

Yakeen NEET 2.0 2026

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DPP: 3

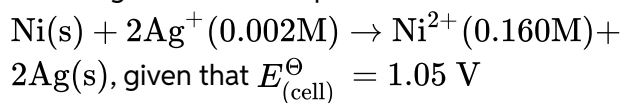
Electrochemistry

- Q1** The standard electrode potential of zinc ions is 0.76 V. What will be the potential of a 2M solution at 300 K ?
 (A) 0.83 V
 (B) 0.76 V
 (C) 0.23 V
 (D) 0.98 V
- Q2** The EMF of H-electrode if pH of electrolyte is 2 is [P = 1 atm]
 (A) $\frac{RT}{F}$
 (B) $\frac{RT}{2F}$
 (C) $\frac{2.303RT}{F}$
 (D) -0.118 V
- Q3** The Nernst equation giving dependence of electrode oxidation potential on concentration is
 (A) $E = E^\circ + \frac{2.303RT}{nF} \log [M^{+n}]$
 (B) $E = E^\circ - \frac{2.303RT}{nF} \log \frac{[M^{n+}]}{[M]}$
 (C) $E = E^\circ - \frac{2.303RT}{nF} \log [M^{n+}]$
 (D) $E = E^\circ + \frac{2.303RT}{nF} \log \frac{[M]}{[M^{n+}]}$
- Q4** The relationship between standard reduction potential of a cell and equilibrium constant is shown by
 (A) $E_{\text{cell}}^\circ = \frac{n}{0.059} \log K_C$
 (B) $E_{\text{cell}}^\circ = \frac{0.059}{n} \log K_C$
 (C) $E_{\text{cell}}^\circ = 0.059n \log K_C$
 (D) $E_{\text{cell}}^\circ = \frac{\log K_C}{n}$
- Q5** Which is the correct representation for the Nernst equation?
 (A) $E_{\text{RP}} = E_{\text{RP}}^\circ + \frac{0.059}{n} \log \frac{[\text{oxidant}]}{[\text{product}]}$
 (B) $E_{\text{OP}} = E_{\text{OP}}^\circ - \frac{0.059}{n} \log \frac{[\text{oxidant}]}{[\text{reductant}]}$
 (C) $E_{\text{OP}} = E_{\text{OP}}^\circ + \frac{0.059}{n} \log \frac{[\text{reductant}]}{[\text{oxidant}]}$
 (D) All of these
- Q6** For the cell $\text{Tl} | \text{Tl}^{+1} (0.001\text{M}) || \text{Cu}^{+2} (0.1\text{M}) | \text{Cu}$, E_{cell} at 25°C is 0.83 V. Which can be increased by
 (A) increasing $[\text{Cu}^{+2}]$
 (B) increasing $[\text{Tl}^{+}]$
 (C) decreasing $[\text{Cu}^{+2}]$
 (D) None of these
- Q7** How much will the potential of a hydrogen electrode change when its solution initially at pH = 0 is neutralized to pH = 7 ?
 (A) increase by 0.059 V
 (B) decrease by 0.059 V
 (C) increase by 0.41 V
 (D) decrease by 0.41 V
- Q8** For a cell involving one electron, $E_{\text{cell}}^\circ = 0.59$ V at 298 K, the equilibrium constant for the reaction is
- $$\frac{2.303RT}{F} = 0.059 \text{ V}$$
- (A) 10^{30} (B) 10^2



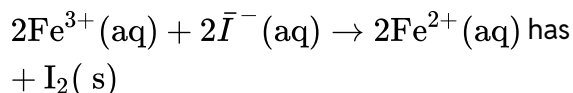
(C) 10^5 (D) 10^{10}

- Q9** Find the value of the emf of the cell in which the following reaction takes place :



(A) 0.23 volt (B) 0.31 volt
(C) 0.51 volt (D) 0.91 volt

- Q10** The cell in which the following reaction occurs :



$E_{(\text{cell})}^{\circ} = 0.236 \text{ V}$ at 298 K. Find the value of the standard Gibbs energy and the equilibrium constant of the cell reaction.

(A) $\Delta_r G^{\ominus} = -45.54 \text{ kJ mol}^{-1}$,

$$K_c = 9.62 \times 10^7$$

(B) $\Delta_r G^{\ominus} = -55.54 \text{ kJ mol}^{-1}$,

$$K_c = 7.62 \times 10^7$$

(C) $\Delta_r G^{\ominus} = -80.54 \text{ kJ mol}^{-1}$,

$$K_c = 3.62 \times 10^7$$

(D) $\Delta_r G^{\ominus} = -10.54 \text{ kJ mol}^{-1}$,

$$K_c = 1.62 \times 10^7$$



Answer Key

Q1 (B)

Q2 (D)

Q3 (C)

Q4 (B)

Q5 (D)

Q6 (A)

Q7 (D)

Q8 (D)

Q9 (D)

Q10 (A)



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