V}$

$\text{O}_2(\text{g}) + 4\text{H}^+ + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}$   $E_{\text{red}}^{\circ} = +1.23 \text{ V}$

Molar mass of Cu =  $63.54 \text{ g mol}^{-1}$

Molar mass of  $\text{O}_2$  =  $32 \text{ g mol}^{-1}$

Faraday Constant =  $96500 \text{ C mol}^{-1}$

Molar volume at STP =  $22.4 \text{ L}$ ]

## ANSWERS AND SOLUTIONS

1. (1)      2. 200      3. (3)      4. 4      5. 78      6. 2      7. (1)      8. 4  
9. 4      10. (3)      11. 6      12. 24      13. 111

- (1) The standard cell potential  $E_{cell}^{\circ}$  is a thermodynamic property defined for a cell under standard conditions (concentrations of all species at 1 M, pressure at 1 bar, and a specific temperature).

For a Daniell cell, the reaction is  $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$ .

The standard cell potential is given by  $E_{cell}^{\circ} = E_{Cu^{2+}/Cu}^{\circ} - E_{Zn^{2+}/Zn}^{\circ}$ .

Since  $E_{cell}^{\circ}$  depends only on the nature of the electrodes, the nature of the electrolytes, and the temperature, it remains constant over time as long as the temperature is constant.

While the actual cell potential  $E_{cell}$  changes as the reaction proceeds (due to changes in concentration according to the Nernst equation), the standard cell potential  $E_{cell}^{\circ}$  is a constant value.

Therefore, the plot of  $E_{cell}^{\circ}$  versus time is a horizontal line parallel to the time axis.
- (200) The standard reduction potential for a metal/metal insoluble salt electrode is related to the standard reduction potential of the metal ion and the solubility product  $K_{sp}$  of the salt.

The half-cell reaction for the  $X^{-}/MX(s)/M$  electrode is:  $MX(s) + e^{-} \rightarrow M(s) + X^{-}(aq)$ .

This can be viewed as a combination of two processes:

  - $MX(s) \rightleftharpoons M^{+}(aq) + X^{-}(aq)$  with equilibrium constant  $K_{sp} = 10^{-10}$
  - $M^{+}(aq) + e^{-} \rightarrow M(s)$  with  $E_{M^{+}/M}^{\circ} = 0.79$  V

The relation between the standard potentials is given by:

$$E_{X^{-}/MX/M}^{\circ} = E_{M^{+}/M}^{\circ} + \frac{2.303RT}{nF} \log_{10} K_{sp}$$

Given  $n = 1$  and  $\frac{2.303RT}{F} = 0.059$  V:

$$E_{X^{-}/MX/M}^{\circ} = 0.79 + 0.059 \log_{10}(10^{-10})$$

$$E_{X^{-}/MX/M}^{\circ} = 0.79 + 0.059(-10)$$

$$E_{X^{-}/MX/M}^{\circ} = 0.79 - 0.59 = 0.20$$
 V

To convert the value into mV:

$$0.20 \text{ V} = 0.20 \times 1000 \text{ mV} = 200 \text{ mV}$$
- (3) The cell  $Ag(s)|AgCl(s)|FeCl_2(aq), FeCl_3(aq)|Pt(s)$  has the reaction:

$$Ag(s) + Cl^{-} + Fe^{3+} \rightarrow AgCl(s) + Fe^{2+}$$

Using Nernst equation:

$$E_{cell} = E_{cell}^{\circ} - \frac{0.059}{n} \log Q \text{ where } Q = \frac{[Fe^{2+}]}{[Fe^{3+}][Cl^{-}]}$$

For  $E_{cell}$  to increase,  $Q$  must decrease. Analyzing each condition:

  - Increase  $[Fe^{2+}]$  increases  $Q$ , decreases  $E_{cell}$  - false
  - Decrease  $[Fe^{3+}]$  increases  $Q$  (smaller denominator), decreases  $E_{cell}$  - false
  - Decrease  $[Fe^{2+}]$  decreases  $Q$ , increases  $E_{cell}$  - true

D. Increase  $[\text{Fe}^{3+}]$  decreases  $Q$ , increases  $E_{\text{cell}}$  - true

E. Increase  $[\text{Cl}^-]$  decreases  $Q$ , increases  $E_{\text{cell}}$  - true

Therefore C, D, and E cause  $E_{\text{cell}}$  to increase.

4. (4) The equation  $\Lambda_m = \Lambda_m^0 - A c^{1/2}$  is linear in  $c^{1/2}$ .

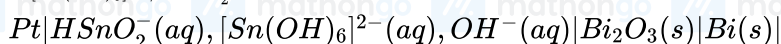
Taking two data points: ( $c=0.04$ ,  $\sqrt{c}=0.2$ ,  $\Lambda_m=96.1$ ) and ( $c=0.25$ ,  $\sqrt{c}=0.5$ ,  $\Lambda_m=94.9$ ).

$$\text{Slope} = \frac{94.9-96.1}{0.5-0.2} = \frac{-1.2}{0.3} = -4.0$$

Since slope =  $-A$ , the constant  $A = 4 \text{ S cm}^2 \text{ mol}^{-1}/(\text{mol/L})^{1/2}$

5. (78) We have considered

$$E^\circ_{[\text{Sn}(\text{OH})_6]^{2-}/\text{HSnO}_2^-} = -0.9 \text{ V}$$



$$0.5M \quad 0.05M$$

$$E_{\text{cell}}^\circ = +0.9 - 0.44 = 0.46 \text{ V}$$

**Oxidation Half :**



**Reduction Half :**



$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.059}{6} \log \frac{(0.5)^3}{(0.05)^3 \times [\text{OH}^-]^3}$$

$$0.2353 = 0.46 - \frac{0.059}{6} \times 3 \log \left[ \frac{10}{[\text{OH}^-]} \right]$$

$$\log \left[ \frac{10}{[\text{OH}^-]} \right] = \frac{2 \times 0.2247}{0.059} = 7.6$$

$$1 + p\text{OH} = 7.6$$

$$p\text{OH} = 6.6$$

$$p\text{H} = 14 - 6.6 = 7.4$$

$$p\text{H} = pK_{a1} + \log \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$$

$$7.4 = 6.11 + \log \frac{5x}{20}$$

$$1.29 = \log \frac{x}{4}$$

$$\frac{x}{4} = 19.5$$

$$x = 78$$

6. (2) For weak acids with  $\lambda^\infty Q^- = \lambda^\infty Z^-$ , the molar conductivity is  $\lambda_m = \alpha \lambda^\infty$ .

$$\text{From } \lambda_m^Q = \frac{1}{30} \lambda_m^Z, \text{ we get } \alpha_Q = \frac{\alpha_Z}{30}.$$

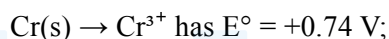
$$\text{Using } K_a = \alpha^2 C \text{ for weak acids: } K_a(\text{HQ}) = \alpha_Q^2 \times 0.18 = \frac{\alpha_Z^2 \times 0.18}{900} \text{ and } K_a(\text{HZ}) = \alpha_Z^2 \times 0.02.$$

Therefore  $\frac{K_a(HZ)}{K_a(HQ)} = \frac{0.02 \times 900}{0.18} = 100$ .

Thus  $pK_a(HQ) - pK_a(HZ) = \log(100) = 2$ .

7. (1) Reducing agent strength is determined by oxidation potential (opposite sign of reduction potential). Stronger reducing agents have more negative oxidation potentials (easier to oxidize).

Oxidation potentials:



Order of decreasing reducing agent tendency:  $\text{Al} > \text{Cr} > \text{Fe}^{2+} > \text{Co}^{2+}$ .

8. (4) In this electrochemical cell, at the anode oxygen evolution occurs from water oxidation. Using Nernst equation for the  $\text{O}_2/\text{H}_2\text{O}$  couple:  $E_{\text{anode}} = E^\circ_{\text{O}_2/\text{H}_2\text{O}} - \frac{0.059}{4} \log \frac{1}{[\text{H}^+]^4} = 1.23 - 0.059 \times pH$

At the cathode:  $E_{\text{cathode}} = 0.994 \text{ V}$  (standard potential for  $\text{M}^{2+}/\text{M}$  reduction)

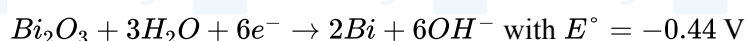
For oxygen to start evolving, the cell potential becomes zero:  $E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}} = 0$

$$0.994 = 1.23 - 0.059 \times pH$$

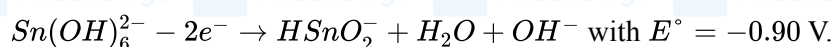
$$0.059 \times pH = 0.236$$

$$pH = 4.0$$

9. (4) The electrochemical cell has cathode:



and anode:



The standard cell EMF is  $E^\circ_{\text{cell}} = -0.44 - (-0.90) = 0.46 \text{ V}$ .

Using the Nernst equation with  $Q = 10^6$  and  $n = 6$  electrons transferred:

$$E_{\text{cell}} = 0.46 - \frac{0.059}{6} \log(10^6) = 0.46 - \frac{0.059 \times 6}{6} = 0.46 - 0.059 = 0.401 \text{ V}.$$

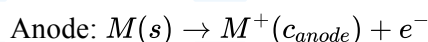
Expressing as  $x \times 10^{-1} \text{ V}$ :

$$0.401 = 4.01 \times 10^{-1} \text{ V, so } x = 4 \text{ (nearest integer).}$$

10. (3) The given setup is a concentration cell with identical electrodes (M) and different electrolyte concentrations ( $c_1$  and  $c_2$ ).

For a concentration cell, the standard cell potential  $E^\circ_{\text{cell}} = 0$ .

The cell reactions are:





Using the Nernst equation at 298 K:

$$E_{cell} = E_{cell}^0 - \frac{0.0591}{n} \log \frac{c_{anode}}{c_{cathode}}$$

$$E_{cell} = 0 - 0.0591 \log \frac{c_{anode}}{c_{cathode}}$$

For a positive cell potential ( $E_{cell} > 0$ ), the term  $\log \frac{c_{anode}}{c_{cathode}}$  must be negative.

This requires  $\frac{c_{anode}}{c_{cathode}} < 1$ , which means  $c_{cathode} > c_{anode}$ .

If  $c_1$  is present at the cathode and  $c_2$  is at the anode, then for  $E_{cell} > 0$ , we must have  $c_1 > c_2$ .

Checking the options:

(1) If  $c_1$  is at cathode,  $c_1 < c_2$  (Incorrect)

(2) If  $c_1$  is at anode,  $c_1 > c_2$  (Incorrect, would require  $c_2 > c_1$ )

(3) If  $c_1$  is at cathode,  $c_1 > c_2$  (Correct)

(4) If  $c_1$  is at anode,  $c_1 = c_2$  (Incorrect,  $E_{cell}$  would be zero)

11. (6) Given: pH = 5, so  $[H^+] = 10^{-5}$  M;  $G = 4 \times 10^{-5}$  S;  $A = 1$  cm<sup>2</sup>;  $l = 15$  cm

Cell constant =  $l/A = 15$  cm<sup>-1</sup>

Specific conductivity:  $\kappa = G \times (l/A) = 4 \times 10^{-5} \times 15 = 6 \times 10^{-4}$  S/cm

For weak acid HX:  $[H^+] = C\alpha = 10^{-5}$  M

Using  $\kappa = \frac{C\alpha \times \Lambda_m^\circ}{1000}$

$$\Lambda_m^\circ = \frac{1000 \times \kappa}{C\alpha} = \frac{1000 \times 6 \times 10^{-4}}{10^{-5}} = 6 \times 10^4 \text{ S cm}^2 \text{ mol}^{-1}$$

Converting:  $\Lambda_m^\circ = 6 \times 10^4 \times 10^{-4} = 6 \text{ S m}^2 \text{ mol}^{-1}$

12. (24)  $3BH_4^- + 4ClO_3^- \rightarrow 4Cl^- + 3H_2BO_3^- + 3H_2O$

n-factor = 8

moles = 3

$\therefore n = 3 \times 8 = 24$

13. (111) Eq of Cu = Eq of  $O_2$

$$\frac{300 \times 10^{-3} \times 2}{63.54} = n_{O_2} \times 4$$

$$2.36 \times 10^{-3} = n_{O_2}$$

When current is further passed

$$n_{O_2} \times 4 = \frac{600 \times 28 \times 60}{96500 \times 1000}$$

$$n_{O_2} = 2.611 \times 10^{-3}$$

Total  $O_2$  released

$$= [10^{-3} \times (2.36 + 2.611)] \times 22400 \text{ ml}$$

$$= 111.35 \text{ ml}$$