JEE Mains 2019 Chapter wise Question Bank

Electrochemistry - Questions

Q1

The anodic half-cell of lead-acid battery is recharged using electricity of 0.05 Faraday. The amount of $PbSO_4$ electrolyzed in g during the process is : (Molar mass of $PbSO_4 = 303\ g\ mol^{-1}$)

(1) 22.8 (2) 15.2
(3) 7.6 (4) 11.4

9 Jan Morning

Q2

If the standard electrode potential for a cell is 2 V at 300 K, the equilibrium constant (K) for the reaction

$$Zn(s) + Cu^{2+}(aq) \rightleftharpoons Zn^{2+}(aq) + Cu(s)$$

at 300 K is approximately

($R = 8 JK^{-1}\ mol^{-1}$, $F = 96000\ C\ mol^{-1}$)

(1) e^{-80} (2) e^{-160}
(3) e^{320} (4) e^{160}

9 Jan Evening

Q3

In the cell

$Pt(s) \mid H_2(g, 1bar) / HCl(aq) \| AgCl(s)/Ag(s) \mid Pt(s)$, the cell potential is 0.92 V when a 10^{-6} molal HCl solution is used. The standard electrode potential of $(AgCl/Ag, Cl^-)$ electrode is:

$$\left\{\text{Given}: \frac{2.303RT}{F} = 0.06\ V \text{ at } 298\ K\right\}$$

(1) 0.94 V (2) 0.76 V
(3) 0.40 V (4) 0.20 V

10 Jan Evening

Q4

For the cell $Zn(s) \mid Zn^{2+}(aq) \| M^{x+}(aq) \mid M(s)$, different half cells and their standard electrode potentials are given below:

$M^{x+}(aq)/M(s)$	$Au^{3+}(aq)/Au(s)$	$Ag^{+}(aq)/Ag(s)$	$Fe^{3+}(aq)/Fe^{2+}(aq)$	$Fe^{2+}(aq)/Fe(s)$
$E^o_{M^{x+}/M}/(V)$	1.40	0.80	0.77	–0.44

If $E^o_{Zn^{2+}/Zn} = -0.76V$, which cathode will give a maximum value of E°_{cell} per electron transferred?

(1) Ag^+/Ag (2) Fe^{3+}/Fe^{2+}
(3) Au^{3+}/Au (4) Fe^{2+}/Fe

11 Jan Morning

Q5

Given the equilibrium constant :

K_C of the reaction :

$Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$ is

10×10^{15} calculate the E^0_{cell} of this reaction at 298 K

$$\left[2.303\frac{RT}{F} \text{ at } 298\ K = 0.059\ V\right]$$

(1) 0.04736 mV (2) 0.4736 mV
(3) 0.4736 V (4) 0.04736 V

11 Jan Evening

Q6

The standard electrode potential E^0 and its temperature coefficient $\left(\frac{dE^0}{dT}\right)$ for a cell are 2 V and -5×10^{-4} VK^{-1} at 300 K respectively. The cell reaction is:

$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu$

The standard reaction enthalpy ($\Delta_r H^0$) at 300 K in kJ mol^{-1} is,

[Use R = 8 JK^{-1} mol^{-1} and F = 96,000 C mol^{-1}]

(1) –412.8 (2) –384.0

(3) 192.0 (4) 206.4

12 Jan Morning

Q7

Λ_m° for NaCl, HCl and NaA are 126.4, 425.9 and 100.5 S $cm^2 mol^{-1}$, respectively. If the conductivity of 0.001 M HA is 5×10^{-5} S cm^{-1}, degree of dissociation of HA is :

(1) 0.50 (2) 0.25

(3) 0.125 (4) 0.75

12 Jan Evening

Q8

Given that $E^\circ_{O_2/H_2O} = +1.23V$;

$E^\circ_{S_2O_8^{2-}/SO_4^{2-}} = 2.05V$

$E^\circ_{Br_2/Br^-} = +1.09V$

$E^\circ_{Au^{3+}/Au} = +1.4V$

The strongest oxidising agent is :

(1) Au^{3+} (2) O_2 (3) $S_2O_8^{2-}$ (4) Br_2

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Q9

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

$Fe^{2+}(aq) + Ag^+(aq) \rightarrow Fe^{3+}(aq) + Ag(s)$

Given that

$E^0_{Ag^+/Ag} = x$ V

$E^0_{Fe^{2+}/Fe} = y$ V

$E^0_{Fe^{3+}/Fe} = z$ V

(1) $x - z$ (2) $x - y$

(3) $x + 2y - 3z$ (4) $x + y - z$

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Q10

The standard Gibbs energy for the given cell reaction in kJ mol^{-1} at 298 K is:

$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$,

$E^\circ = 2$ V at 298 K

(Faraday's constant, F = 96000 C mol^{-1})

(1) – 384 (2) 384 (3) 192 (4) – 192

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Q11

A solution of Ni $(NO_3)_2$ is electrolysed between platinum electrodes using 0.1 Faraday electricity. How many mode of Ni will be deposited at the cathode?

(1) 0.05 (2) 0.20 (3) 0.15 (4) 0.10

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Q11

Consider the statements S1 and S2 :

S1 : Conductivity always increases with decrease in the concentration of electrolyte.

S2 : Molar conductivity always increases with decrease in the concentration of electrolyte.

The correct option among the following is :

(1) Both S1 and S2 are wrong
(2) S1 is wrong and S2 is correct
(3) Both S1 and S2 are correct
(4) S1 is correct and S2 is wrong

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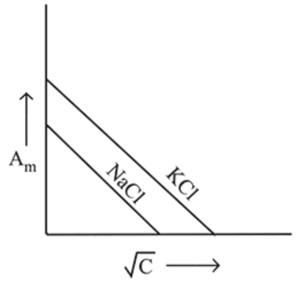
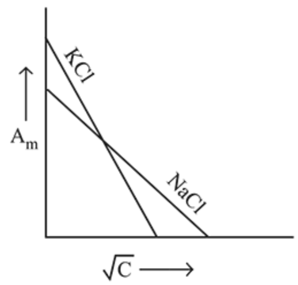
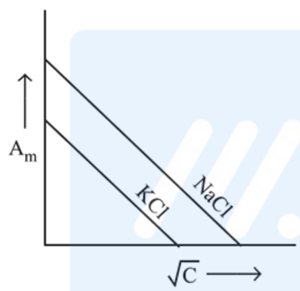
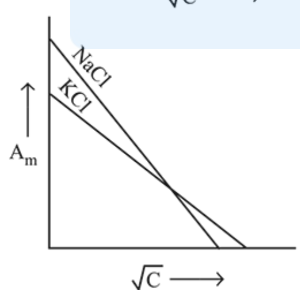
Q12

Which one of the following graphs between molar conductivity (A_m) versus $\sqrt{C}$ is correct?

(1) 

(2) 

(3) 

(4) 

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Q13

Given :

$Co^{3+} + e^- \rightarrow Co^{2+}$; $E^o = + 1.81$ V

$Pb^{4+} + 2e^- \rightarrow Pb^{2+}$; $E^o = + 1.67$ V

$Ce^{4+} + e^- \rightarrow Ce^{3+}$; $E^o = + 1.61$ V

$Bi^{3+} + 3e^- \rightarrow Bi$; $E^o = + 0.20$ V

oxidizing power of the species will increase in the order:

(1) $Ce^{4+} < Pb^{4+} < Bi^{3+} < Co^{3+}$

(2) $Bi^{3+} < Ce^{4+} < Pb^{4+} < Co^{3+}$

(3) $Co^{3+} < Ce^{4+} < Bi^{3+} < Pb^{4+}$

(4) $Co^{3+} < Pb^{4+} < Ce^{4+} < Bi^{3+}$

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JEE Mains 2019 Chapter wise Question Bank

Electrochemistry - Answers

Q1

(3) Half cell reaction: $PbSO_4 \rightarrow Pb^{4+} + 2e^-$

According to the reaction:

$PbSO_4 \rightarrow Pb^{4+} + 2e^-$

We require 2F for the electrolysis of 1 mol or 303 g of $PbSO_4$

$\therefore$ Amount of $PbSO_4$ electrolysed by

$0.05F = \times .05 = 7.575$ g ≈ 7.6 g

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Q2

(4) We know that, $\ln K = \dfrac{nFE^\circ_{cell}}{RT}$

After putting the given values, we get

$\ln K = \dfrac{2 \times 96000 \times 2}{8 \times 300} = 160$

$\therefore \quad K = e^{160}$

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Q3

(4) Given that:

$Pt(s) \mid H_2(g, 1\ bar) / HCl(aq) \parallel AgCl(s) / Ag(s) \mid Pt(s)$

$E_{cell} = 0.92$ V

Now, $E_{cell} = E^\circ_{H_2(g)/H^+(aq)} + E^\circ_{AgCl(s)/Ag(s),Cl^-} - \dfrac{0.06}{n}\log Q$

Cell reaction: $\dfrac{1}{2}H_2(g) \longrightarrow H^+(aq) + e^-$

$AgCl(s) + e^- \longrightarrow Ag(s) + Cl^-(aq)$

Net cell reaction:

$\dfrac{1}{2}H_2(g) + AgCl(s) \longrightarrow Ag(s) + H^+ + Cl^-(aq)$

$\therefore \quad Q = \dfrac{[H^+][Cl^-]}{(P_{H_2})^{1/2}}$

Here, 10^{-6} molal HCl solution is used

So $\quad Q = \dfrac{10^{-6} \times 10^{-6}}{1} = 10^{-12}$

(assuming molality = molarity)

Now, $\quad 0.92 = E^\circ_{AgCl(s)/Ag(s),Cl^-} - \dfrac{0.06}{1}\log 10^{-12}$

$E^\circ_{AgCl(s)/Ag(s),Cl^-} = 0.92 + [0.06 \times (-12)]$

$= 0.92 - 0.72 = 0.20$ V

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Q4

(1) $Zn + 2Ag^{+} \longrightarrow 2Ag + Zn^{2+} \quad (n = 2e^{-})$

$E^{\circ}_{cell} = (E^{\circ}_{R.P.})_{cathode} - (E^{\circ}_{R.P.})_{anode}$

$= 0.80 - (-0.76) = 1.56 \text{ V for } 2e^{-}$

$\therefore \; E^{\circ}_{cell} \text{ for } 1e^{-} = \dfrac{1.56}{2} = 0.78 \text{ V}$

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Q5

(3) $E^{0}_{cell} = \dfrac{2.303RT}{nF}\log K_C$ or $E^{0}_{cell} = \dfrac{0.059V}{n}\log K_C$

$= \dfrac{0.059V}{2}\log 10^{16} = 0.4736 \text{ V}$

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Q6

(1) $\Delta G^{\circ} = \Delta_r H^{\circ} - T\Delta S^{\circ}$

$\Delta_r H^{\circ} = \Delta G^{\circ} + T\Delta S^{\circ}$

$\Delta_r H^{\circ} = -nFE^{\circ} + TnF\dfrac{dE^{\circ}}{dT}$

$\Delta_r H^{\circ} = -nFE^{\circ} + nFT\dfrac{dE^{\circ}}{dT}$

Cell reaction: $Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$

$\Delta_r H^{\circ} = -nF\left(E^{\circ} = \dfrac{TdE^{\circ}}{dT}\right)$

$\Delta_r H^{\circ} = -2 \times 96000(2 - 300 \times -5 \times 10^{-4})$

$\Delta_r H^{\circ} = -2 \times 96000(2 + 300 \times 5 \times 10^{-4})$

$= -2 \times 96000(2 + 0.15)$

$= -412.8 \times 10^{3}$ J/mol; $= -412.8$ kJ/mol

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Q7

(3) Given: $\Lambda^{\circ}_{m}(NaCl) = 126.4 \text{ S cm}^2 \text{ mol}^{-1}$, $\Lambda^{\circ}_{m}(HCl) = 425.9 \text{ S cm}^2 \text{ mol}^{-1}$ and $\Lambda^{\circ}_{m}(NaA) = 100.5 \text{ S cm}^2 \text{ mol}^{-1}$

$\Lambda^{\circ}_{m}(HA) = \lambda^{\circ}_{H^+} + \lambda^{\circ}_{A^-}$

$= \lambda^{\circ}_{H^+} + \lambda^{\circ}_{Cl^-} + \lambda^{\circ}_{A^-} + \lambda^{\circ}_{Na^+} - \lambda^{\circ}_{Cl^-} - \lambda^{\circ}_{Na^+}$

$= \Lambda^{\circ}_{m}(HCl) + \Lambda^{\circ}_{m}(NaA) - \Lambda^{\circ}_{m}(NaCl)$

$\Lambda^{\circ}_{m}(HA) = 429.5 - 126.4 + 100.5 = 400 \text{ S cm}^2 \text{ mol}^{-1}$

$\kappa(HA) = 5 \times 10^{-5} \text{ S cm}^{-1}$

$\Lambda_m \; (HA) = \dfrac{\kappa(HA) \times 1000}{\text{Molality of HA}} = \dfrac{5 \times 10^{-5} \times 1000}{0.001} = 50 \text{ S cm}^2 \text{ mol}^{-1}$

$\alpha = \dfrac{\Lambda_m(HA)}{\Lambda^{\circ}_m(HA)} = \dfrac{50}{400} = 0.125$

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Q8

(3) More positive is the reduction potential stronger is the oxidising agent. Reduction potential is maximum for $S_2O_8^{2-}$, therefore, it is the strongest oxidising agent amongst the given species.

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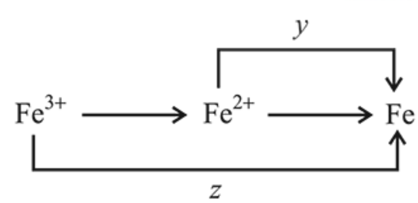
Q9

(3) $Ag^{+}(aq) + Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + Ag(s)$

$E^{0}_{cell} = E^{0}_{Ag^+/Ag} - E^{0}_{Fe^{3+}/Fe^{2+}}$



$E^{0}_{Fe^{3+}/Fe^{2+}} = 3z - 2y$

$E^{0}_{Ag^+/Ag} = x$ (given)

$E^{0}_{cell} = x - 3z + 2y$

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Q10

(1) $\Delta G^{\circ} = -nFE^{\circ}_{cell}$

$= -2 \; (96000) \; 2 \text{ V} = -384000 \text{ J/mol}$

$= -384 \text{ kJ/mol}$

9 April Morning

Q11

(1) According to the Faraday's law of electrolysis, nF of current is required for the deposition of 1 mol

According to the reaction,

$$Ni(NO_3)_2 \longrightarrow Ni^{2+} + 2NO_3^-$$

2 F of current deposits = 1 mol

$\therefore$ 0.1 F of current deposits = $\frac{0.1}{2}$ = 0.05 mol

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Q11

(2) Conductivity of an electrolyte is the conductance of 1 cm^3 of the given electrolyte. It increases with the increase in concentration of electrolyte due to increase in the number of ions per unit volume. Molar conductivity (λ_m) is the conductance of a solution containing 1 mole of the electrolyte. It increases with the decrease of concentration due to increase in the total volume having one mole of electrolyte. Thus, interionic attraction increses and degree of ionisation decreases. Therefore, (S_1) is wrong and (S_2) is correct.

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Q12

(1) Since, KCl is more conducting than NaCl, therefore, graph (1) is correct.

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Q13

(2) Higher the reduction potential, higher will be oxidising power. So,

$$Bi^{3+} < Ce^{4+} < Pb^{4+} < Co^{3+}$$

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