Electrochemistry

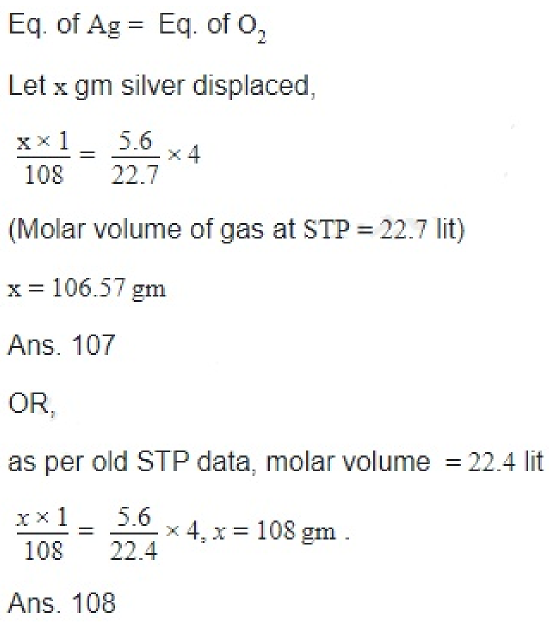
Question1

The mass of silver (Molar mass of Ag : $108 \mathrm{gmol}^{-1}$) displaced by a quantity of electricity which displaces 5600mL of O_2 at S.T.P. will be _____g.

[27-Jan-2024 Shift 1]

Answer: 108

Solution:



Eq. of Ag = Eq. of O_2

Let x gm silver displaced,

$$\frac{x \times 1}{108} = \frac{5.6}{22.7} \times 4$$

(Molar volume of gas at STP = 22.7 lit)

$x = 106.57$ gm

Ans. 107

OR,

as per old STP data, molar volume = 22.4 lit

$$\frac{x \times 1}{108} = \frac{5.6}{22.4} \times 4, x = 108 \text{ gm}.$$

Ans. 108

Question2

Which of the following statements is not correct about rusting of iron?

[27-Jan-2024 Shift 2]

Options:

A.

Coating of iron surface by tin prevents rusting, even if the tin coating is peeling off.

B.

When pH lies above 9 or 10 , rusting of iron does not take place.

C.

Dissolved acidic oxides SO_2, NO_2 in water act as catalyst in the process of rusting.

D.

Rusting of iron is envisaged as setting up of electrochemical cell on the surface of iron object.

Answer: A

Solution:

As tin coating is peeled off, then iron is exposed to atmosphere.

Question3

The hydrogen electrode is dipped in a solution of pH = 3 at 25°C. The potential of the electrode will be____ $\times 10^{-2}$V.

$\left(\frac{2.303\,RT}{F} = 0.059V\right)$

[27-Jan-2024 Shift 2]

Options:

Answer: 18

Solution:

$$2H^{+}_{(aq.)} + 2e^{-} \longrightarrow H_2(g)$$

$$E_{cell} = E_{cell}^{0} - \frac{0.059}{2} \log \frac{P_{H_2}}{[H^{+}]^{2}}$$

$$= 0 - 0.059 \times 3 = -0.177 \text{ volts.} \quad = -17.7 \times 10^{-2} V.$$

Question4

The mass of zinc produced by the electrolysis of zinc sulphate solution with a steady current of 0.015 A for 15 minutes is ____ $\times 10^{-4}$g.

(Atomic mass of zinc = 65.4amu)

[29-Jan-2024 Shift 1]

Options:

Answer: 45.75

Solution:

$$Zn^{+2} + 2e^{-} \longrightarrow Zn$$

$$W = Z \times i \times t$$

$$= \frac{65.4}{2 \times 96500} \times 0.015 \times 15 \times 60$$

$$= 45.75 \times 10^{-4} \text{ gm}$$

Question5

A constant current was passed through a solution of $AuCl_4^-$ion between gold electrodes. After a period of 10.0 minutes, the increase in mass of cathode was 1.314g. The total charge passed through the solution is _____ $\times 10^{-2}$F.

(Given atomic mass of $\overline{Au} = 197$)

[29-Jan-2024 Shift 2]

Options:

Answer: 2

Solution:

$$\frac{W}{E} = \frac{\text{charge}}{1F}$$

$$\frac{1.314}{\frac{197}{3}} = \frac{Q}{1F}$$

$$Q = 2 \times 10^{-2} F$$

Question6

Reduction potential of ions are given below:

ClO_4^-	IO_4^-	BrO_4^-
$E^\circ = 1.19V$	$E^\circ = 1.65V$	$E^\circ = 1.74V$

The correct order of their oxidising power is:

[30-Jan-2024 Shift 2]

Options:

A.

$ClO_4^- > IO_4^- > BrO_4^-$

B.

$BrO_4^- > IO_4^- > ClO_4^-$

C.

$BrO_4^- > ClO_4^- > IO_4^-$

D.

$IO_4^- > BrO_4^- > ClO_4^-$

Answer: B

Solution:

Higher the value of ⊕ ve SRP (Std. reduction potential) more is tendency to undergo reduction, so better is oxidising power of reactant.

Hence, ox. Power:- $BrO_4^- > IO_4^- > ClO_4^-$

Question7

Identify the factor from the following that does not affect electrolytic conductance of a solution.

[31-Jan-2024 Shift 1]

Options:

A.

The nature of the electrolyte added.

B.

The nature of the electrode used.

C.

Concentration of the electrolyte.

D.

The nature of solvent used.

Answer: B

Solution:

Conductivity of electrolytic cell is affected by concentration of electrolyte, nature of electrolyte and nature of solvent.

Question8

The metals that are employed in the battery industries are
A. Fe
B. Mn
C. Ni
D. Cr
E. Cd

Choose the correct answer from the options given below:

[31-Jan-2024 Shift 1]

Options:

A.

B, C and E only

B.

A, B, C, D and E

C.

A, B, C and D only

D.

B, D and E only

Answer: A

Solution:

Mn, Ni and Cd metals used in battery industries.

Question9

One Faraday of electricity liberates $x \times 10^{-1}$ gram atom of copper from copper sulphate, x is_____

[31-Jan-2024 Shift 1]

Options:

Answer: 5

Solution:

$Cu^{2+} + 2e^- \longrightarrow Cu$

2 Faraday $\longrightarrow 1\ \text{mol Cu}$

1 Faraday $\longrightarrow 0.5\ \text{mol Cu}$ deposit

$0.5\ \text{mol} = 0.5\text{g atom} \quad = 5 \times 10^{-1}$

$x = 5$

Question10

The values of conductivity of some materials at $298.15 K^{-1} Sm^{-1}$ are 2.1×10^3, 1.0×10^{-16}, 1.2×10, 3.91, 1.5×10^{-2}, 1×10^{-7}, 1.0×10^3. The number of conductors among the materials is____

[31-Jan-2024 Shift 2]

Answer: 4

Solution:

Conductivity (Sm^{-1})

$$\left.\begin{array}{c} 2.1 \times 10^3 \\ 1.2 \times 10 \\ 3.91 \\ 1 \times 10^3 \end{array}\right\} \text{conductors at } 298.15\text{K}$$

1×10^{-16} Insulator at 298.15K

$$\left.\begin{array}{c} 1.5 \times 10^{-2} \\ 1 \times 10^{-7} \end{array}\right\} \text{Semiconductor at } 298.15\text{K}$$

Therefore number of conductors is 4 .

Question11

The potential for the given half cell at 298K is $\times 10^{-2}$V.

$2H^+_{(aq)} + 2e^- \longrightarrow H_2(g)$

$[H^+] = 1M,\ P_{H_2} = 2\ \text{atm}$

(Given: $2.303\ RT/F = 0.06V,\ \log 2 = 0.3$)

[1-Feb-2024 Shift 1]

Options:

Answer: 1

Solution:

$$E = E^{\circ}_{H^+/H_2} - \frac{0.06}{2}\log\frac{P_{H_2}}{[H^+]^2}$$

$$E = 0.00 - \frac{0.06}{2}\log\frac{2}{[1]^2}$$

$$E = -0.03 \times 0.3 = -0.9 \times 10^{-2}\,V$$

Question12

The amount of electricity in Coulomb required for the oxidation of 1mol of H_2O to O_2 is ______ × 10^5C.

[1-Feb-2024 Shift 2]

Answer: 2

Solution:

$$2H_2O \rightarrow O_2 + 4H^+ + 4e^-$$

$$\frac{W}{E} = \frac{Q}{96500}$$

$$\text{mole} \times \text{n-factor} = \frac{Q}{96500}$$

$$1 \times 2 = \frac{Q}{96500}$$

$$Q = 2 \times 96500C$$

$$= 1.93 \times 10^5 C$$

Question13

At 298K, a 1 litre solution containing 10 mmol of $Cr_2O_7^{2-}$ and 100 mmol of Cr^{3+} shows a pH of 3.0.

Given : $Cr_2O_7^{2-} \rightarrow Cr^{3+}$; $E^0 = 1.330V$ and $\frac{2.303\,RT}{F} = 0.059V$
The potential for the half cell reaction is $x \times 10^{-3}$ V. The value of x is____
[24-Jan-2023 Shift 1]

Answer: 917

Solution:

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$$

$$E = 1.33 - \frac{0.059}{6}\log\frac{(0.1)^2}{(10^{-2})(10^{-3})^{14}}$$

$$E = 1.33 - \frac{0.059}{6} \times 42 = 0.917$$

$$E = 917 \times 10^{-3}$$

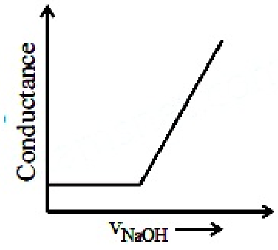
$$x = 917$$

Question14

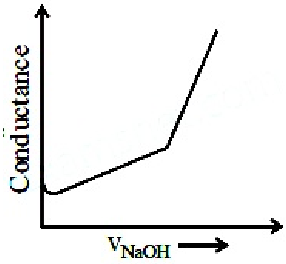
Choose the correct representation of conductometric titration of benzoic acid vs sodium hydroxide.
[24-Jan-2023 Shift 2]

Options:

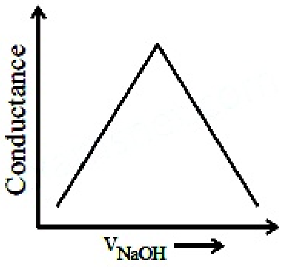
A.



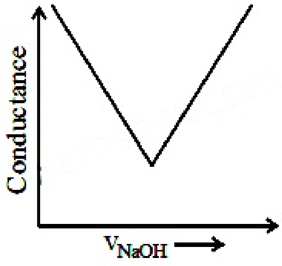
B.



C.



D.

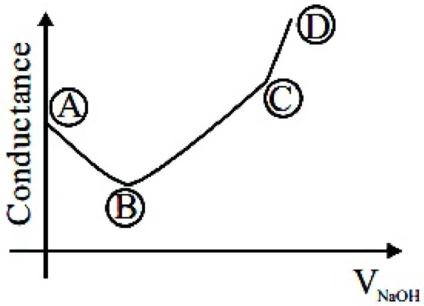


Answer: B

Solution:

Solution:

$C_6H_5COOH + NaOH \rightarrow C_6H_5COONa + H_2O$
(WA) (SB) (Salt)



(A) → (B) Free H^+ions are replaced by $Na^{\oplus}$ which decreases conductance.
(B) → (C) Un-dissociated benzoic acid reacts with NaOH and forms salt which increases ions & conductance increases.
(C) → (D) After equivalence point at (3), NaOH added further increases $Na^{\oplus}$&$OH^{\ominus}$ ions which further increases the conductance.

Question15

Consider the cell
Pt(s) | H_2(s)(latm)H^+(aq, $[H^+] = 1$)
Fe^{3+}(aq), Fe^{2+}(aq) | Pt(s)
Given: $E^{\circ}_{Fe^{3+}//Fe^{2+}} = 0.771V$ and $E^{\circ}_{H^+/\frac{1}{2}H_2} = 0V$, T = 298K

If the potential of the cell is 0.712V the ratio of concentration of Fe^{2+} to Fe^{2+} is ________ (Nearest integer)
[25-Jan-2023 Shift 1]

Answer: 10

Solution:

$$\frac{1}{2}H_2(g) + Fe^{3+}(aq.) \longrightarrow H^+(aq) + Fe^{2+}(aq.)$$

$$E = E^o - \frac{0.059}{1}\log\frac{[Fe^{2+}]}{[Fe^{3+}]}$$

$$\Rightarrow 0.712 = (0.771 - 0) - \frac{0.059}{1}\log\frac{[Fe^{2+}]}{[Fe^{3+}]}$$

$$\Rightarrow \log\frac{[Fe^{2+}]}{[Fe^{3+}]} = \frac{(0.771 - 0712)}{0.059} = 1$$

$$\Rightarrow \frac{[Fe^{2+}]}{[Fe^{3+}]} = 10$$

Question16

Pt(s) $H_2(g)$(1 bar) | $H^+(aq)$(1M)| | $M^{3+}(aq)$, $M^+(aq)$ | Pt(s)

The E_{cell} for the given cell is 0.1115V at 298K when $\frac{[M^+(aq)]}{[M^{3+}(aq)]} = 10^a$

The value of a is ________

Given : $E^{\theta}M^{3+} / M^+ = 0.2V$

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

[25-Jan-2023 Shift 2]

Answer: 3

Solution:

Overall reaction :-

$$H_{2(g)} + M_{(aq)}^{3+} \longrightarrow M_{(aq)}^{+} + 2H_{(aq)}^{+}$$

$$E_{Cell} = E_{Cathode}^{0} - E_{anode}^{0} - \frac{0.059}{2}\log\frac{[M^+] \times 1^2}{[M^{+3}]1}$$

$$0.1115 = 0.2 - \frac{0.059}{2}\log\frac{[M^+]}{[M^{+3}]}$$

$$3 = \log\frac{[M^+]}{[M^{+3}]}$$

$$\therefore a = 3$$

Question17

The standard electrode potential (M^{3+} / M^{2+}) for V, Cr, Mn& Co are −0.26V, −0.41V, +1.57V and +1.97V, respectively. The metal ions which can liberate H_2 from a dilute acid are

[29-Jan-2023 Shift 1]

Options:

A. V^{2+} and Mn^{2+}

B. Cr^{2+} and CO^{2+}

C. V^{2+} and Cr^{2+}

D. Mn^{2+} and Co^{2+}

Official Ans. by NTA (3)

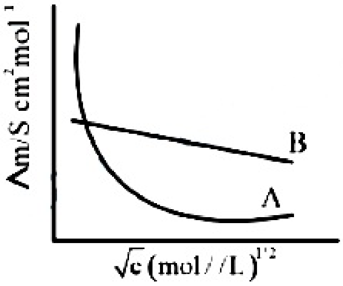
Answer: C

Solution:

Metal cation with (−) value of reduction potential (M^{+3} / M^{+2}) or with (+) value of oxidation potential (M^{+2} / M^{+3}) will liberate H_2
Therefore they will reduce H^+
i.eV eV^{+2} and Cr^{+2}

Question18

Following figure shows dependence of molar conductance of two electrolytes on concentration.
Λ_m^0 is the limiting molar conductivity



The number of Incorrect statement(s) from the following is ________
(A) Λm for electrolyte A is obtained by extrapolation
(B) For electrolyte B, vx Λm vs $\sqrt{c}$ graph is a straight line with intercept equal to Λ_m^0
(C) At infinite dilution, the value of degree of dissociation approach zero for electrolyte B.
(D) Λm for any electrolyte A or B can be calculated using λ^0 for individual ions.
[29-Jan-2023 Shift 1]

Answer: 2

Solution:

Solution:
Statement (A) and Statement (C) are incorrect

Question19

The equilibrium constant for the reaction
$Zn(s) + Sn^{2+}(aq) \rightleftharpoons Zn^{2+}(aq) + Sn(s)$ is 1×10^{20}

at 298K. The magnitude of standard electrode potential of Sn / Sn^{2+} if $E_{Zn^{2+}/Zn}{}^{0} = -0.76V$ is ______ $\times 10^{-2}$V. (Nearest integer)
[29-Jan-2023 Shift 2]

Answer: 17

Solution:

Given : $\frac{2.303\,RT}{F} = 0.059V$

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

$\Delta G^{\circ} = -2.303\,RT \log_{10} Keq$

$-nF(E_{cell}{}^{0}) = -2.303RT \log_{10} Keq$

$0.76 + E_{SQ^{2+}/Su}{}^{0} = \frac{0.059}{2}\log_{10}10^{20}$

$0.76 + E_{Sn^{2+}/Sn}{}^{0} = \frac{0.059 \times 20}{2}$

$E_{Sn^{2+}/Sa}{}^{0} = 0.59 - 0.76 = -0.17$

$E_{Su/Sn^{2+}}{}^{0} = 17 \times 10^{-2}V$

Ans. = 17

Question20

Consider the cell
$Pt_{(s)} \mid H_2(g, 1\,atm) \mid H^{+}(aq, 1M)$
$Fe^{3+}(aq), Fe^{2+}(aq) \mid Pt(s)$
When the potential of the cell is 0.712V at 298K, the ratio $[Fe^{2+}]/[Fe^{3+}]$ is ______.
(Nearest integer)
Given: $Fe^{3+} + e^{-} = Fe^{2+}$, $E^{\circ}Fe^{3+}, Fe^{2+} \mid Pt = 0.771$
$\frac{2.303\,RT}{F} = 0.06V$
[30-Jan-2023 Shift 1]

Answer: 10

Solution:

$Pt_{(s)} \mid H_2(g, 1\,latm) \mid H^{+}(aq, 1M) \mid Fe^{3+}(aq), Fe^{2+}(aq) \mid Pt(s)$

at anode $H_2 \rightarrow 2H^{+} + 2e^{-}$

At cathode $Fe_{aq}{}^{3+} + e^{-} \rightarrow Fe_{aq}{}^{2+}$

$$E^{\circ} = E_{H_2 \mid H^{+}}{}^{\circ} + E_{Fe^{\circ +} \mid Fe^{2+}} = 0 \cdot 771V$$

$$E = E^{\circ} - \frac{0 \cdot 06}{1} \log \frac{Fe^{2+}}{Fe^{3+}}$$

$$0 \cdot 712 = (0 + 0 \cdot 771) - \frac{0 \cdot 06}{1} \log \frac{Fe^{2+}}{Fe^{3+}}$$

$$\log \frac{Fe^{2+}}{Fe^{3+}} = \frac{0 \cdot 059}{0 \cdot 06} \approx 1$$

$$\frac{Fe^{2+}}{Fe^{3+}} = 10$$

Question21

The electrode potential of the following half cell at
298K .
$X \mid X^{2+}(0.001M) \mid Y^{2+}(0.01M) \mid Y$
is ______ $\times 10^{-2}V$ (Nearest integer).
Given: $E_{x^{2+} \mid x}{}^{0} = -2.36V$
$E_{Y^{2+}{}_{Y}}{}^{0} = +0.36V$
$\frac{2.303\,RT}{F} = 0.06V$
[30-Jan-2023 Shift 2]

Answer: 275

Solution:

$$X + Y^{2+} \rightarrow Y + X^{2+}$$
$$E_{Cell}{}^{0} = 0.36 - (-2.36) = 2.72V$$
$$E_{\ Cell} = 2.72 - \frac{0.06}{2} \log \frac{0.001}{0.01}$$
$$= 2.72 + 0.03 = 2.75V$$
$$= 275 \times 10^{-2}V$$

Question22

Which one of the following statements is correct for electrolysis of brine solution?
[31-Jan-2023 Shift 1]

Options:

A. Cl_2 is formed at cathode

B. O_2 is formed at cathode

C. H_2 is formed at anode

D. $OH^-$is formed at cathode

Answer: D

Solution:

Solution:
Electrolysis of brine solution
$NaCl\ (aq.) \rightarrow Na_{(aq)}^{+} + Cl_{(aq)}^{+}$
At anode : $2Cl_{(aq.)}^{+} \rightarrow \underset{\text{Major}}{Cl_2(g)} + 2e^-$
$2H_2O_{(\ell)} \rightarrow \underset{\text{Minor}}{O_{2(g)}} + 4H_{(aq)}^{+} + 4e^-$
At Cathode : $2H_2O_{(\ell)} + 2e^- \rightarrow H_{2(g)} \uparrow + 2OH_{(aq)}^{-}$
$2Na^+ + 2OH^- \rightarrow 2\,NaOH$

Question23

The resistivity of a 0.8M solution of an electrolyte is $5 \times 10^{-3}\Omega\,cm$. Its molar conductivity is _______ $\times 10^4 \Omega^{-1} cm^2 mol^{-1}$. (Nearest integer)
[31-Jan-2023 Shift 2]

Answer: 25

Solution:

$$\Lambda_m = \frac{\kappa \times 1000}{M}$$
$$\Lambda_m = \frac{1}{\rho} \times \frac{1000}{M}$$
$$\frac{1}{5 \times 10^{-3}} \times \frac{1000}{0.8}$$
Ans. $25 \times 10^4 \Omega^{-1} cm^{-2} mol^{-1}$

Question24

At what pH, given half cell $MnO_4^-(0.1M) \mid Mn^{2+}$ (0.001 M) will have electrode potential of 1.282 V ? _______ (Nearest Integer)
Given $E^{o}_{MnO_4^- / Mn^{2+}} = 1.54V$, $\frac{2.303RT}{F} = 0.059V$
[1-Feb-2023 Shift 1]

Answer: 3

Solution:

$$MnO_4^- + 8H^+ + 5e^- \rightleftharpoons Mn^{2+} + 4H_2O$$

$$E = E^\circ - \frac{0.059}{5}\log\frac{[Mn^{2+}]}{[MnO_4^-][H^+]^8}$$

$$1.282 = 1.54 - \frac{0.059}{5}\log\frac{10^{-3}}{10^{-1}\times[H^+]^8}$$

$$\frac{0.258\times5}{0.059} = \log\frac{10^{-2}}{[H^+]^8}$$

$$\Rightarrow\ 21.86 = -2 + 8\,pH$$

$$\therefore\ pH = 2.98$$

$$\simeq 3$$

Question25

1×10^{-5}M AgNO 3 is added to 1L of saturated solution of AgBr. The conductivity of this solution at 298K is _____ $\times10^{-8}Sm^{-1}$.

[Given : $K_{sp}(AgBr) = 4.9\times10^{-13}$ at 298K

$\lambda_{Ag^+}{}^0 = 6\times10^{-3}Sm^2mol^{-1}$

$\lambda_{Br^-}{}^0 = 8\times10^{-3}Sm^2mol^{-1}$

$\lambda_{NO_3^-}{}^0 = 7\times10^{-3}Sm^2mol^{-1}$]

[1-Feb-2023 Shift 2]

Answer: 14

Solution:

$[Ag^+] = 10^{-5}$

$[NO_3^-] = 10^{-5}$

$$[Br^-] = \frac{Ksp}{[Ag^+]} = 4.9\times10^{-8}$$

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

For Ag^+

$$6\times10^{-3} = \frac{K_{Ag^+}}{1000\times10^{-5}}$$

$K_{Ag+} = 6\times10^{-5}$

$\Rightarrow 6000\times10^{-8}$

for Br^-

$$8\times10^{-3} = \frac{K_{Br}}{1000\times4.9\times10^{-8}}$$

$K_{Br-} = 39.2\times10^{-8}$

for NO_3^-

$7 \times 10^{-3} = \dfrac{K_{NO_3^-}}{1000 \times 10^{-5}}$

$K_{NO_3^-} = 7 \times 10^{-5}$

$= 7000 \times 10^{-8}$

Conductivity of solution

$\Rightarrow (6000 + 7000 + 39.2) \times 10^{-8}$

$\Rightarrow 13039.2 \times 10^{-8} Sm^{-1}$

Question26

The standard electrode potential of M^+ / M in aqueous solution does not depend on
[6-Apr-2023 shift 1]

Options:

A. Ionisation of a solid metal atom

B. Sublimation of a solid metal

C. Ionisation of a gaseous metal atom

D. Hydration of a gaseous metal ion

Answer: A

Solution:

Solution:

Question27

The product, which is not obtained during the electrolysis of brine solution is
[6-Apr-2023 shift 2]

Options:

A. NaOH

B. Cl_2

C. H_2

D. HCl

Answer: D

Solution:

Brine is aq. Solution of NaCl

$NaCl_{(aq)} \rightarrow Na^+ + Cl^-$

Cathode reaction

$2H_2O + 2e^- \xrightarrow{\text{reduction}} H_{2(g)} + 2OH^-$

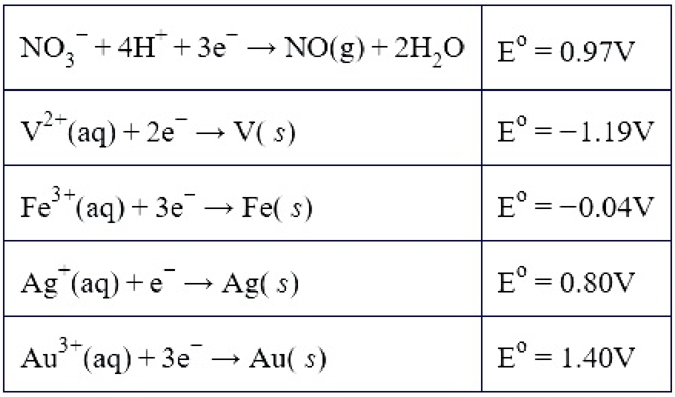
Anode reaction

$2Cl^- \xrightarrow{\text{oxidation}} Cl_{2(g)} + 2e^-$

So HCl will not form during electrolysis.

Question28

The standard reduction potential at 298K for the following half cells are given below :-



$NO_3^- + 4H^+ + 3e^- \rightarrow NO(g) + 2H_2O$	$E^o = 0.97V$
$V^{2+}(aq) + 2e^- \rightarrow V(s)$	$E^o = -1.19V$
$Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$	$E^o = -0.04V$
$Ag^+(aq) + e^- \rightarrow Ag(s)$	$E^o = 0.80V$
$Au^{3+}(aq) + 3e^- \rightarrow Au(s)$	$E^o = 1.40V$

The number of metal(s) which will be oxidized by NO_3^- in aqueous solution is
[6-Apr-2023 shift 2]

Answer: 3

Solution:

Solution:
Metal having lower SRP than 0.97V will be oxidised by NO_3^-.

Question29

The reaction
$\frac{1}{2}H_2(g) + Ag(Cl)(s) \rightleftharpoons H^+(aq) + Cl^-(aq) + Ag(s)$
occurs in which of the given galvanic cell.
[8-Apr-2023 shift 1]

Options:

A. Pt | H_2 (g) |HCl(sol ^{n})|$AgNO_3$(sol ^{n}) | Ag

B. Pt |H_2(g)| HCl(sol ^{n}) | AgCl(s) | Ag

C. $\mathrm{Pt}\,|\mathrm{H_2(g)}|\,\mathrm{KCl(\ sol^{\ n})}\ |\ \mathrm{AgCl(\ s)}\ |\ \mathrm{Ag}$

D. $\mathrm{Ag}\ |\ \mathrm{AgCl(\ s)}\ |\ \mathrm{KCl(\ sol^{\ n})}\ |\mathrm{AgNO_3}|\,\mathrm{Ag}$

Answer: B

Solution:

Solution:
Anode $\rightarrow \mathrm{H_2} \rightarrow 2\mathrm{H^+} + 2\mathrm{e^-}$
Cathode $\rightarrow \mathrm{AgCl} + \mathrm{e^-} \rightarrow \mathrm{Ag} + \mathrm{Cl^-}$

Question30

The specific conductance of 0.0025M acetic acid is $5 \times 10^{-5}\mathrm{Scm^{-1}}$ at a certain temperature. The dissociation constant of acetic acid is ________ $\times 10^{-7}$. (Nearest integer)
Consider limiting molar conductivity of $\mathrm{CH_3COOH}$ as $400\mathrm{Scm^2mol^{-1}}$.
[10-Apr-2023 shift 2]

Answer: 66

Solution:

Solution:

$$\Lambda_m = \frac{k}{C} \times 1000$$

Given $k = 5 \times 10^{-5}\mathrm{Scm^{-1}}$
$C = 0.0025\mathrm{M}$

$$\Lambda_m = \frac{5 \times 10^{-5} \times 10^3}{0.0025} = \frac{5 \times 10^{-2}}{2.5 \times 10^{-3}}$$

$$= 20\mathrm{Scm^2mol^{-1}}$$

$$\alpha = \frac{20}{400} = \frac{1}{20}$$

$$K_a = \frac{C\alpha^2}{1-\alpha} = \frac{0.0025 \times \frac{1}{20} \times \frac{1}{20}}{\frac{19}{20}}$$

$$= \frac{0.0025}{19 \times 20} = 6.6 \times 10^{-6}$$

$$= 66 \times 10^{-7}$$

Question31

In an electrochemical reaction of lead, at standard temperature, if $E_{(pb^{2+}/pb)}^{\circ}$ = m volt and $E_{(pb^{+}/Pb_b)}^{\circ}$ = n volt, then the value of $E_{(Pb^{2+}/Pb^{+})}^{\circ}$ is given by m − xn. The value of x is ________ (Nearest integer)

[11-Apr-2023 shift 1]

Answer: 2

Solution:

$$\begin{array}{lll} Pb^{2+} + 2e^- \rightarrow Pb & E^\circ = m & \Delta G_1^\circ = -2Fm \\ Pb^{4+} + 4e^- \rightarrow Pb & E^\circ = n & \Delta G_2^\circ = -4Fn \\ \hline Pb^{2+} \rightarrow Pb^{4+} + 2e^- & \Delta G_3^\circ = \Delta G_1^\circ - \Delta G_2^\circ & \\ \hline & -2FE^\circ = -2Fm + 4Fn & \\ & \boxed{E^\circ = m - 2n} & \\ & \boxed{x = 2} & \end{array}$$

Question32

The number of correct statements from the following is _______
A. E_{cell} is an intensive parameter
B. A negative E⊖ means that the redox couple is a stronger reducing agent than the H^+ / H_2 couple.
C. The amount of electricity required for oxidation or reduction depends on the stoichiometry of the electrode reaction.
D. The amount of chemical reaction which occurs at any electrode during electrolysis by a current is proportional to the quantity of electricity passed through the electrolyte.
[11-Apr-2023 shift 2]

Answer: 4

Solution:

Solution:
Given statements A, B, C and D are correct.

Question33

For lead storage battery pick the correct statements
A. During charging of battery, $PbSO_4$ on anode is converted into PbO_2

B. During charging of battery, $PbSO_4$ on cathode is converted into PbO_2
C. Lead storage battery consists of grid of lead packed with PbO_2 as anode
D. Lead storage battery has ~38% solution of sulphuric acid as an electrolyte Choose the correct answer from the options given below:
[12-Apr-2023 shift 1]

Options:

A. B, D only

B. B, C only

C. B, C, D only

D. A,B,D only

Answer: A

Solution:

Solution:
Lead storage battery consists of lead anode and a grid of lead packed with lead oxide (PbO_2) as cathode, a 38% solution of H_2SO_4 is used as an electrolyte.
On charging the battery the reaction is reversed and $PbSO_4(s)$ on anode and cathode is converted into Pb and PbO_2 respectively.

Question34

A metal surface of $100 cm^2$ area has to be coated with nickel layer of thickness 0.001 mm. A current of 2A was passed through a solution of $Ni(NO_3)_2$ for ' x ' seconds to coat the desired layer. The value of x is (Nearest integer)
(ρ_{Ni}(.. density of Nickel) is $10 gmL^{-1}$, Molar mass of Nickel is ________ $60 gmol^{-1}$ $F = 96500 Cmol^{-1}$)
[13-Apr-2023 shift 1]

Answer: 161

Solution:

$$\text{Volume of nickel required} = 100 \times 0.001 \times 10^{-3} \times 100$$
$$= 0.01 cm^3$$
$$\text{Mass of Nickel required} = 0.01 \times 10$$
$$= 0.1 gm$$
$$\text{Moles} = \frac{0.1}{60} = \frac{1}{600} \text{mol}$$
$$Ni^{2+} + 2e^- \rightarrow Ni(s)$$

for coating of $1\,\text{mol Ni}$, charge required $= 2 \times 96500\text{C}$

for coating of $\frac{1}{600}$ mol, charge required $= 2 \times 96500 \times \frac{1}{600}\text{C}$

$= \frac{965}{3}\text{C}$

$I = \frac{q}{t}$

$t = \frac{965/3}{2} = 160.83 \text{ sec} \approx 161$

Question35

At 298K, the standard reduction potential for Cu^{2+} / Cu electrode is 0.034V.

Given : K_{sp} $Cu(OH)_2 = 1 \times 10^{-20}$

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

The reduction potential at pH = 14 for the above couple is $(-)x \times 10^{-2}V$. The value of x is _______.

[13-Apr-2023 shift 2]

Answer: 25

Solution:

$Cu(OH)_2(s)\ \ Cu^{2+}(aq) + 2OH^-(aq)$

$Ksp = [Cu^{2+}][OH^-]^2$

$pH = 14; pOH = 0; [OH^-] = 1M$

$\therefore [Cu^{2+}] = \frac{Ksp}{[1]^2} = 10^{-20}M$

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

$E = E^\circ - \frac{0.059}{2}\log_{10}\frac{1}{[Cu^{2+}]}$

$= 0.34 - \frac{0.059}{2}\log_{10}\frac{1}{10^{-20}}$

$= -0.25 = -25 \times 10^{-2}$

Question36

The number of correct statements from the following is _______.

(A) Conductivity always decreases with decrease in concentration for both strong and weak electrolysis.

(B) The number of ions per unit volume that carry current in a solution increases on dilution.

(C) Molar conductivity increases with decrease in concentration.

(D) The variation in molar conductivity is different for strong and weak electrolysis.

(E) For weak electrolysis, the change in molar conductivity with dilution is due to decrease in degree of dissociation.
[15-Apr-2023 shift 1]

Answer: 3

Solution:

Solution:
(A) Conductivity decreases with dilution for strong electrolyte as well as weak electrolyte.
(B) On dilution, The number of ions per unit volume that carry current in a solution decreases.
(C) Molar conductivity increases with dilution.
(D) Molar conductivity of strong electrolyte follows DHO equation but it is not applicable for weak electrolyte.
(E) On dilution degree of dissociation of weak electrolyte increases.
So answer is (A), (C) & (D)

Question37

The cell potential for the following cell
Pt | H_2(g) | H^+(aq) | Cu^{2+}(0.01M) | Cu(s)
is 0.576V at 298K. The pH of the solution is _____(Nearest integer)
(Given : $E^{o}_{Cu^{2+}/Cu}$ = 0.34V and $\frac{2.303RT}{F}$ = 0.06V)
[24-Jun-2022-Shift-1]

Answer: 5

Solution:

$$E_{cell} = E_{cell}^{0} - \frac{0.06}{2}\log\frac{\left[H^{\oplus}\right]^2}{[Cu^{+2}]}$$
$$0.576 = 0.34 - 0.03\log\frac{\left[H^{\oplus}\right]^2}{[0.01]}$$
$$0.576 - 0.34 = -0.03\log\left[H^{\oplus}\right]^2 + 0.03\log(0.01)$$
$$= 0.06\,pH - 0.06$$
$$pH \simeq 4.93 \simeq 5$$

Question38

The resistance of a conductivity cell containing 0.01M KCl solution at 298K is 1750Ω. If the conductivity of 0.01M KCl solution at 298K is 0.152×10^{-3} Scm^{-1}, then the cell constant of the conductivity cell is

_____$\times 10^{-3} cm^{-1}$
[24-Jun-2022-Shift-2]

Answer: 266

Solution:

Molarity of KCl solution $= 0.1M$
Resistance $= 1750\,ohm$
Conductivity $= 0.152 \times 10^{-3} Scm^{-1}$
Conductivity $= \dfrac{\text{Cell constant}}{\text{Resistance}}$
∴ Cell constant $= 0.152 \times 10^{-3} \times 1750$
$= 266 \times 10^{-3} cm^{-1}$

Question39

In a cell, the following reactions take place

$Fe^{2+} \rightarrow Fe^{3+} + e^- \quad E^o_{Fe^{3+}/Fe^{2+}} = 0.77V$

$2I^- \rightarrow I_2 + 2e^- \quad E^o_{I_2/I^-} = 0.54V$

The standard electrode potential for the spontaneous reaction in the cell is x × 10^{-2}V298K . The value of x is ___(Nearest Integer)
[25-Jun-2022-Shift-1]

Answer: 23

Solution:

$$\underset{\text{Cathode}}{Fe^{+3}} + \underset{\text{anode}}{I^-} \longrightarrow I_2 + Fe^{+2}$$
$E_{cell}^{\circ} = E_{cathode}^{\circ} - E_{anode}^{\circ}$
$= 0.77 - 0.54$
$= 0.23V$
$= 23 \times 10^{-2}V$

Question40

The correct order of reduction potentials of the following pairs is
A. Cl_2 / Cl^-
B. $1^2 / 1^-$

C. Ag^+ / Ag
D. Na^+ / Na
E. Li^+ / Li
Choose the correct answer from the options given below.
[25-Jun-2022-Shift-2]

Options:

A. A > C > B > D > E

B. A > B > C > D > E

C. A > C > B > E > D

D. A > B > C > E > D

Answer: A

Solution:

$E_{C_2/CI}^{\circ} = +1.36V$
$E_{I_2/I^-}^{0} = +0.54V$
$E_{Ag^+/Ag}^{\circ} = +0.80V$
$E_{Na^+/Na}^{\circ} = -2.71V$
$E_{L^+/Li} = -3.05V$

Question41

A solution of $Fe_2(SO_4)_3$ is electrolyzed for ' x ' min with a current of 1.5A to deposit 0.3482g of Fe. The value of x is [nearest integer]
Given : $1F = 96500 Cmol^{-1}$
Atomic mass of Fe = $56 gmol^{-1}$
[25-Jun-2022-Shift-2]

Answer: 20

Solution:

$Fe^{3+} + 3e^- \longrightarrow Fe$
3F ⟶ 1 mole Fe is deposited
For 56g ⟶ 3 × 96500 (required charge)
For lg $\longrightarrow \dfrac{3 \times 96500}{56}$ (required charge)
For 0.3482g $\longrightarrow \dfrac{3 \times 96500}{56} \times 0.3482$
$= 1800.06$
$Q = it$
$1800.06 = 1.5t$
$t = 20 \text{ min}$

Question42

The $\left(\frac{\partial E}{\partial T}\right)_P$ of different types of half cells are as follows:

A	B	C	D
1×10^{-4}	2×10^{-4}	0.1×10^{-4}	0.2×10^{-4}

(Where E is the electromotive force)
Which of the above half cells would be preferred to be used as reference electrode?
[26-Jun-2022-Shift-1]

Options:

A. A

B. B

C. C

D. D

Answer: C

Solution:

Solution:
A cell with less variation in EMF with temperature is preferred as a reference electrode because it can be used for a wider range of temperatures without much derivation from standard value so a cell with less $\left(\frac{\partial E}{\partial T}\right)_P$ is preferred.

Question43

$Cu(s) + Sn^{2+}(0.001M) \rightarrow Cu^{2+}(0.01M) + Sn(s)$
The Gibbs free energy change for the above reaction at 298K is $x \times 10^{-1}\,kJ\,mol^{-1}$. The value of x is ______ [nearest integer]
[Given : $E^{\ominus}_{Cu^{2+}/Cu} = 0.34V$; $E^{\ominus}_{Sn^{2+}/Sn} = -0.14V$; $F = 96500 C mol^{-1}$]
[26-Jun-2022-Shift-2]

Answer: 983

Solution:

$Cu_{(s)} + Sn^{2+}(0.001M) \rightarrow Cu^{2+}(0.01M) + Sn_{(s)}$
$E_{cell}{}^{\circ} = E_{cathode}{}^{\circ} - E_{anode}{}^{\circ}$
$= -0.14 - (0.34)$
$= -0.48V$
$E_{cell} = E_{cell}{}^{\circ} - \frac{0.059}{2} \log \frac{[Cu^{2+}]}{[Sn^{2+}]}$
$= -0.48 - \frac{0.059}{2} \log \frac{0.01}{0.001}$
$= -0.509$
$\Delta G = -nFE_{cell}$
$= -2 \times 96500 \times (-0.5095)$
$= 98333.5J / mol$
$= 98.335\,kJ / mol$
$= 983.35 \times 10^{-1}\,kJ / mol$

Question44

In 3d series, the metal having the highest M^{2+} / M standard electrode potential is
[27-Jun-2022-Shift-2]

Options:

A. Cr

B. Fe

C. Cu

D. Zn

Answer: C

Solution:

$Cr^{+2} / Cr \rightarrow -0.90V$
$Fe^{+2} / Fe \rightarrow -0.44V$
$Cu^{+2} / Cu \rightarrow +0.34V$
$Zn^{+2} / Zn \rightarrow -0.76V$
So Ans. is Cu^{+2} / Cu

Question45

For the reaction taking place in the cell :
$Pt(s) \mid H_2(g) \mid H^+(aq) \mid Ag^+(aq) \mid Ag(s)\ E_{cell}{}^{o} = +0.5332V$.
The value of $\Delta_f G^{\circ}$ is ____kJ mol^{-1}. (in nearestinteger)
[27-Jun-2022-Shift-2]

Answer: 103

Solution:

$H_2 \left| H^{+} \right| \left| Ag^{+} \right| Ag$

Anode Cathode

At anode, oxidation occur

$H_2 \rightarrow 2H^{+} + 2e^{-}$

\# At cathode, reduction occur

$2Ag^{+} + 2e^{-} \rightarrow 2Ag$

Adding equation (1) and (2), we get $n = 2$, where $n =$ cancelled out electron

Now,

$\Delta G^{\circ} = -nFE_{cell}^{o}$

$= -2 \times 96500 \times 0.5332$

$= -102907.6$

$= -102.9\,kJ / mol$

$= -103\,kJ / mol$

Question46

The solubility product of a sparingly soluble salt A_2X_3 is 1.1×10^{-23}. If specific conductance of the solution is $3 \times 10^{-5} Sm^{-}$ 1 , the limiting molar conductivity of the solution is $x \times 10^{-3} Sm^2 mol^{-1}$. The value of x is____

[28-Jun-2022-Shift-1]

Answer: 3

Solution:

$A_2X_3 \rightleftharpoons \underset{2S}{2A} + \underset{3S}{3X}$

$K_{sp} = (2s)^2(3s)^3 = 1.1 \times 10^{-23}$

$S \approx 10^{-5}$

For sparingly soluble salts

$\Lambda_m = \Lambda_m^{0}$

$\Lambda_m = \dfrac{k}{S \times 10^3}$

$= \dfrac{3 \times 10^{-5}}{10^{-5}} \times 10^{-3}$

$= 3 \times 10^{-3} Sm^2 mol^{-1}$

Question47

The quantity of electricity in Faraday needed to reduce 1 mol of $Cr_2O_7^{2-}$ to Cr^{3+} is ______

[28-Jun-2022-Shift-1]

Answer: 6

Solution:

$Cr_2O_7^{-2} + 6e^- \longrightarrow 2Cr^{+3}$
1 mol 6 mol
⇒ number of faradays = moles of electrons = 6

Question48

For the given reactions
$Sn^{2+} + 2e^- \rightarrow Sn$
$Sn^{4+} + 4e^- \rightarrow Sn$
the electrode potentials are ;$E_{Sn^{2+}/Sn}^{o} = -0.140V$ and $E_{Sn^{4+}/Sn}^{o} = -0.010V$. The magnitude of standard electrode potential for Sn^{4+}/Sn^{2+} i.e. $E_{Sn^{4+}/Sn^{2+}}^{o}$ is _____$\times 10^{-2}$V. (Nearest integer)
[28-Jun-2022-Shift-2]

Answer: 16

Solution:

$Sn \longrightarrow Sn^{2+} + 2e^- \quad E_1^0 = 0.140V$
$Sn^{4+} + 4e^- \longrightarrow Sn \quad E_2^0 = 0.010V$
$Sn^{4+} + 2e^- \longrightarrow Sn^{2+} \quad E_{cell}^{0}$
$E_{cell}^{O} = \frac{n_2E_2^o + n_1E_1^0}{n} = \frac{4(0.010) + 2(0.140)}{2}$
$E_{cell}^{0} = 0.16V = 16 \times 10^{-2}V$

Question49

A dilute solution of sulphuric acid is electrolysed using a current of 0.10A for 2 hours to produce hydrogen and oxygen gas. The total volume of gases produced a STP is_____cm^3. (Nearest integer)
[Given : Faraday constant F = $96500Cmol^{-1}$ at STP, molar volume of an ideal gas is $22.7Lmol^{-1}$]
[29-Jun-2022-Shift-1]

Answer: 127

Solution:

Solution:
At anode
$2H_2O \rightarrow O_2(g) + 4H^+ + 4e^-$
At cathode
$2H^+ + 2e^- \rightarrow H_2(g)$

Now number of gm eq. $= \frac{i \times t}{96500}$

$= \frac{0.1 \times 2 \times 60 \times 60}{96500}$

$= 0.00746$

$V_{O_2} = \frac{0.00746}{4} \times 22.7 = 0.0423$

$V_{H_2} = \frac{0.00746}{2} \times 22.7 = 0.0846$

$V_{\text{Total}} \approx 127\,\text{ml}$ or cc

Question50

The cell potential for the given cell at 298K
$Pt \mid H_2(g, 1\,\text{bar}) \mid H^+(aq) \mid Cu^{2+}(aq) \mid Cu(s)$
is 0.31V. The pH of the acidic solution is found to be 3 , whereas the concentration of Cu^{2+} is 10^{-x}M. The value of x is_____
(Given : $E^{\Theta}_{Cu^{2+}/Cu} = 0.34V$ and $\frac{2.303RT}{F} = 0.06V$)
[29-Jun-2022-Shift-2]

Answer: 7

Solution:

$Q = \frac{[H^+]^2}{[Cu^2]pH_2} = \frac{10^{-6}}{C}$ $pH_2 = 1$

$E = E_{\text{cell}}{}^{\circ} - \frac{0.06}{n} \log Q$

$0.31 = 0.34 - \frac{0.06}{2} \log \frac{10^{-6}}{C}$

$\log \frac{10^{-6}}{C} = 1$

$C = 10^{-7}M$
$x = 7$

Question51

The cell potential for $Zn \mid Zn^{2+}(aq) \mid\mid Sn^{x+} \mid Sn$ is $0.801V$ at $298K$. The reaction quotient for the above reaction is 10^{-2}. The number of electrons involved in the given electrochemical cell reaction is (Given : $E^{o}_{Zn^{2+} \mid Zn} = -0.763V$, $E^{o}_{Sn^{x+} \mid Sn} = +0.008V$ and $\frac{2.303RT}{F} = 0.06V$)
[25-Jul-2022-Shift-1]

Answer: 4

Solution:

$A : Zn \rightarrow Zn^{2+} + 2e^{-}$
$C : Sn^{+x} + xe^{-} \rightarrow Sn$
$E_{Cell}^{\circ} = E_{Zn \mid Zn^{2+}}^{\circ} + E_{Sn^{+x} \mid Sn}^{\circ}$
$\Rightarrow 0.763 + 0.008 = 0.771V$
From the Nernst equation,
$$E_{Cell} = E_{Cell}^{\circ} \frac{-2.303\,RT}{nF} \log Q$$
$$0.801 = 0.771 - \frac{0.06}{n} \log 10^{-2}$$
$$0.03 = \frac{0.06}{n} \times 2$$
$n = 4$

Question52

The molar conductivity of a conductivity cell filled with 10 moles of $20mL\, NaCl$ solution is Λ_{m1} and that of 20 moles another identical cell heaving $80mL\, NaCl$ solution is Λ_{m2}. The conductivities exhibited by these two cells are same. The relationship between Λ_{m2} and Λ_{m1} is
[25-Jul-2022-Shift-2]

Options:

A. $\Lambda_{m2} = 2\Lambda_{m1}$

B. $\Lambda_{m2} = \Lambda_{m1} / 2$

C. $\Lambda_{m.2} = \Lambda_{m1}$

D. $\Lambda_{m2} = 4\Lambda_{m1}$

Answer: A

Solution:

$$\Lambda_{m_1} = \frac{k_1 \times 1000}{M_1} = \frac{k \times 1000}{\frac{10}{0.02}}$$

$$\Lambda_{m_2} = \frac{k_2 \times 1000}{\frac{20}{0.08}}$$

It is given that $k_1 = k_2$

$$k_1 = \frac{\Lambda_{m_1}}{2} \quad k_2 = \frac{\Lambda_{m_2}}{4}$$

Applying the given condition on conductivity.

$$\frac{\Lambda_{m_1}}{2} = \frac{\Lambda_{m_2}}{4}$$

$$\Lambda_{m_2} = 2\Lambda_{m_1}$$

Question53

The spin-only magnetic moment value of M^{3+} ion (in gaseous state) from the pairs Cr^{3+} / Cr^{2+}, Mn^{3+} / Mn^{2+}, Fe^{3+} / Fe^{2+} and Co^{3+} / Co^{2+} that has negative standard electrode potential, is _____ B.M. [Nearest integer]

[25-Jul-2022-Shift-2]

Answer: 4

Solution:

Solution:

Among the pairs given, Cr^{3+} / Cr^{2+} has negative reduction potential which is $-0.41V$.

$Cr(III) \Rightarrow d^3$

Number of unpaired electrons $= 3$

$\mu = \sqrt{3(3+2)} = \sqrt{15} \simeq 4$ B.M.

Question54

The amount of charge in F (Faraday) required to obtain one mole of iron from Fe_3O_4 is ______. (Nearest Integer)

[26-Jul-2022-Shift-1]

Answer: 3

Solution:

For Fe_3O_4,

$x = \frac{+8}{3}$

where x is oxidation state of Fe.

$Fe_3O_4 + 8H^+ + 8e^- \longrightarrow 3\,Fe + 4H_2O$

Charge required $= \frac{8}{3} \times F = \frac{8F}{3} \simeq 3F$

Question55

Given below are two statements :
Statement I : For KI, molar conductivity increases steeply with dilution
Statement II : For carbonic acid, molar conductivity increases slowly with dilution
In the light of the above statements, choose the correct answer from the options given below:
[27-Jul-2022-Shift-2]

Options:

A. Both Statement I and Statement II are true

B. Both Statement I and Statement II are false

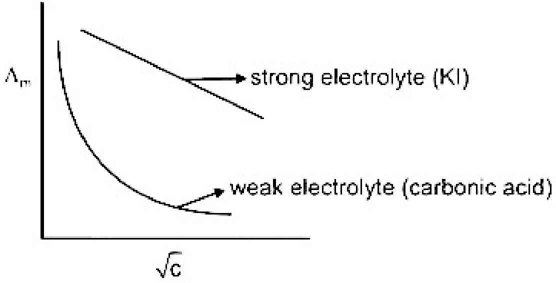
C. Statement I is true but Statement II is false

D. Statement I is false but Statement II is true

Answer: B

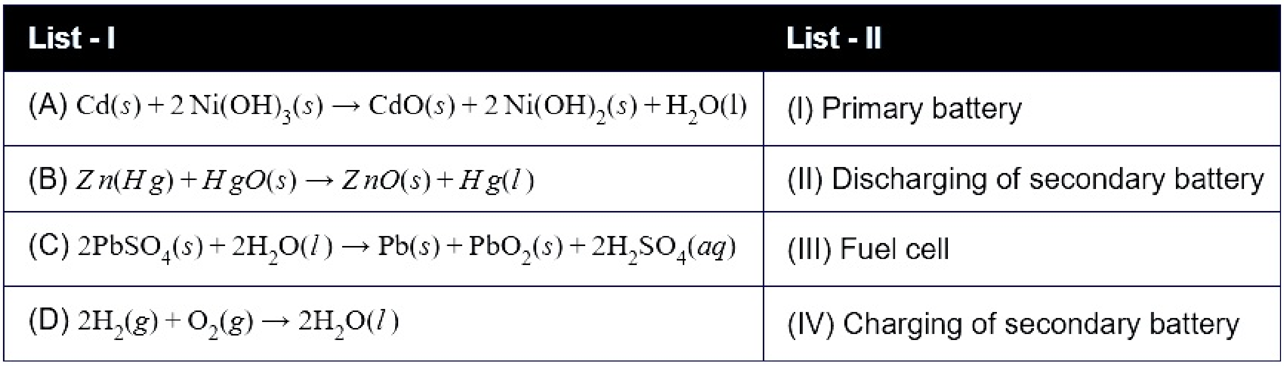
Solution:

Solution:
Statement I: KI is strong electrolyte thus almost constant on dilution.
Statement II: In weak electrolyte it increases, sharply.



Question56

Match List - I with List - II.



List - I	List - II
(A) $Cd(s)+2\,Ni(OH)_3(s) \rightarrow CdO(s)+2\,Ni(OH)_2(s)+H_2O(l)$	(I) Primary battery
(B) $Zn(Hg)+HgO(s) \rightarrow ZnO(s)+Hg(l)$	(II) Discharging of secondary battery
(C) $2PbSO_4(s)+2H_2O(l) \rightarrow Pb(s)+PbO_2(s)+2H_2SO_4(aq)$	(III) Fuel cell
(D) $2H_2(g)+O_2(g) \rightarrow 2H_2O(l)$	(IV) Charging of secondary battery

Choose the correct answer from the options given below:
[28-Jul-2022-Shift-1]

Options:

A. (A) – (I), (B) – (II), (C) – (II), (D) – (IV)

B. (A) – (IV), (B) – (I), (C) – (I), (D) – (II)

C. (A) – (II), (B) – (I), (C) – (IV), (D) – (III)

D. (A) – (II) ,(B) – (I), (C) – (II), (D) – (IV)

Answer: C

Solution:

Solution:
(a) $Cd(s) + 2\,Ni(OH)_3(s) \rightarrow CdO(s) + 2\,Ni(OH)_2(s) + H_2O(l)$ Discharge of secondary Battery
(b) $Zn(Hg) + HgO(s) \rightarrow ZnO(s) + Hg(l)$ (Primary Battery Mercury cell)
(c) $2PbSO_4(s) + 2H_2O(l) \rightarrow Pb(s) + PbO_2(s) + 2H_2SO_4(aq)$ (Charging of secondary Battery)
(d) $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ (Fuel cell)

Question57

Resistance of a conductivity cell (cell constant $129\,m^{-1}$) filled with 74.5 ppm solution of KCl is 100Ω (labelled as solution 1). When the same cell is filled with KCl solution of 149 ppm, the resistance is 50Ω (labelled as solution 2). The ratio of molar conductivity of solution 1 and solution 2 is i.e. $\frac{\Lambda_1}{\Lambda_2} = x \times 10^{-3}$. The value of x is ______.(Nearest integer)
Given, molar mass of KCl is $74.5\,g\,mol^{-1}$.
[29-Jul-2022-Shift-1]

Answer: 1000

Solution:

$\frac{l}{A} = 129\mathrm{m}^{-1}$
KCl solution 1 ⇒ 74.5 ppm, $R_1 = 100\Omega$
KCl solution 2 ⇒ 149 ppm, $R_2 = 50\Omega$
Here, $\frac{ppm_1}{ppm_2} = \frac{M_1}{M_2} = \left(\frac{w_{1/M_0}}{V} \times \frac{V}{w_{2/M_0}} \right)$

$$\frac{\Lambda_1}{\Lambda_2} = \frac{k_1 \times \frac{1000}{M_1}}{k_2 \times \frac{1000}{M_2}}$$
$$= \frac{k_1}{k_2} \times \frac{M_1}{M_2}$$
$$= \frac{50}{100} \times 2$$
$$= \frac{\Lambda_1}{\Lambda_2} = 1000 \times 10^{-3}$$
=1000

Question58

For a cell, $Cu(s) \mid Cu^{2+}(0.001M) \mid Ag^{+}(0.01M) \mid Ag(s)$ the cell potential is found to be 0.43V at 298K. The magnitude of standard electrode potential for Cu^{2+}/Cu is ______ $\times 10^{-2}$V

[Given : $E^{\Theta}_{Ag^{+}/Ag} = 0.80V$ and . $\frac{2.303RRT}{F} = 0.06V$]

[29-Jul-2022-Shift-2]

Answer: 34

Solution:

Anode : $Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^{-}$
Cathode: $[Ag^{+} + e^{-} \longrightarrow Ag(s)]2$
$Cus(s) + 2Ag^{+}(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$
$$E_{cell} = E_{cell}^{0} - \frac{0.06}{2} \log \frac{[Cu^{2+}]}{[Ag^{+}]^{2}}$$
$$0.43 = E_{cell}^{0} - \frac{0.06}{2} \log \left(\frac{10^{-3}}{(10^{-2})^{2}} \right)$$
$$0.43 = E_{cell}^{0} - 0.03 \log 10$$
$$E_{cell}^{0} = 0.46V$$
$$E_{cell}^{0} = E_{Ag^{+}/Ag}^{0} - E_{Cu^{2+}/Cu}^{0}$$
$$E_{Cu^{2+}/Cu^{0}} = (0.80 - 0.46) = 0.34V = 34 \times 10^{-2}$$

Question59

Emf of the following cell at 298K in V is $x \times 10^{-2}$, $Zn \mid Zn^{2+}(0.1M) \mid Ag^{+}(0.01M) \mid Ag$ The value of x is ....

(Rounded off to the nearest integer).
[Given, $E^{\circ}_{Zn^{2+}/Zn} = -0.76V$, $E^{\circ}_{Ag^{+}/Ag} = +0.80V$, $\frac{2.303RT}{F} = 0.059$]
[26 Feb 2021 Shift 2]

Answer: 147

Solution:

$Zn \mid Zn^{2+}(0.1M) \mid Ag^{+}(0.01M) \mid Ag$
Cell reaction : $Zn(s) + 2Ag^{+} \rightleftharpoons Zn^{2+} + 2Ag(s)$
$\Rightarrow Q = \frac{[Zn^{2+}][Ag]^2}{[Zn][Ag^{+}]^2} = \frac{0.1M}{1 \times (0.01)^2} = 10^3$
$n = 2$ and $E_{cell}^{\circ} = E^{\circ}_{Ag^{+}/Ag} - E_{Zn^{\circ}}^{2+}/Zn$
$= 0.8 - (-0.76) = 1.56V$
$\Rightarrow E_{cell} = E_{cell}^{\circ} - \frac{0.059}{n} \log Q = 1.56 - \frac{0.059}{2} \log 10^3$
$= 1.4715V$
$= 147.15 \times 10^{-2}V = 147 \times 10^{-2}V$
$= x \times 10^{-2}V$
$x = 147$

Question60

Copper reduces NO_3^{-} into NO and NO_2 depending upon the concentration of HNO_3 in solution. (Assuming fixed $[Cu^{2+}]$ and $p_{NO} = p_{NO_2}$), the HNO_3 concentration at which the thermodynamic tendency for reduction of NO_3^{-} into NO and NO_2 by copper is same is $10^{x}M$. The value of 2x is ....(Rounded off to the nearest integer)
[Given, $E^{\circ}_{Cu^{2+}/Cu} = 0.34V$, $E^{\circ}_{NO_3^{-}/NO} = 0.96V$
$E^{\circ}_{NO_3^{\circ}/NO_2} = 0.79V$ and at 298K, $\frac{RT}{F}$
(2.303) = 0.059]
[25 Feb 2021 Shift 2]

Answer: 4

Solution:

Cell-I ($HNO_3 \longrightarrow NO$) $3Cu + 2NO_3^{-} + 8H^{+} \longrightarrow 3Cu^{2+} + 2NO + 4H_2O$
$$Q_1 = \frac{[Cu^{2+}]^3 \times (p_{NO})^2}{[NO_3^{-}]^2 \times [H^{+}]^8}$$

$\because\ E_1^\circ = 0.96 - (-0.34) = 1.3\text{V}$

$E_1 = 1.3 - \frac{0.059}{6}\log Q_1$

Cell-II ($HNO_3 \rightarrow NO_2$)

$Cu + 2NO_3^- + 4H^+ \longrightarrow Cu^{2+} + 2NO_2 + 2H_2O$

$Q_2 = \frac{[Cu^{2+}] \times (p_{NO_2})^2}{[NO_3^-]^2 \times [H^+]^4}$

$E_2^\circ = 0.79 - (-0.34)\text{V} = 1.13\text{V}$

$E_2 = 1.13 - \frac{0.059}{2}\log Q_2$

$\because\ E_2$

Now, $E_1 = E_2$

$1.3 - \frac{0.059}{6}\log Q_1 = 1.13 - \frac{0.059}{2}\log Q_2$

$0.17 = \frac{0.059}{6}[\log Q_1 - 3\log Q_2] = \frac{0.059}{6}\log\frac{Q_1}{Q_2}$

$= \frac{0.059}{6}\log\frac{[Cu^{2+}]^3 \times (p_{NO})^2}{[NO_3^-]^2 \times [H^+]^8} \times \frac{[NO_3^-]^6 \times [H^+]^{12}}{[Cu^{2+}]^3 \times (p_{NO_2})^6} = \frac{0.059}{6}\log\frac{[H^+]^4 \times [NO_3^-]^4}{(p_{NO_2})^4}\quad [\because p_{NO} = p_{NO_2}]$

$= \frac{0.059}{6}\log\frac{[HNO_3]^4}{(p_{NO_2})^4}$

Now, $p_{NO_2} \equiv [HNO_3]$

So, $0.17 = \frac{0.059}{6}\log[HNO_3]^8$

$= \frac{0.059}{6} \times 8\log[HNO_3]$

$\log[HNO_3] = 2.16$

$[HNO_3] = 10^{2.16}\text{M} = 10^x\text{M}$

$\therefore\ x = 2.16$

$\Rightarrow\ 2x = 2 \times 2.16 = 4.32 \sim eq4$

Question61

The magnitude of the change in oxidising power of the MnO_4^- / Mn^{2+} couple is $x \times 10^{-4}$ V, if the H^+concentration is decreased from 1M to 10^{-4}M at 25°C. (Assume concentration of $MnO_4^-$and Mn^{2+} to be same on change in H^+concentration). The value of x is .... (Rounded off to the nearest integer). [Given, $\frac{2.303RT}{F} = 0.059$]

[24 Feb 2021 Shift 2]

Answer: 3776

Solution:

Reaction, $MnO_4^- + H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$

Applying Nernst equation, $E_{cell} = E_{cell}^\circ - \frac{0.0591}{n}\log\frac{[P]}{[R]}$

or $E_{cell} = E_{cell}^\circ - \frac{0.0591}{5}\log\frac{[Mn^{2+}]}{[MnO_4^-]}\left[\frac{1}{H^+}\right]^8$

(I) Given, $[H^+] = 1\text{M}$

$E_1 = E^\circ - \frac{0.0591}{5}\log\frac{[Mn^{2+}]}{[MnO_4{}^-]}$

(II) Now, $[H^+] = 10^{-4}M$

$E_2 = E^\circ - \frac{0.0591}{5}\log\frac{[Mn^{2+}]}{[MnO_4{}^-]}\times\frac{1}{(10^{-4})^8}$

$\therefore |E_1 - E_2|$

$|E_1 - E_2| = \frac{0.0591}{5}\times 32 = 0.3776V = 3776\times10^{-4}$

$x = 3776$

Question62

Compound A used as a strong oxidising agent is amphoteric in nature. It is the part of lead storage batteries. Compound A is [26 Feb 2021 Shift 1]

Options:

A. PbO_2

B. PbO

C. $PbSO_4$

D. Pb_3O_4

Answer: A

Solution:

Solution:
In the set of four lead compounds Pb (II) compounds are PbO and $PbSO_4$. PbO_2 is a Pb (IV) compound whereas Pb_3O_4 is a mixed oxide of Pb (II) and Pb (IV) i.e. $2PbO\cdot PbO_2$.

Pb is a member of group 14 and it shows +2 and +4 oxidation states. But due to inert pair effect, Pb^{2+} is more stable than Pb^{4+}. So, Pb (IV) compounds are strong oxidising agents as Pb^{4+} gets easily reduced to more stable Pb^{2+}.

$Pb^{+4} + 2e^- \longrightarrow Pb^{2+}$, $\Delta C^\circ < 0$ (spontaneous)

So, PbO_2 or Pb_3O_4 can be the compound A. But out of these two compounds only PbO_2 is used in lead storage batteries where a grid of lead packed with PbO_2 acts as cathode and also it is amphoteric in nature. It reacts with both acids and alkali.

(i) $PbO_2 + 2HCl \longrightarrow PbCl_2 + Cl_2 + H_2O$

Here, PbO_2 acts as a basic oxide as well as an oxidising agent.

(ii) $PbO_2 + 2NaOH \longrightarrow Na_2PbO_2 + H_2O$

Here, PbO_2 acts as an acidic oxide.

So, the compound A is PbO_2 (option-a).

Question63

Consider the following reaction, $MnO_4{}^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$, $E^\circ = 1.51V$. The quantity of electricity required in Faraday to reduce five moles of $MnO_4{}^-$ is .... [26 Feb 2021 Shift 1]

Answer: 25

Solution:

In the reduction half-cell,
$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H$
$1\text{mol}\ \ 5F$
$5\text{mol}\ \ 25F$
As one mole of MnO_4^-required $5F$ of charge,
5 moles of MnO_4^-will require charge,
$Q = 5 \times 5$ Faraday $= 25$ Faraday

Question64

A 5.0mmol dm^{-3} aqueous solution of KCl has a conductance of 0.55mS when measured in a cell constant 1.3cm^{-1}.
The molar conductivity of this solution is $\text{mSm}^2\text{mol}^{-1}$.
(Round off to the nearest integer)
[16 Mar 2021 Shift 2]

Answer: 14

Solution:

$\Lambda_m = \dfrac{K}{1000 \times C}$
where, $\Lambda_m =$ molar conductivity $\kappa =$ conductivity, $C =$ molarity
Conductivity $(\kappa) = G\dfrac{I}{A} = 0.55 \times (1.3 \times 100)$
$= 55 \times 1.3$
$\Lambda_m = \dfrac{55 \times 1.3}{1000 \times 5 \times 10^{-3}}$
$\Lambda_m = 14.3$
Closest integer, $\Lambda_m = 14$

Question65

The molar conductivities at infinite dilution of barium chloride, sulphuric acid and hydrochloric acid are 280,860 and $426\text{Scm}^2\ \text{mol}^{-1}$ respectively. The molar conductivity at infinite dilution of barium sulphate is .... $\text{Scm}^2\text{mol}^{-1}$ (Round off to the nearest Integer).
[18 Mar 2021 Shift 2]

Answer: 288

Solution:

Molar conductivity of $\mathrm{BaCl_2} = 280\,\mathrm{Scm^2mol^{-1}}$
Molar conductivity of $\mathrm{H_2SO_4} = 860\,\mathrm{Scm^2mol^{-1}}$
Molar conductivity of $\mathrm{HCl} = 426\,\mathrm{Scm^2mol^{-1}}$
Molar conductivity of $\mathrm{BaSO_4} = ?$
From Kohlrausch's law,
$\Lambda_m^{\infty}(\mathrm{BaSO_4}) = \lambda_m^{\omega}(\mathrm{Ba^{2+}}) + \lambda_m^{\omega}(\mathrm{SO_4^{2-}})$
$\Lambda_m^{\infty}(\mathrm{BaSO_4}) = \Lambda_m^{-}(\mathrm{BaCl_2}) + \Lambda_m^{\infty}(\mathrm{H_2SO_4}) - 2\Lambda_m^{\infty}(\mathrm{HCl})$
On putting given values we get,
$= 280 + 860 - 2(426)$
$= 288\,\mathrm{Scm^2mol^{-1}}$

Question66

For the reaction,
$2\mathrm{Fe^{3+}(aq)} + 2\mathrm{l^-(aq)} \longrightarrow 2\mathrm{Fe^{2+}(aq)} + \mathrm{I_2(s)}$
The magnitude of the standard molar Gibbs free energy change,
$\Delta,\ G_m^{\circ} = -.\,.\,.\,.\,.\,.\,.\,.\,.\,.\,.\ \mathrm{kJ}$
(Round off to the nearest integer).

$$\left[\begin{array}{ll} E_{\mathrm{Fe^{2+}/Fe(s)}}^{\circ} = -0.440\mathrm{V}; & E_{\mathrm{Fe^{3+}/Fe(s)}}^{\circ} = -0.036\mathrm{V} \\ E_{\mathrm{I_2/2I^-}} = 0.539\mathrm{V}; & F = 96500\mathrm{C} \end{array}\right]$$

[18 Mar 2021 Shift 1]

Answer: 45

Solution:

Given reaction is
$2\mathrm{Fe^{3+}(aq)} + 2\mathrm{I^-(aq)} \longrightarrow 2\mathrm{Fe^{2+}(aq)} + \mathrm{I_2(s)}$

Now, $\mathrm{Fe^{3+}} \xrightarrow{E_1^{\circ}} \mathrm{Fe^{2+}} \xrightarrow{E_2^{\circ}} \mathrm{Fe}$

$$\overset{E_3^{\circ}}{|\!\!\longrightarrow\!\!\!\!\longrightarrow\!\!\!\!\longrightarrow\!\!|}$$

$nE_1^{\circ} + nE_2^{\circ} = nE_3^{\circ}$ [$n =$ number of electron transferred]
$E_1^{\circ} + 2E_2^{\circ} = 3E_3^{\circ}$
$E_{\mathrm{Fe^{3+}/Fe^{2+}}}^{0} + 2E_{\mathrm{Fe^{2+}/Fe(s)}}^{0} = 3E_{\mathrm{Fe^{3+}/Fe(s)}}^{0}$
$E_{\mathrm{Fe^{3+}/Fe^{2+}}}^{\circ} + (-0.440) \times 2 = (-0.036) \times 3$
$E_{\mathrm{Fe^{3+}/Fe^{2+}}}^{\circ} = 0.772\mathrm{V}$

$E_{cell}^{\circ} = E_{cathode}^{\circ} - E_{anode}^{\circ}$
$= 0.772 - 0.539 = 0.233V$
Using standard Gibb's free energy,
$\Delta G^{\circ} = nFE_{cell}^{\circ} = +2 \times 96500 \times 0.233$
$\Delta G^{\circ} = 44969J = 44.9kJ = 45kJ$

Question67

Potassium chlorate is prepared by electrolysis of KCl in basic solution as shown by following equation.
$6OH^{-} + Cl^{-} \rightarrow ClO_3^{-} + 3H_2O + 6e^{-}$
A current of x A has to be passed for 10h to produce 10.0g of potassium chlorate. the value of x is _____. (Nearest integer)
(Molar mass of $KClO_3 = 122.6g\ mol^{-1}$, $F = 96500C$)
[20 Jul 2021 Shift 2]

Answer: 1

Solution:

$W = \frac{E}{F} \times I \times t$
$10 = \frac{122.6}{96500 \times 6} \times x \times 10 \times 3600$
$X = 1.311$

Question68

For the cell
$Cu(s) \mid Cu^{2+}(aq)(0.1M) \parallel Ag^{+}(aq)(0.01M) \mid Ag(s)$
the cell potential $E_1 = 0.3095V$ For the cell
$Cu(s) \mid Cu^{2+}(aq)(0.01M) \parallel Ag^{+}(aq)(0.001M) \mid Ag(s)$
the cell potential = ——$10^{-2}V$. (Round off the Nearest Integer).
$\left[\text{Use}: \frac{2.303RT}{F} = 0.059\right]$
[27 Jul 2021 Shift 2]

Answer: 28

Solution:

Cell reaction is :

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

Now, $E_{cell} = E^{\circ}_{Cell} - \frac{0.059}{2} \log \frac{[Cu^{2+}]}{[Ag^{+}]^{2}}$ ......(1)

$\therefore E_{1} = 0.3095 = E^{\circ}_{Cell} - \frac{0.059}{2} .\log \frac{0.01}{(0.001)^{2}}$

From (1) and (2), $E_{2} = 0.28V = 28 \times 10^{-2}V$

Question69

Consider the cell at 25°C

$Zn \mid Zn^{2+}(aq), (1M) \parallel Fe^{3+}(aq), Fe^{2+}(aq))\, Pt(s)$

The fraction of total iron present as Fe^{3+} ion at the cell potential of $1.500V$ is $x \times 10^{-2}$. The value of x is _______ . (Nearest integer)

(Given : $E^{\circ}_{Fe^{3+}/Fe^{2+}} = 0.77V$, $E^{\circ}_{Zn^{2+}/Zn} = -0.76V$)

[25 Jul 2021 Shift 1]

Answer: 24

Solution:

$Zn \rightarrow Zn^{2+} + 2e^{-}$

$2Fe^{3+} \rightarrow 2e^{-} + 2e^{2+}$

$Zn + 2Fe^{3+} \rightarrow Zn^{2+} + 2Fe^{2+}$

$E^{\circ}_{cell} = 0.77 - (0.76)$

$= 1.53V$

$1.50 = 1.53 - \frac{0.06}{2} \log \left(\frac{Fe^{2+}}{Fe^{3+}} \right)^{2}$

$\log \left(\frac{Fe^{2+}}{Fe^{3+}} \right) = \frac{0.03}{0.06} = \frac{1}{2}$

$\frac{[Fe^{2+}]}{[Fe^{3+}]} = 10^{1/2} = \sqrt{10}$

$\frac{[Fe^{3+}]}{[Fe^{2+}]} = \frac{1}{\sqrt{10}}$

$\frac{[Fe^{3+}]}{[Fe^{2+}] + [Fe^{3+}]} = \frac{1}{1 + \sqrt{10}} = \frac{1}{4.16}$

$= 0.2402$

$= 24 \times 10^{-2}$

Question70

Assume a cell with the following reaction

$Cu_{(s)} + 2Ag^{+}(1 \times 10^{-3}M) \rightarrow Cu^{2+}(0.250M) + 2Ag_{(s)}$

$E_{cell}^{\ominus} = 2.97V$

E_{cell} for the above reaction is ________ V.
(Nearest integer)
[Given : $\log 2.5 = 0.3979$, $T = 298K$]
[22 Jul 2021 Shift 2]

Answer: 3

Solution:

$$E = E^\circ - \frac{0.059}{2}\log\frac{[Cu^{+2}]}{[Ag^{+}]^{2}}$$
$$= 2.97 - \frac{0.059}{2}\log\frac{0.25}{(10^{-3})^{2}} = 2.81V$$

Question71

The conductivity of a weak acid HA of concentration $0.001 mol\,L^{-1}$ is $2.0\times10^{-5} Scm^{-1}$. If $\Lambda_m^{o}(HA) = 190 Scm^{2}mol^{-1}$, the ionization constant (K_a) of HA is equal to ________ $\times10^{-6}$.
(Round off to the Nearest Integer)
[27 Jul 2021 Shift 1]

Answer: 12

Solution:

$$\Lambda_m = 1000\times\frac{K}{M}$$
$$= 1000\times\frac{2\times10^{-5}}{0.001} = 20 Scm^{2}mol^{-1}$$
$$\Rightarrow\alpha = \frac{\Lambda_m}{\Lambda_m^{\infty}} = \frac{20}{190} = \left(\frac{2}{19}\right)$$
$$HA \rightleftharpoons H^{+} + A^{-}$$
$$0.001(1-\alpha)0.001\alpha0.001\alpha$$
$$\Rightarrow k_a = 0.001\left(\frac{\alpha^{2}}{1-\alpha}\right) = \frac{0.001\times\left(\frac{2}{19}\right)^{2}}{1-\left(\frac{2}{19}\right)}$$
$$= 12.3\times10^{-6}$$

Question72

Given below are two statements.
Statement I The limiting molar conductivity of KCl (strong electrolyte) is higher compared to that of CH_3COOH (weak electrolyte).
Statement II Molar conductivity decreases with decrease in concentration of electrolyte. In the light of the above statements, choose the most appropriate answer from the options given below
[26 Aug 2021 Shift 1]

Options:

A. Statement I is true but statement II is false.

B. Statement I is false but statement II is true.

C. Both statement I and statement II are true.

D. Both statement I and statement II are false.

Answer: D

Solution:

Solution:
Limiting molar conductivity of KCl
$\Lambda_m^\circ \text{KCl} = \Lambda_m^\infty(K^+) + \Lambda_m^\infty(Cl^-)$
$= 73.55 S\ cm^2 / mol + 76.3 S\ cm^2 / mol$
$= 149.3 S\ cm^2 / mol$
Limiting molar conductivity of CH_3COOH
$\Lambda_m^\circ(CH_3COOH) = \Lambda_m^\infty(CH_3COO^-) + \Lambda_m^\infty(H^+)$
$= 349.8 S\ cm^2 / mol + 40.9 S\ cm^2 / mol = 390.7 S\ cm^2 / mol$
∴ Limiting molar conductivity of CH_3COOH is more than KCl.
Hence, statement I is false.
Also, as the concentration decreases, dilution of the electrolyte increases which will increase the degree of dissociation of weak electrolyte. Thus, the number of ions in the solution increases and hence, the molar conductance of electrolyte increases. Therefore, we can see that with decrease in concentration of electrolyte, molar conductance increases.
Hence, statement II is also false.

Question73

These are physical properties of an element.
A. Sublimation enthalpy
B. Ionisation enthalpy
C. Hydration enthalpy
D. Electron gain enthalpy
The total number of above properties that affect the reduction potential is.......... (Integer answer)
[26 Aug 2021 Shift 1]

Answer: 3

Solution:

Solution:
The physical properties of an element is affected by sublimation enthalpy, ionisation enthalpy and hydration enthalpy as these 3 enthalpies will affect the reduction potential.

Question74

For the galvanic cell,
$Zn(s) + Cu^{2+}(0.02M) \rightarrow Zn^{2+}(0.04M) + Cu(s)$,
E_{cell} = ... $\times 10^{-2}$V.(Nearest integer)
[Use $E°_{Cu / cu^{2+}} = -0.34V$, $E°_{Zn / Zn^{2+}} = +0.76V$,
$\frac{2.303RT}{F} = 0.059V$]
[26 Aug 2021 Shift 2]

Answer: 109

Solution:

Solution:
Cell rection,

$$\underset{0.02M}{Cu^{2+}(aq)} + Zn(s) \rightleftharpoons Cu(s) + \underset{0.04M}{Zn^{2+}(aq)}$$

$$E°_{cell} = E°_{Cu^{2+} / Cu} - E°_{Zn^{2+} / Zn}$$
$$= 0.34 - (-0.75) = 1.1V$$
$$E_{cell} = E°_{cell} - \frac{0.059}{2} \log \frac{[Zn^{2+}]}{[Cu^{2+}]}$$
$$= 1.1 - \frac{0.059}{2} \log \frac{[0.04]}{[0.02]}$$
$$= 1.1 - 0.03 \log 2 = 1.1 - 0.03 \times 0.30$$
$$= 1.09V = 109 \times 10^{-2}V$$
∴ Answer is 109.

Question75

The resistance of a conductivity cell with cell constant $1.14cm^{-1}$, containing 0.001 M KCl at 298K is 1500Ω. The molar conductivity of 0.001 M KCl solution at 298K in Scm^2mol^{-1} is ....(Integer answer)
[27 Aug 2021 Shift 2]

Answer: 760

Solution:

Solution:
Given, $C^* = 1.14cm^{-1} = 1.14 \times 10^3m^{-1}$

R = 1500 Ω
C = 0.001 M
G* = $R \times \kappa$
[where, R = resistance
G* = cell constant
κ = conductivity
C = concentration]
$\therefore \kappa = \frac{G^*}{R} = \frac{1.14 \times 10^3 m^{-1}}{1500}$
$\Lambda m = \frac{\kappa}{C}$
[where , Λ_m = molar conductivity]
$= \frac{1.14 \times 10^3}{1500 \times 0.001} = 760 Scm^2 mol^{-1}$.

Question76

Section B : Numerical Type Questions
Consider the following cell reaction,
$Cd(s) + Hg_2SO_4(s) + \frac{9}{5}H_2O(l) \rightleftharpoons CdSO_4 . \frac{9}{5}H_2O(s) + 2\,Hg(l)$
The value of $E_{cell}°$ is 4.315V at 25°C. If $\Delta H° = -825.2\,kJ\,mol^{-1}$, the standard entropy change $\Delta S°$ in JK^{-1} is.... (Nearest integer)
[Given, Faraday constant = $96487 Cmol^{-1}$]
[31 Aug 2021 Shift 1]

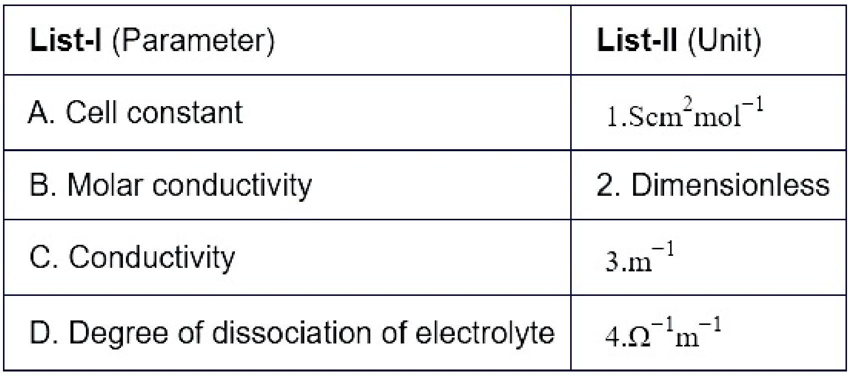
Answer: 25

Solution:

Solution:
$\Delta G° = \Delta H° - T\Delta S°$ ...(i)
where, $\Delta G°$ = standard Gibb's free energy change
$\Delta H°$ = standard change in enthalpy
$\Delta S°$ = standard entropy change
Also, $\Delta G° = -nFE°$
where, n = number of electrons
F = Faraday constant
E° = standard cell potential
∴ Equation (i) is,
$-nFE° = \Delta H° - T\Delta S°$
$-nFE° - \Delta H° = T\Delta S°$
$\Rightarrow \Delta S° = \frac{nFE° + \Delta H°}{T}$ $(2 \times 96487 Cmol^{-1} \times 4.315 V)$
$= \frac{+(-825.2\,kJ / mol)}{298K}$
$= \frac{832.682 \times 10^3 J\,mol - 825.2 \times 10^3 J / mol}{298K}$
$= 25.11 J / Kmol$

Question77

Match List-I with List-II.



List-I (Parameter)	**List-II** (Unit)
A. Cell constant	1.Scm^2mol^{-1}
B. Molar conductivity	2. Dimensionless
C. Conductivity	3.m^{-1}
D. Degree of dissociation of electrolyte	4.$\Omega^{-1}m^{-1}$

Choose the most appropriate answer from the options given below [31 Aug 2021 Shift 2]

Options:

A. A-3 B-1 C-4 D-2

B. A-3 B-1 C-2 D-4

C. A-1 B-4 C-3 D-2

D. A-2 B-1 C-3 D-4

Answer: A

Solution:

Solution:

(A) Cell constant $= \frac{I}{A} m^{-1}$

(B) Molar conductivity $(\lambda_m) = \frac{\kappa \times 1000}{\text{molarity}} \mathrm{Scm^2 mol^{-1}}$

(C) Conductivity $(\kappa) = \frac{1}{\rho} = \frac{1}{RA} \Omega^{-1} m^{-1}$

(D) Degree of dissociation of electrolyte = Number of moles dissociated out of one mole.

It is a ratio. Hence, it is dimensionless.

Thus, the correct match is

A → 3, B → 1, C → 4, D → 2.

Question78

If the conductivity of mercury at 0°C is $1.07 \times 10^6 Sm^{-1}$ and the resistance of a cell containing mercury is 0.243Ω, then the cell constant of the cell is $x \times 10^4 m^{-1}$. The value of x is........ .(Nearest integer) [1 Sep 2021 Shift 2]

Answer: 26

Solution:

Conductance (G) is reciprocal of resistance (R).

$R = \frac{1}{G}$ or $G = \frac{1}{R} = \frac{1}{0.243\Omega} = 4.115\Omega^{-1}$
Relation between conductance (G),
conductivity (kappa) and cell constant $\left(\frac{I}{A}\right)$ is given as

$k = \frac{Gl}{A}$

$\Rightarrow \frac{l}{A} = \frac{k}{G} = \frac{1.07 \times 10^6 Sm^{-1}}{4.115\Omega^{-1}} = 26 \times 10^4 m^{-1}$

$\Rightarrow x = 26$
∴ Answer is 26.

Question79

The equation that is incorrect is:
[Jan. 07, 2020 (II)]

Options:

A. $(\Lambda_m{}^0)_{NaBr} - (\Lambda_m{}^0)_{NaCl} = (\Lambda_m{}^0)_{KBr} - (\Lambda_m{}^0)_{KCl}$

B. $(\Lambda_m{}^0)_{KCl} - (\Lambda_m{}^0)_{NaCl} = (\Lambda_m{}^0)_{KBr} - (\Lambda_m{}^0)_{NaBr}$

C. $(\Lambda_m{}^0)_{H_2O} = (\Lambda_m{}^0)_{HCl} + (\Lambda_m{}^0)_{NaOH} - (\Lambda_m{}^0)_{NaCl}$

D. $(\Lambda_m{}^0)_{NaBr} - (\Lambda_m{}^0)_{Nal} = (\Lambda_m{}^0)_{KBr} - (\Lambda_m{}^0)_{NaBr}$

Answer: D

Solution:

Solution:
$(\Lambda_m{}^0)_{NaBr} - (\Lambda_{mNal}{}^0)_{Nal}$
$= \Lambda_m{}^0Na^+ + \Lambda_m{}^0Br^- - (\Lambda_m{}^0N^+ + \Lambda_m{}^0I^-)$
$= \Lambda_m{}^0N_a{}^+ + \Lambda_m{}^0Br^- - \Lambda_m{}^0N^+ - \Lambda_m{}^0I$
$= \Lambda_m{}^0Br^- + \Lambda_m{}^0I^-$
$(\Lambda_m{}^0)_{KBr} - (\Lambda_m{}^0)_{NaBr}$
$= \Lambda_m{}^0K^+ + \Lambda_m{}^0Br^- - (\Lambda_m{}^0Na^+ + \Lambda_m{}^0Br^-)$
$= \Lambda_m{}^0K^+ + \Lambda_m{}^0Br^- - \Lambda_m{}^0N^+ - \Lambda_m{}^0Br^-$
$= \Lambda_m{}^0K^+ + \Lambda_m{}^0Na^+$
$\therefore (\Lambda_m{}^0)_{NaBr} - (\Lambda_m{}^0)_{Nal} \neq (\Lambda_m{}^0)_{KBr} - (\Lambda_m{}^0)_{NaBr}$

Question80

108g of silver (molar mass 108gmol $^{-1}$) is deposited at cathode from $AgNO_3(aq)$ solution by a certain quantity of electricity. The volume (in L) of oxygen gas produced at 273K and 1 bar pressure from water by the same quantity of electricity is
[NV, Jan. 09, 2020(I)]

Answer: 5.68

Solution:

Solution:
No. of moles of silver deposited.

$$= \frac{108}{108} = 1\text{mol}$$

$Ag^{+} + e^{-} \rightarrow Ag$

$1F$ charge is required to deposit 1mol e of Ag

$$H_2O \rightarrow \frac{1}{2}O_2 + 2H^{+} + 2e^{-}$$

$2F$ charge deposit → $\frac{1}{2}$ moles of oxygen

$1F$ charge will deposit → $\frac{1}{4}$ moles of oxygen

$$V_{O_2} = \frac{nRT}{P}$$

$$= \frac{1}{4} \times \frac{0.08314\text{LbarK}^{-1}\text{mol}^{-1} \times 273\text{K}}{1\text{ bar}}$$

$$= \frac{1}{4} \times 22.7$$

$$V_{O_2} = 5.675L$$

Question81

What would be the electrode potential for the given half cell reaction at pH = 5 ?

________ .

$$\mathbf{2H_2O \rightarrow O_2 + 4H^{\oplus} + 4c^{-}; E_{red}^{0} = 1.23V}$$

$(R = 8.314J\ mol^{-1}K^{-1}$; Temp = 298K ; oxygen under std. atm. pressureof 1 bar)

[NV, Jan. 08, 2020 (I)]

Answer: 1.52

Solution:

$$E = 1.23 - \frac{0.0591}{4}\log[H^{+}]^{4}$$

$$= 1.23 + 0.0591 \times pH$$

$$= 1.23 + 0.0591 \times 5$$

$$= 1.23 + 0.2955$$

$$= 1.52V$$

Question82

For an electrochemical cell
Sn(s) | Sn^{2+}(aq, 1M) | Pb^{2+}(aq, 1M) | Pb(s) the ratio $\frac{[Sn^{2+}]}{[Pb^{2+}]}$
when this cell attains equilibrium is
[NV, Jan. 08, 2020 (II)]

Answer: 2.15

Solution:

At equilibrium state $E_{cell} = 0$; $E_{cell}^{\ 0} = 0.01V$

$Sn + Pb^{2+} \longrightarrow Sn^{2+} + Pb$

$$E = E_{cell}^{\ 0} - \frac{0.06}{n}\log\frac{[P]}{[R]}$$

$$0 = 0.01 - \frac{0.06}{2}\log\frac{[Sn^{2+}]}{[Pb^{2+}]}$$

$$-0.01 = -\frac{0.06}{2}\log\frac{[Sn^{2+}]}{[Pb^{2+}]}$$

$$\frac{1}{3} = \log\frac{[Sn^{2+}]}{[Pb^{2+}]}$$

$$\frac{[Sb^{2+}]}{[Pb^{2+}]} = 10^{1/3} = 2.15$$

Question83

Given that the standard potentials (E) of Cu^{2+} / Cu and Cu^{+}/ Cu are 0.34V and 0.522V respectively, the E^{0} of Cu^{2+} / Cu is:
[Jan. 07,2020(I)]

Options:

A. +0.182V

B. +0.158V

C. −0.182V

D. −0.158V

Answer: B

Solution:

Solution:

$Cu^{2+} + 2e^{-} \longrightarrow Cu,\ \Delta G_1^{\circ} = -2F(0.34)\ . . .$ (i)

$Cu^{+} + e^{-} \longrightarrow Cu,\ \Delta G_2^{\circ} = -F(0.522)\ . . .$ (ii)

Subtract (ii) from (i)

$Cu^{2+} + e^{-} \longrightarrow Cu^{+},\ \Delta G_3^{\circ} = -F(E^{0})$

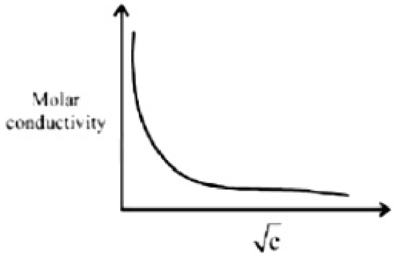
$\therefore \Delta G_1^{\circ} - \Delta G_2^{\circ} = G_3^{\circ}$

$\Rightarrow -F E^{\circ} = -2F(0.34) + F(0.522)$

$\Rightarrow E^{\circ} = 0.68 - 0.522 = 0.158V$

Question84

The variation of molar conductivity with concentration of an electrolyte (X) in aqueous solution is shown in the given figure.



The electrolyte X is :
[Sep. 05, 2020 (II)]

Options:

A. HCl

B. NaCl

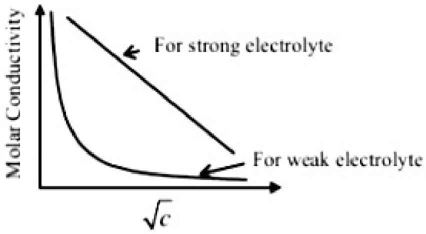
C. KNO_3

D. CH_3COOH

Answer: D

Solution:

Solution:



Among given electrolytes, CH_3COOH is weak electrolyte.

Question85

Let C_{NaCl} and C_{BaSO_4} be the conductances (in S) measured for saturated aqueous solutions of NaCl and $BaSO_4$, respectively, at a temperature T. Which of the following is false?
[Sep. 03, 2020 (I)]

Options:

A. Ionic mobilities of ions from both salts increase with T

B. $C_{BaSO_4}(T_2) > C_{BaSO_4}(T_1)$ for $T_2 > T_1$

C. $C_{NaCl}(T_2) > C_{NaCl}(T_1)$ for $T_2 > T_1$

D. $C_{NaCl} \gg C_{BaSO_4}$ at a given T

E. None of above

Answer: E

Solution:

Solution:
(N) Ionic mobility increases with increase in temperature, which increases the conductance of the solution, while conductance of $NaCl$ solution is independant of temperature above $400^\circ C$. Because the temperature is not given, so none of the option is correct.

Question86

Potassium chlorate is prepared by the electrolysis of KCl in basic solution $6OH^- + Cl^- \rightarrow ClO_3^- + 3H_2O + 6e^-$ If only 60% of the current is utilized in the reaction, the time (rounded to the nearest hour) required to produce 10 g of $KClO_3$ using a current of 2A is _______ .
(Given : $F = 96,500 C mol^{-1}$; molar mass of $KClO_3 = 122 g\ mol^{-1}$)
[NV, Sep. 06, 2020 (I)]

Answer: 11

Solution:

Solution:
$6OH^- + Cl^- \longrightarrow ClO_3^- + 3H_2O + 6e^-$
For synthesis of 1 mole of ClO_3^-, $6F$ of charge is required.
$\because$ Current efficiency $= 60\%$
$\therefore$ To synthesis 1 mole of ClO_3^-, $10F$ of charge is required.
Tosynthesis $\frac{10}{122}$ moles of $KClO_3$, charge $= \frac{10 \times 10}{122}F$ $Q = I\ .\ t$
$t = \frac{100 \times 96500}{122 \times 2} = 39549.18s$
$= \frac{79098.365}{3600s} = 10.99h$
$\therefore t = 11h.$

Question87

250mL of a waste solution obtained from the workshop of a goldsmith contains $0.1M\ AgNO_3$ and $0.1M\ AuCl$. The solution was electrolyzed at $2V$ by passing a current of 1 A for 15 minutes. The metal\/metals

electrodeposited will be:
[Sep. 04, 2020(II)]

$$\left(E^{0}_{Ag^{+}/Ag} = 0.80V , E^{0}_{Au^{+}/Au} = 1.69V \right)$$

Options:

A. only gold

B. silver and gold in proportion to their atomic weights

C. only silver

D. silver and gold in equal mass proportion

Answer: A

Solution:

Solution:
Millimoles of $Au^{+} = 0.1 \times 250 = 25$
Mole of $Au^{+} = \frac{25}{1000} = \frac{1}{40} = 0.025$
Similarly, moles of $Ag^{+} = 0.025$
Charge passed $= I \times t = 1 \times 15 \times 60 = 900C$
Moles of e- passed $= \frac{900}{96500} = 0.0093mol$.
Species withhigher value of SRP will get deposited first at cathode.

Question88

An acidic solution of dichromate is electrolyzed for 8 minutes using 2 A current. As per the following equation
$$Cr_2O_7^{2-} + 14H^{+} + 6e^{-} \rightarrow 2Cr^{3+} + 7H_2O$$
The amount of Cr^{3+} obtained was 0.104g. The efficiency of the process (in %) is (Take : $F = 96000C$, At. mass of chromium $= 52$) __________.
[NV, Sep. 03,2020 (II)]

Answer: 60

Solution:

Solution:
Charge (Q) $= It = 2 \times 8 \times 60 = 960C$
$$\Rightarrow \frac{960}{96000} = 0.01F$$
$$Cr_2O_7^{2-} + \underset{0.01F}{14H^{+} + 6e^{-}} \longrightarrow \underset{\frac{1}{3} \times 0.01\text{ mole}}{2Cr^{3+}} + 7H_2O$$
Theoritical mass of $Cr^{3+} = \frac{1}{3} \times \frac{960}{96000} \times 52 = 0.173g$

So, efficiency $= \frac{W_{\text{actual}}}{W_{\text{Theoritical}}} \times 100 = \frac{0.104}{0.173} \times 100 = 60\%$

Question89

For the given cell;
$Cu(s) \mid Cu^{2+}(C_1M) \mid Cu^{2+}(C_2M) \mid Cu(s)$
change in Gibbs energy (ΔG) is negative, if:
[Sep. 06, 2020 (II)]

Options:

A. $C_1 = C_2$

B. $C_2 = \frac{C_1}{\sqrt{2}}$

C. $C_1 = 2C_2$

D. $C_2 = \sqrt{2}C_1$

Answer: D

Solution:

For the concentration cell, $E_{\text{cell}}^{0} = 0$

Anode : $Cu(s) \longrightarrow Cu^{2+}(aq)_A$

$$\begin{array}{l} \text{Cathode} \quad : Cu^{2+}(aq)_e \longrightarrow Cu(s) \\ \hline \text{Overall} \quad \underset{(C_2M)}{Cu^{2+}(aq)_C} \longrightarrow \underset{(C_1M)}{Cu^{2+}(aq)_A} \end{array}$$

As $\Delta G = -nF\,E$

$\text{If} \quad \Delta G = -ve, \quad \text{then} \quad E_{\text{cell}} \quad \text{is} \quad +v_e \ .$

$$E_{\text{cell}} = E_{\text{cell}}^{0} - \frac{RT}{2F} \ln \frac{C_1}{C_2}$$

$$E_{\text{cell}} = 0 - \frac{RT}{2F} \ln \frac{C_1}{C_2}$$

$$E_{\text{cell}} = \frac{RT}{2F} \ln \frac{C_2}{C_1}$$

$\text{So,} \ C_2 > C_1 \ .$

Thus, $C_2 = \sqrt{2}C_1$ relation is correct.

Question90

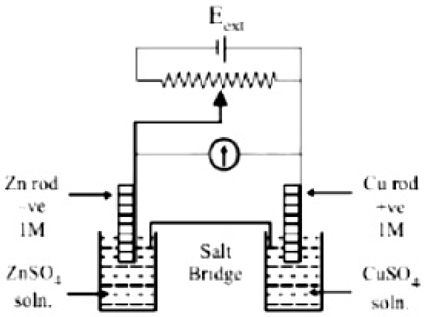
An oxidation-reduction reaction in which 3 electrons are transferred has a ΔG^0 of $17.37\,kJ\,mol^{-1}$ at $25^{\circ}C$. The value of $E_{cell}^{0}(in\,V)$ is ______ $\times 10^{-2}$
$(1F = 96,500\,C\,mol^{-1})$
[NV, Sep. 05, 2020(I)]

Answer: -6

Solution:

$\Delta G^0 = -nFE_{cell}^{\,0}$
$17.37 \times 10^3 = -3 \times 96500 \times E_{cell}^{\,o}$
$E_{cell}^{\,o} = -0.06V \sim eq - 6.0 \times 10^{-2}V$

Question91



$E_{Cu^{2+}|Cu}^{\,o} = +0.34V$; $E_{Zn^{2+}|Zn}^{\,o} = -0.76V$
Identify the incorrect statement from the options below for the above cell :
[Sep. 04, 2020(I)]

Options:

A. If $E_{ext} > 1.1V$, e^- flows from Cu to Zn

B. If $E_{cxt} > 1.1V$, Zn dissolves at Zn electrode and Cu deposits at Cu electrode

C. If $E_{ext} < 1.1V$, Zn dissolves at anode and Cu deposits at cathode

D. If $E_{ext} = 1.1V$, no flow of e^- or current occurs

Answer: B

Solution:

Solution:
$E_{cell}^{\,\circ} = E_{Cu^{2+}|Cu}^{\,o} - E_{Zn^{2+}|Zn}^{\,\circ} = 1.1V$
So, if $E_{ext.} = 1.1V$ no electron will flow
At, $E_{ext.} > 1.1V$ cell act as electrolytic cell and electron will flowfrom Cu to Zn.
At, $E_{ext.} < 1.1V$ cell act as electrochemical cells so Zn dissolve and Cu deposit.

Question92

The photoelectric current from Na (work function, $w_0 = 2.3\text{eV}$) is stopped by the output voltage of the cell Pt(s) | H_2(g, 1bar) | H Cl (aq., pH = 1) | AgCl (s) | Ag(s) The pH of aq. HCl required to stop the photoelectric current from K ($w_0 = 2.25\text{cV}$), all other conditions remaining the same, is ______ $\times 10^{-2}$ (to the nearest integer).

Given,

$2.303\frac{RT}{F} = 0.06\text{V}$; $E^{0}_{AgCl \mid Ag \mid C^{-}} = 0.22\text{V}$

[NV, Sep. 03 , 2020(I)]

Answer: 142

Solution:

Solution:

Sodium metal :

$E = E_0 + (KE)_{max}$; $E_{cell}^{0} = 0.22\text{V}$

Cell reaction

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

Anode : $\frac{1}{2}H_2(g) \longrightarrow H^+(aq) + e^-$

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

$E_{cell} = E_{cell}^{0} - \frac{0.06}{1}\log[H^+][Cl^-]$

$E_{cell} = 0.22 - \frac{0.06}{1}\log[10^{-1}][10^{-1}]$

$= 0.22 + 0.12 = 0.34\text{V}$

$(KE)_{max} = E_{cell} = 0.34\text{eV}$

So, $E = 2.3 + 0.34 = 2.64\text{eV}$ = Energy of photon incident

For potassium metal :

$E = E_0 + (KE)_{max}$

$2.64 = 2.25 + (KE)_{max}$

$(KE)_{max} = 0.39 = E_{cell}$

Cell reaction

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

Anode : $\frac{1}{2}H_2(g) \longrightarrow H^+(aq) + e^-$

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

$E_{cell} = E_{cell}^{0} - \frac{0.06}{1}\log[H^+][Cl^-]$

$0.39 = 0.22 - 0.06\log[H^+]^2$

$0.39 = 0.22 - 0.12\log[H^+]$

$0.17 = 0.12 \times pH$

$pH = 17/12 = 1.4166 \sim eq1.42 = 142 \times 10^{-2}$

Question93

The Gibbs energy change (in J) for the given reaction at $[Cu^{2+}] = [Sn^{2+}] = 1M$ and $298K$ is :

$Cu(s) + Sn^{2+}(aq.) \longrightarrow Cu^{2+}(aq.) + Sn(s)$
$\left(E^0_{Sn^{2+}|Sn} = -0.16V, E^0_{Cu^{2+}|Cu} = 0.34V, \text{ Take } F = 96500 C mol^{-1}\right)$
[NV, Sep. 02, 2020(I)]

Answer: 96500

Solution:

Solution:
Sol. $E_{cell}^0 = E_{Sn^{2+}/Sn}^0 - E_{Cu^{2+}/Cu}^0$
$= -0.16 - 0.34$
$= -0.50V$
$\Delta G^0 = -nFE_{cell}^0$
$= -2 \times 96500 \times (-0.5)$
$= 96500J$
$= 96.5KJ = 96500J$

Question94

For the disproportionation reaction $2Cu^+(aq) \rightleftharpoons Cu(s) + Cu^{2+}(aq)$ at 298K, $\ln K$
(where K is the equilibrium constant) is $\times 10^{-1}$.
Given
[NV, Sep. 02,2020(II)]
$\left(E_{Cu^{2+}/Cu^+}^0 = 0.16V; E_{Cu^+/Cu}^0 = 0.52V; \frac{RT}{F} = 0.025\right)$

Answer: 144

Solution:

Solution:
$2Cu^+(aq) \rightleftharpoons Cu(s) + Cu^{2+}(aq)$
$E_{cell}^0 = E_{Cu^+/Cu}^o - E_{Cu^{2+}/Cu^+}^o = 0.52 - 0.16 = 0.36V$
$E_{cell}^0 = \frac{RT}{nF}\ln K_{eq}$
$\Rightarrow 0.36 = \frac{0.025}{1}\ln K$
$\Rightarrow \ln K = 14.4 = 144 \times 10^{-1}$

Question95

Λ_m° for $NaCl$, HCl and NaA are 126.4, 425.9 and 100.5Scm^2mol^{-1}, respectively. If the conductivity of 0.001 M HA is $5 \times 10^{-5}Scm^{-1}$, degree of dissociation of HA is :
[Jan. 12, 2019 (II)]

Options:

A. 0.50

B. 0.25

C. 0.125

D. 0.75

Answer: C

Solution:

Solution:
Given: $\Lambda_m^\circ(NaCl) = 126.4Scm^2mol^{-1}$, $\Lambda_m^\circ(HCl) = 425.9Scm^2mol^{-1}$ and $\Lambda_m^\circ(NaA) = 100.5Scm^2mol^{-1}$

$\Lambda_m^\circ(HA) = \lambda_H^\circ + \lambda_{A^-}^\circ$

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

$= \Lambda_m^\circ(HCl) + \Lambda_m^\circ(NaA) - \Lambda_m^\circ(NaCl)$

$\Lambda_m^\circ(HA) = 429.5 - 126.4 + 100.5 = 400Scm^2mol^{-1}$

$\kappa(HA) = 5 \times 10^{-5}Scm^{-1}$

$$(HA) = \frac{\kappa(HA) \times 1000}{\text{Molality of HA}} = \frac{5 \times 10^{-5} \times 1000}{0.001} = 50Scm^2mol^{-1}$$

$$\alpha = \frac{\Lambda_m(HA)}{\Lambda_m^\circ(HA)} = \frac{50}{400} = 0.125$$

Question96

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 $PbSO_4 = 303gmol^{-1}$)
[Jan. 9,2019(I)]

Options:

A. 22.8

B. 15.2

C. 7.6

D. 11.4

Answer: A

Solution:

According to the Faraday's law of electrolysis, nF of current is required for the deposition of $1 mol$
According to the reaction,
$Ni(NO_3)_2 \longrightarrow Ni^{2+} + 2NO_3^-$
$2F$ of current deposits $= 1 mol$
$\therefore 0.1F$ of currentdeposits $= \frac{0.1}{2} = 0.05 mol$

Question97

The standard electrode potential E^0 and its temperature coefficient $\left(\frac{dE^0}{dT}\right)$ for a cell are $2V$ and $-5 \times 10^{-4} V K^{-1}$ at $300K$ respectively. The cell reaction is:
$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu$
The standard reaction enthalpy ($\Delta_r H^0$) at $300K$ in $kJ mol^{-1}$ is,
[Use $R = 8 J K^{-1} mol^{-1}$ and $F = 96,000 C mol^{-1}$]
[Jan. 12, 2019 (I)]

Options:

A. -412.8

B. -384.0

C. 192.0

D. 206.4

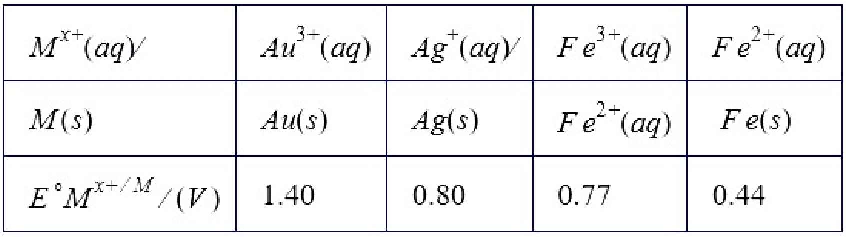
Answer: A

Solution:

$\Delta G^\circ = \Delta_r H^\circ - T \Delta S^\circ$
$\Delta_r H^\circ = \Delta G^\circ + T \Delta S^\circ$
$\Delta_r H^\circ = -nFE^\circ + TnF \frac{dE}{dT}$
$\Delta_r H^\circ = -nFE^\circ + nFT \frac{dE}{dT}$
Cell reaction: $Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$
$\Delta_r H^\circ = -nF\left(E^\circ - \frac{TdE^\circ}{dT}\right)$
$\Delta_r H^\circ = -2 \times 96000(2 - 300 \times -5 \times 10^{-4})$
$\Delta_r H^\circ = -2 \times 96000(2 + 300 \times 5 \times 10^{-4})$
$= -2 \times 96000(2 + 0.15)$
$= -412.8 \times 10^3 J / mol$; $= -412.8 kJ / mol$

Question98

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



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

If $E_{Zn^{2+}/Zn}{}^{2} = -0.76V$, which cathode will give a maximum value of $E_{ecll}{}^{\circ}$ per electron transferred?
[Jan. 11, 2019 (I)]

Options:

A. Ag^{+} / Ag

B. Fe^{3+} / Fe^{2+}

C. Au^{3+} / Au

D. Fe^{2+} / Fe

Answer: A

Solution:

$$Zn + 2Ag^{+} \longrightarrow 2Ag + Zn^{2+}(n = 2e^{-})$$
$$E_{cell}{}^{\circ} = (E_{R.P}{}^{\circ})_{cathode} - (E_{R.P}{}^{\circ})_{anode}$$
$$= 0.80 - (-0.76) = 1.56V \text{ for } 2e$$
$$\therefore E_{cell}{}^{\circ} \text{ for } 1e^{-} = \frac{1.56}{2} = 0.78V$$

Question99

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 $E_{cell}{}^{0}$ of this reaction at 298K

$\left[2.303 \frac{RT}{F} \text{at } 298K = 0.059V\right]$
[Jan. 11, 2019 (II)]

Options:

A. 0.04736mV

B. 0.4736mV

C. 0.4736V

D. 0.04736V

Answer: C

Solution:

Solution:

$E_{cell}^{0} = \frac{2.303RT}{nF} \log K_C$ or $E_{cell}^{0} = \frac{0.059V}{n} \log K_C$

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

Question100

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

$\{ \text{Given} : \frac{2.303RT}{F} = 0.06V \text{ at } 298K \}$
[Jan. 10, 2019 (II)]

Options:

A. 0.94V

B. 0.76V

C. 0.40V

D. 0.20V

Answer: D

Solution:

Solution:
Given that:
$Pt(s) \mid H_2(g, 1bar) / HCl(aq) \mid AgCl(s) / Ag(s) \mid Pt(s)$
$E_{cell} = 0.92V$
Now, $E_{cell} = E^{\circ}_{H_2(g) / H^+(aq)} + E^{\circ}_{AgCl(s)} Ag(s) Cl^- - \frac{0.06}{n} \log Q$
Cell reaction: $\frac{1}{2} H_2(g) \longrightarrow H^+(aq) + e^-$
$AgCl(s) + e^- \longrightarrow Ag(s) + Cl^-(aq)$
Net cell reaction:
$\frac{1}{2} H_2(g) + AgCl(s) \longrightarrow Ag(s) + H^+ + Cl^-(aq)$
$\therefore Q = \frac{[H^+][Cl^-]}{(P_{H_2})^{1/2}}$
Here, 10^{-6} molal HCl solution is used
So $Q = \frac{10^{-6} \times 10^{-6}}{1} = 10^{-12}$
(assuming molality = molarity)
Now, $0.92 = E_{AgCl(s)} Ag(s), Cl^- - \frac{0.06}{1} \log 10^{-12}$
$E^{\circ}_{AgCl(s) Ag(s), Cl^-} = 0.92 + [0.06 \times (-12)]$
$= 0.92 - 0.72 = 0.20V$

Question101

If the standard electrode potential for a cell is 2V at 300K, the equilibrium constant (K) for the reaction
$Zn(s) + Cu^{2+}(aq) \rightleftharpoons Zn^{2+}(aq) + Cu(s)$
at 300K is approximately (R = 8J $K^{-1}mol^{-1}$, F = 96000$Cmol^{-1}$)
[Jan. 9, 2019 (II)]

Options:

A. e^{-80}

B. e^{-160}

C. e^{320}

D. c^{160}

Answer: D

Solution:

Solution:
We know that,

$\ln K = \frac{nFE_{cell}^{\circ}}{RT}$

$\Delta G = -nFE_{cell}^{\circ}$

$\Delta G = -RT \ln k$

Now,

After putting the given values, we get $\ln K = \frac{2 \times 96000 \times 2}{8 \times 300} = 160$

$\therefore K = e^{160}$

Question102

The decreasing order of electrical conductivity of the following aqueous solutions is:

0.1 M Formic acid (A),
0.1 MAcetic acid (B),
0.1 M Benzoic acid (C).
[April 12, 2019 (II)]

Options:

A. A > C > B

B. C > B > A

C. A > B > C

D. C > A > B

Answer: A

Solution:

Solution:
Order of acidic strength is $HCOOH > C_6H_5COOH > CH_3COOH$
More the acidic strength more will be the dissociation of acid into ions and more will be the conductivity.
Thus, order of conductivity will be,
$\underset{(A)}{HCOOH} \underset{>}{>} \underset{(C)}{C_6H_5COOH} \underset{>}{>} \underset{(B)}{CH_3COOH}$

Question103

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: [April 10,2019 (I)]

Options:

A. Both S1 and S2 are wrong

B. S1 is wrong and S2 is correct

C. Both S1 and S2 are correct

D. S1 is correct and S2 is wrong

Answer: B

Solution:

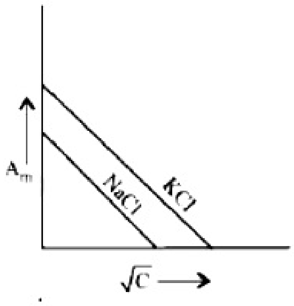
Solution:
Conductivity of an electrolyte is the conductance of $1 \mathrm{cm}^3$ of the given electrolyte. It increases with the increase in concentration of electrolyte due to increase in the number of ions per unit volume. Molar conductivity (λ_m) is theconductance of a solution containing 1 mole of 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.

Question104

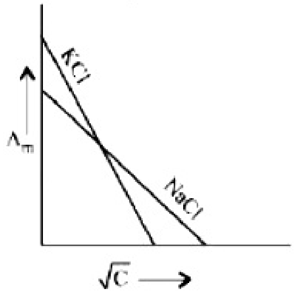
Which one of the following graphs between molar conductivity (A_m) versus $\sqrt{C}$ is correct?
[April 10, 2019 (II)]

Options:

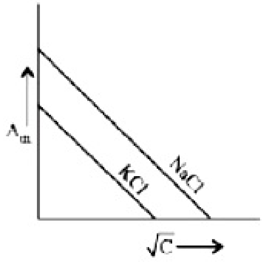
A.



B.



C.



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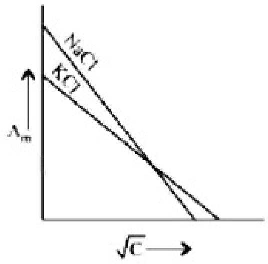


Answer: A

Solution:

Solution:
Since, KCl is more conducting than $NaCl$, therefore, graph (a) is correct.

Question105

A solution of $Ni(NO_3)_2$ is electrolysed between platinum electrodes using 0.1 Faraday electricity. How many mole of Ni will be deposited at the cathode?
[April 9, 2019 (II)]

Options:

A. 0.05

B. 0.20

C. 0.15

D. 0.10

Answer: C

Solution:

The reaction involved in the process is given below
$PbSO_4 \rightarrow Pb^{2+} + SO_4^{2-}$
$Pb^{2+} + H_2O \rightarrow PbO_2 + 4H^+ + 2e^-$
So, half cell reaction is
$PbSO_4 \rightarrow Pb^{4+} + 2e^-$
According tothe reaction:
$PbSO_4 \rightarrow Pb^{4+} + 2e^-$
We require $2F$ for the electrolysis of $1 mol$ or $303 g$ of $PbSO_4$
∴ Amount of $PbSO_4$ electrolysed by $0.05F$

$$= \frac{303}{2} \times 0.05 = 7.575g \approx 7.6g$$

Question106

The standard Gibbs energy for the given cell reaction in $kJ\,mol^{-1}$ at $298K$ is:
$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$
$E^{\circ} = 2V$ at $298K$
(Faraday's constant, $F = 96000 C mol^{-1}$)
[April 9, 2019(I)]

Options:

A. -384

B. 384

C. 192

D. -192

Answer: A

Solution:

Solution:
$\Delta G^{\circ} = -nF E_{cell}^{\circ}$
$= -2 \times (96000) \times 2V = -384000 J / mol = -384 kJ / mol$

Question107

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

$E_{Ag^{+}/Ag}^{0} = xV$; $E_{Fc^{2+}/Fe}^{0} = yV$; $E_{Fe^{3+}/Fe}^{0} = zV$
[April 8, 2019 (II)]

Options:

A. $x - z$

B. $x - y$

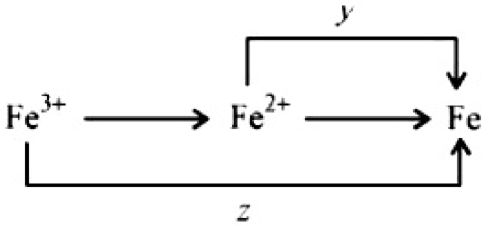
C. $x + 2y - 3z$

D. $x + y - z$

Answer: C

Solution:

$Ag^{+}(aq) + Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + Ag(s)$
$E_{cell}^{0} = E_{Ag^{+}/Ag}^{0} - E_{Fe^{3+}/Fe^{2+}}^{0}$



$E_{Fe^{3+}/Fe^{2+}}^{0} = 3z - 2y$
$E_{Ag^{+}/Ag}^{0} = x$ (given)
$\therefore E_{cell}^{0} = x - 3z + 2y$

Question108

How long (approximate) should water be electrolysed by passing through 100 amperes current so that the oxygen released can completely burn 27.66g of diborane? (Atomic weight of B = 10.8u) [2018]

Options:

A. 6.4 hours

B. 0.8 hours

C. 3.2 hours

D. 1.6 hours

Answer: C

Solution:

Solution:

$B_2H_6 + 3O_2 \longrightarrow B_2O_3 + 3H_2O$

27.66g of B_2H_6(1mol e) requires 3 moles of oxygen (O_2) for complete burning.

Now the oxygen is produced by the electrolysis of H_2O.

On electrolysis: $2H_2O \overset{4F}{\longrightarrow} 2H_2 + O_2$

1 mole O_2 is produced by $4F$ charge

$\therefore 3\text{mol e}O_2$ will be produced by $12F$ charge.

$\because Q = I\,t$

$12 \times 96500C = I \times t$

$12 \times 96500C = 100 \times t$

$t = \dfrac{12 \times 96500}{100}$second

$= \dfrac{12 \times 96500}{100 \times 3600}$ hour

$= 3.2$ hours

Question109

When 9.65 ampere current was passed for 1.0 hour into nitrobenzene in acidic medium, the amount of p -aminophenol produced is [Online April 16, 2018]

Options:

A. 109.0g

B. 98.1g

C. 9.81g

D. 10.9g

Answer: C

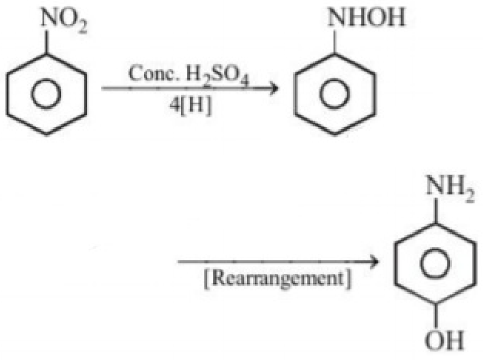
Solution:

Solution:

9.65 ampere current was passed for 1.0 hour (3600 seconds)

Number of moles of electrons passed

$= \dfrac{I\,(A) \times t(s)}{96500} = \dfrac{9.65A \times 3600s}{96500} = 0.36\text{mol es}$



$C_6H_5NO_2 + 4e^- + 4H^+ \longrightarrow$ p -Aminophenol $+ H_2O$

$\because$4 moles of electrons reduces 1 mole of nitrobenzene to p -aminophenol. $\therefore$0.36 moles of electrons will reduce

$\frac{0.36}{4}$ = 0.09 moles of nitrobenzene to p -aminophenol.
p -aminophenol molar mass $= 109.14\text{g / mol}$
Mass of p -aminophenol obtained $= 109.14\text{g / mol} \times 0.09\text{mol} = 9.81\text{g}$

Question110

When an electric current is passes through acidified water, 112mL of hydrogen gas at N.T.P. was collected at the cathode in 965 seconds. The current passed, in ampere, is :
[Online April 15,2018 (I)]

Options:

A. 2.0

B. 0.1

C. 0.5

D. 1.0

Answer: D

Solution:

Solution:
Reduction at cathode:
$2e^{-} + 2H_2O \rightarrow H_2 + 2OH^{-}$ (valence factor) $H_2 = 2$
At NTP 22400mL of $H_2 = 1\text{mole}$ of H_2
112mL of $H_2 = \frac{1}{22400} \times 112 = 0.005\text{mole}$ of H_2
Moles of H_2produced $= \frac{I \times t}{96500} \times$ molar ratio
$0.005 = \frac{I \times 965}{96500} \times \frac{1\text{mole of } H_2}{2\text{mole of } e^{-}}\ I = 1.0\text{A}$

Question111

Given
$$E^{0}_{Cl_2/Cl^{-}} = 1.36V, E^{\circ}_{Cr^{3+}/Cr} = -0.74V$$
$$E_{Cr_2/O_7^{2-}/Cr^{3+}} = 1.33V, E^{0}_{MnO\bar{4}^{-}Mn^{2+}} = 1.51V$$
Among the following, the strongest reducing agent is
[2017]

Options:

A. Cr

B. Mn^{2+}

C. Cr^{3+}

D. Cl

Answer: A

Solution:

Solution:

$E_{MnO_4^- / Mn^{2+}}^{\circ} = 1.51V$

$E_{Cl_2 / Cl^-}^{\circ} = 1.36V$

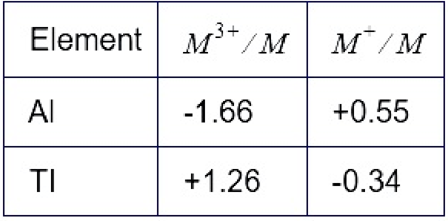
$E_{C_2O_7^{2-} / Cr^{3+}} = 1.33V$

$E_{Cr^{3+} / Cr}^{\circ} = -0.74$

Since Cr^{3+} is having least reduction potential, so Cris the best reducing agent.

Question112

Consider the following standard electrode potentials (E° in volts) in aqueous solution:



Element	M^{3+}/M	M^{+}/M
Al	-1.66	+0.55
Tl	+1.26	-0.34

Based on these data, which of the following statements is correct? [Online April 8, 2017]

Options:

A. Tl^{+} is more stable than Al^{3+}

B. Al^{+} is more stable than Al^{3+}

C. Tl^{+} is more stable than Al^{+}

D. Tl^{3+} is more stable than Al^{3+}

Answer: C

Solution:

Solution:

(i) $\underset{\text{more stable}}{Al^{3+}} \xrightarrow{E^{\circ} = -1.66} Al \xleftarrow{E^{\circ} = +0.55} \underset{\text{Less stable}}{Al^{+}}$

(ii) $\underset{\text{Less stable}}{Tl^{3+}} \xrightarrow{E^{\circ} = +1.26} Tl \xleftarrow{E^{\circ} = -0.34} \underset{\text{more stable}}{Tl^{+}}$

Tl^{+} has negative electrode potential ($E^{\circ} = -0.34$) means, it does not prefer to convert into Tl but reverse must be preferred that's why it is more stable than Tl^{3+} ($E^{\circ} = +1.26$). In Al, Al^{3+} is more stable ($E^{\circ} = -1.66$) than Al^{+} ($E^{\circ} = +0.55$) and also from Tl^{+} due to more negative value of E°. Therefore, by comparison it confirms that Tl^{+} is more stable than Al^{+}.

Question113

What is the standard reduction potential (E°) for $Fc^{3+} \rightarrow Fe$?
Given that :
$Fe^{2+} + 2e^{-} \rightarrow Fe; E^{\circ}Fe^{2+} / Fe = -0.47V$
$Fe^{3+} + e^{-} \rightarrow Fe^{2+}; E^{\circ}Fe^{3+} / Fe^{2+} = +0.77V$
[Online April 8, 2017]

Options:

A. $-0.057V$

B. $+0.057V$

C. $+0.30V$

D. $-0.30V$

Answer: A

Solution:

Solution:
$\Delta G^{\circ} = -nFE^{\circ}$
(i) $Fe^{2+} + 2e^{-} \longrightarrow Fe; \quad E^{0} = -0.47V;$
(ii) $Fe^{3+} + e^{-} \longrightarrow Fe^{2+}; E^{0} = +0.77V$:
(iii) $Fe^{3+} + 3e^{-} \longrightarrow Fe$
(i) $\Delta G^{0} = -nFE^{0} = -2(-0.47)F = 0.94F$
(ii)$\Delta G^{0} = -nFE^{0} = -1(+0.77)F = -0.77F$
(iii) On adding : $\Delta G^{0} = +0.17F$
$\Delta G^{0} = -nFE^{0}$
E^{0} for $(Fe^{3+} \longrightarrow Fe) = \dfrac{\Delta G^{0}}{-nF} = \dfrac{0.17F}{-3F} = -0.057V$

Question114

To find the standard potential of M^{3+} / M electrode, the following cell is constituted : $Pt / M \ / M^{3+}(0.001 mol L^{-1}) \backslash / Ag^{+}(0.01 mol L^{-1}) / Ag$

The emf of the cell is found to be 0.421 volt at $298K$. The standard potential of half reaction $M^{3+} + 3e \longrightarrow M$ at $298K$ will be:
(Given $E^{\circ}_{Ag^{+} / Ag}$ at $298K = 0.80V olt$)
[Online April 9, 2017]

Options:

A. 0.38 Volt

B. 0.32 Volt

C. 1.28 Volt

D. 0.66 Volt

Answer: B

Solution:

Cell reaction:

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

$$E = E^{0} - \frac{0.0591}{n} \log \frac{[\text{ Reduced state }]}{[\text{ Oxidised state }]}$$

$$0.421 = E^{0} - \frac{0.0591}{3} \log \frac{0.001}{(0.01)^{3}}$$

$$E^{0} = 0.48$$

$$E^{0} = E_{Ag^{+}/Ag}^{0} - E_{M^{3+}/M}^{0}$$

$$E_{M^{3+}/M}^{0} = 0.8V - 0.48V = 0.32\text{volt}$$

Question115

What will occur if a block of copper metal is dropped into a beaker containing a solution of 1M $ZnSO_4$?

[Online April 9, 2016]

Options:

A. The copper metal will dissolve with evolution of oxygen gas

B. The copper metal will dissolve with evolution of hydrogen gas

C. Noreaction will occur

D. The copper metal will dissolve and zinc metal will be deposited.

Answer: C

Solution:

Solution:
No reaction will occur as the Zn is placed above Cu in electrochemical series. Hence there will be no $_2$ displacement reaction.

Question116

Galvanization is applying a coating of:

[2016]

Options:

A. Cu

B. Zn

C. Pb

D. Cr

Answer: B

Solution:

Solution:
Galvanization is the process by which zinc is coated over corrosive (easily rusted) metals to prevent them from corrosion.

Question117

Identify the correct statement:
[Online April 10,2016]

Options:

A. Corrosion of iron can be minimized by forming a contact with another metal with a higher reduction potential

B. Iron corrodes in oxygen free water

C. Corrosion of iron can be minimized by forming an impermeable barrier at its surface

D. Iron corrodes more rapidly in salt water because its electrochemical potential is higher

Answer: C

Solution:

Solution:
Corrosion of iron can be minimized by forming an impermeable barrier at its surface.

Question118

Two Faraday of electricity is passed through a solution of $CuSO_4$. The mass of copper deposited at the cathode is
(at. mass of Cu = 63.5amu)
[2015]

Options:

A. 2g

B. 127g

C. 0g

D. 63.5g

Answer: D

Solution:

Solution:
$Cu^{2+} + 2e^{-} \longrightarrow Cu$
2F i.e. $2 \times 96500C$ deposit Cu = 1mol
$= 63.5g$

Question119

A variable, opposite external potential (E_{ext}) is applied to the cell $Zn | Zn^{2+}(1M) | Cu^{2+}(1M) | Cu$, of potential $1.1V$ When $E_{cxt} < 1.1V$ and $E_{cxt} > 1.1V$, respectively electrons flow from:
[Online April 10, 2015]

Options:

A. anode to cathode in both cases

B. cathode to anode and anode to cathode

C. anode to cathode and cathode to anode

D. cathode to anode in both cases

Answer: C

Solution:

Solution:
EMF of galvanic cell $= 1.1$ volt
If $E_{ext} < EMF$ then electrons flow steadily from anode to cathode, while if $E_{ext} > EMF$ then electrons flow from cathode to anode as polarity is changed.

Question120

At $298K$, the standard reduction potentials are $1.51V$ for $MnO_4^- | Mn^{2+}$, $1.36V$ for $Cl_2 | Cl^-$, $1.07V$ for $Br_2 | Br^-$ and $0.54V$ for $I_2 | I$. At $pH = 3$, permanganate is expected to oxidize : $\left(\frac{RT}{F} = 0.059V \right)$
[Online April 11, 2015]

Options:

A. Cl , Br and I

B. Br and I^-

C. Cl^- and Br^-

D. I^- only

Answer: B

Solution:

Solution:
$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$

$$E = 1.51 - \frac{0.059}{5} \log \frac{[Mn^{2+}]}{[MnO_4^-][H^+]^8}$$

Taking Mn^{2+} and MnO_4^- in standard state i.e. $1M$,

$E = 1.51 - \frac{0.059}{5} \times 8 \log \frac{1}{[H^+]}$

$= 1.51 - \frac{0.059}{5} \times 8 \times 3 = 1.2268V$

Hence at this pH, MnO_4^- will oxidise only Br^- and I^- as SRP of Cl_2 / Cl^- is $1.36V$ which is greater than that for MnO_4^- / Mn^{2+}

Question121

Resistance of $0.2M$ solution of an electrolyte is 50Ω. The specific conductance of the solution is $1.4Sm^{-1}$. The resistance of $0.5M$ solution of the same electrolyte is $280\ \Omega$. The molar conductivity of $0.5M$ solution of theelectrolyte in Sm^2mol^{-1} is:
[2014]

Options:

A. 5×10^{-4}

B. 5×10^{-3}

C. 5×10^{3}

D. 5×10^{2}

Answer: A

Solution:

Solution:
Given for $0.2M$ solution
$R = 50\Omega$
$\kappa = 1.4Sm^{-1} = 1.4 \times 10^{-2}Scm^{-1}$
Now, $R = \rho \frac{l}{a} = \frac{1}{\kappa} \times \frac{l}{a}$
$\Rightarrow \frac{l}{a} = R \times \kappa = 50 \times 1.4 \times 10^{-2}cm^{-1}$
For 0.5 M solution
$R = 280\Omega$
$\kappa = ?$
$\frac{l}{a} = 50 \times 1.4 \times 10^{-2}cm^{-1}$
$\Rightarrow R = \rho \frac{l}{a} = \frac{1}{\kappa} \times \frac{l}{a}$
$\Rightarrow \kappa = \frac{1}{280} \times 50 \times 1.4 \times 10^{-2}$
$= \frac{1}{280} \times 70 \times 10^{-2}$
$= 2.5 \times 10^{-3}Scm^{-1}$
Now, $\Lambda_m = \frac{\kappa \times 1000}{M}$
$= \frac{2.5 \times 10^{-3}Scm^{-1} \times 1000cm^3 / L}{0.5mol / L}$
$= 5Scm^2mol^{-1} = 5 \times 10^{-4}Sm^2mol^{-1}$

Question122

The equivalent conductance of $NaCl$ at concentration C and at infinite dilution are λ_C and λ_∞, respectively. The correct relationship between λ_C and λ_∞ is given as:
(Where the constant B is positive)
[2014]

Options:

A. $\lambda_C = \lambda_\infty + (B)C$

B. $\lambda_C = \lambda_\infty - (B)C$

C. $\lambda_C = \lambda_\infty - (B)\sqrt{C}$

D. $\lambda_C = \lambda_\infty + (B)\sqrt{C}$

Answer: C

Solution:

Solution:
According to Debye Huckle onsager equation,
$\lambda_C = \lambda_\infty - B\sqrt{C}$

Question123

The standard electrode potentials $\left(E^\circ_{M^+/M}\right)$ of four metals A, B, C and D are $-1.2V$, $0.6V$, $0.85V$ and $-0.76V$, respectively. The sequence of deposition of metals on applying potential is:
[Online April 9, 2014]

Options:

A. A, C, B, D

B. B, D, C, A

C. C, B, D, A

D. D, A, B, C

Answer: C

Solution:

Solution:
As the value of reduction potential decreases the reducing power increases i.e.
$C < B < D < A$
$(0.85) \cdot (0.6) \cdot (-0.76)(-1.2)$

Question124

A current of 10.0A flows for 2.00h through an electrolytic cell containing a molten salt of metal X. This results in the decomposition of 0.250 mol of metal X at the cathode. The oxidation state of X in the molten salt is: (F = 96, 500C)
[Online April 9, 2014]

Options:

A. 1+

B. 2+

C. 3+

D. 4

Answer: C

Solution:

Solution:
According to Faraday's first law of electrolysis

$$W = \frac{E \times i \times t}{96500}$$

Where E = equivalent weight

$$= \frac{\text{mol. mass of metal (M)}}{\text{oxidation state of metal (x)}}$$

Substituting the value in the formula

$$W = \frac{M}{x} \times \frac{i \times t}{96500}$$

or $x = \frac{M}{W} \times \frac{i \times t}{96500} = \frac{10 \times 2 \times 60 \times 60}{96500 \times 0.250} = 3$

$\left[\text{Given : no. of moles } = \frac{W}{M} = 0.250\right]$

Hence oxidation state of metal is (+3)

Question125

Given

$$Fe^{3+}(aq) + e^{-} \rightarrow Fe^{2+}(aq); E^{\circ} = +0.77V$$

$$Al^{3+}(aq) + 3e^{-} \rightarrow Al(s); E^{\circ} = -1.66V$$

$$Br_2(aq) + 2e \rightarrow 2Br; E^{\circ} = +1.09V$$

Considering the electrode potentials, which of the following represents the correct order of reducing power?
[Online April 11, 2014]

Options:

A. $Fe^{2+} < Al < Br^{-}$

B. $Br^{-} < Fe^{2+} < Al$

C. $Al < Br^{-} < Fe^{2+}$

D. $Al < Fe^{2+} < Br$

Answer: D

Solution:

Solution:
Reducing character decreases down the series. Hence the correct order is
$Al < Fe^{2+} < Br^{-}$

Question126

How many electrons would be required to deposit 6.35g of copper at the cathode during the electrolysis of an aqueous solution of copper sulphate? (Atomic mass of copper $= 63.5u$, $N_A =$ Avogadro's constant) [Online April 12, 2014]

Options:

A. $\frac{N_A}{20}$

B. $\frac{N_A}{10}$

C. $\frac{N_A}{5}$

D. $\frac{N_A}{2}$

Answer: C

Solution:

Solution:
$Cu \longrightarrow Cu^{2+} + 2e^{-}$
i.e, to deposit 1 mole of Cu at cathode from $Cu^{2+}SO_4{}^{2-}$ solution $=$ 2 moles of electrons are required
i.e, To deposit 6.35g $= \frac{6.35}{63.5} \times 2 = \frac{2}{10} = \frac{1}{5}$mol
Thus total no. of electronsrequired $= \frac{N_A}{5}$

Question127

Given below are the half-cell reactions:
$Mn^{2+} + 2e^{-} \rightarrow Mn; E^{\circ} = -1.18V$
$2(Mn^{3+} + e^{-} \rightarrow Mn^{2+}); E^{\circ} = +1.51V$
The E for $3Mn^{2+} \rightarrow Mn + 2Mn^{3+}$ will be:
[2014]

Options:

A. $-2.69\mathrm{V}$; the reaction will not occur

B. $-2.69\mathrm{V}$; the reaction will occur

C. $-0.33\mathrm{V}$; the reaction will not occur

D. $-0.33\mathrm{V}$; the reaction will occur

Answer: A

Solution:

Solution:
(a) $Mn^{2+} + 2e^- \rightarrow Mn;\ E^\circ = -1.18\mathrm{V}$;...(i)
(b) $Mn^{3+} + e^- \rightarrow Mn^{2+};\ E^0 = -1.51\mathrm{V}$; ... (ii)
Now multiplying equation (ii) by two and subtracting from equation (i)
$3Mn^{2+} \rightarrow Mn^{+} + 2Mn^{3+}$
$E^0 \ = E_{Ox.} + E_{Red}$
$= -1.18 + (-1.51) = -2.69\mathrm{V}$
(-ve value of EMF (i.e. $\Delta G = +ve$) shows that the reaction is non-spontaneous)

Question128

Electrode potentials (E°) are given below:

$Cu^+ / Cu = +0.52\mathrm{V}$
$Fe^{3+} / Fe^{2+} = +0.77\mathrm{V}$
$\frac{1}{2}I_2(s) / I^- = +0.54\mathrm{V}$
$Ag^+ / Ag = +0.88\mathrm{V}$
Based on the above potentials, strongest oxidizing agent will be:
[Online April 9, 2013]

Options:

A. Cu^+

B. Fe^{3+}

C. Ag^+

D. I_2

Answer: C

Solution:

Solution:
Higher the value of reduction potential stronger will be the oxidising agent, hence based on the given values Ag^+ will be strongest oxidizing agent.

Question129

Given:

$E_{\frac{1}{2}Cl_2/Cl^-} = 1.36V$, $E^\circ_{Cr^{3+}/Cr} = -0.74V$

$E^\circ Cr_2O_7^{2-}/Cr^{3+} = 1.33V$, $E^\circ MnO_4^-/Mn^{2+} = 1.51V$

The correct order of reducing power of the species(Cr, Cr^{3+}, Mn^{2+} and Cl^-) will be:
[Online April 23, 2013]

Options:

A. $Mn^{2+} < Cl^- < Cr^{3+} < Cr$

B. $Mn^{2+} < Cl^{3+} < Cl^- < Cr$

C. $Cr^{3+} < Cl^- < Mn^{2+} < Cr$

D. $Cr^{3+} < Cl^- < Cr < Mn^{2+}$

Answer: A

Solution:

Solution:
Lower the value of reduction potential higher will be reducing power hence the correct order will be
$Mn^{2+} < Cl^- < Cr^{3+} < Cr$

Question130

A solution of copper sulphate ($CuSO_4$) is electrolysed for 10 minutes with a current of 1.5 amperes. The mass of copper deposited at the cathode (at. mass of Cu = 63u) is:
[Online April 25, 2013]

Options:

A. 0.3892g

B. 0.2938g

C. 0.2398g

D. 0.3928g

Answer: B

Solution:

Solution:
$W = Zit$
where $Z =$ Electrochemical equivalent
Eq. wt. of copper $= \frac{63}{2} = 31.5$
$Z = \frac{31.5}{96500}$
$W = Zit = \frac{31.5}{96500} \times 1.5 \times 10 \times 60 = 0.2938g$

Question131

Given:

$$\mathbf{E^{\circ}_{Cr^{3+}/Cr} = -0.74V ;\ E^{0}_{MnO_4^-/Mn^{2+}} = 1.51V}$$

$$\mathbf{E_{Cr_2O_7^{2-}/Cr^{3+}} = 1.33V ;\ E^{\circ}_{Cl/Cl^-} = 1.36V}$$

Based on the data given above, strongest oxidising agentwill be: [2013]

Options:

A. Cl

B. Cr^{3+}

C. Mn^{2+}

D. MnO_4^-

Answer: D

Solution:

Solution:
Higher the value of standard reduction potential, stronger is the oxidising agent, hence MnO_4^- is the strongest oxidising agent.

Question132

The standard reduction potentials for Zn^{2+}/Zn, Ni^{2+}/Ni and Fe^{2+}/Fe are -0.76, -0.23 and $-0.44V$ respectively.
The reaction $X + Y^{2+} \longrightarrow X^{2+} + Y$ will be spontaneous when : [2012]

Options:

A. $X = Ni, Y = Fe$

B. $X = Ni, Y = Zn$

C. $X = Fe, Y = Zn$

D. $X = Zn, Y = Ni$

Answer: D

Solution:

Solution:
For a spontaneous reaction ΔG° must be -ve
Since $\Delta G^{\circ} = -nFE^{\circ}$

Hence for ΔG° to be − ve, ΔE° has to be positive. Which is possible when $X = Zn, Y = Ni$
$Zn + Ni^{2+} \longrightarrow Zn^{2+} + Ni$
$E_{cell}^{\circ} = E_{Ni^{2+}/Ni}^{\circ} - E_{Zn^{2+}/Zn}^{\circ}$
$= -0.23 - (-0.76) = +0.53$ (positive)

Question133

A battery is constructed of Cr and $Na_2Cr_2O_7$. The unbalanced chemical equation when such a battery discharges is following:
$Na_2Cr_2O_7 + Cr + H^+ \rightarrow Cr^{3+} + H_2O + Na^+$
If one Faraday of electricityis passed through the batteryduring the charging, the number of moles of Cr^{3+} removed from the solution is
[Online May 7, 2012]

Options:

A. $\frac{4}{3}$

B. $\frac{1}{3}$

C. $\frac{3}{3}$

D. $\frac{2}{3}$

Answer: C

Solution:

Solution:
Reduction half reaction:
$Cr_2O_7^{2-} + 6e^- + 14H^+ \longrightarrow 2Cr^{3+} + 7H_2O$
Oxidation half reaction:
$Cr \longrightarrow Cr^{3+} + 3e^-$
Overall reaction :
$Cr_2O_7^{2-} + Cr + 14H^+ + 3e^- \longrightarrow 3Cr^{3+} + 7H_2O$
$3F$ of electricity $= 3$moles of Cr^{3+}
$1F$ of electricity $= \frac{3}{3}$ moles of Cr^{3+}

Question134

The standard potentials of Ag^+/Ag, $Hg_2^{2+}/2Hg$, Cu^{2+}/Cu and Mg^{2+}/Mg electrodes are 0.80, 0.79, 0.34 and $-2.37V$, respectively. An aqueous solution which contains one mole per litre of the salts of each of the four metals is electrolyzed. With increasing voltage, the correct sequence of deposition of the metals at the cathode is
[Online May 19, 2012]

Options:

A. Ag, Hg, Cu, Mg

B. Cu, Hg, Ag only

C. Ag, Hg, Cu only

D. Mg, Cu, Hg, Ag

Answer: C

Solution:

Solution:
More the reduction potential, more is the deposition of metals at cathode. Cation having E° value less than $-0.83V$ (reduction potential of H_2O) will not deposit from aqueous solution. Hence correct order of deposition of themetal at the cathode is
$Ag > Hg > Cu$

Question135

Given $E^\circ_{Cu^{2+}/Cu} = 0.34V, E^\circ_{Cu^{2+}/Cu^+} = 0.15V$ Standard electrode potential for the half cell Cu^+/Cu is [Online May 26,2012]

Options:

A. 0.38V

B. 0.53V

C. 0.19V

D. 0.49V

Answer: B

Solution:

$Cu^{2+} + e^- \longrightarrow Cu^+; E_1^\circ = 0.15V; \Delta G_1^\circ = -n_1E_1^\circ F$
$Cu^{2+} + 2e \longrightarrow Cu; E_2^\circ = 0.34V; \Delta G_2^\circ = -n_2E_2^\circ F$
On subracting eq.(i) from eq. (ii) we get
$Cu^+ + e^- \longrightarrow Cu; \quad \Delta G^\circ = \Delta G_2^\circ - \Delta G_1^\circ$
$-nE^\circ F = -(n_2E_2^\circ F - n_1E_1^0 F)$

$$E^\circ = \frac{n_2E_2^\circ F - n_1E_1^\circ F}{nF}$$

$$= \frac{2 \times 0.34 - 0.15}{1} = 0.53V$$

Question136

Resistance of $0.2M$ solution of an electrolyte is 50Ω. The specific

conductance of the solution is $1.3 \mathrm{Sm}^{-1}$. If resistance of the 0.4M solution of the same electrolyte is 260 Ω, its molar conductivity is : [2011RS]

Options:

A. $6.25 \times 10^{-4} \mathrm{Sm}^{2} \mathrm{mol}^{-}$

B. $625 \times 10^{-4} \mathrm{Sm}^{2} \mathrm{mol}^{-1}$

C. $62.5 \mathrm{Sm}^{2} \mathrm{mol}^{-1}$

D. $6250 \mathrm{Sm}^{2} \mathrm{mol}$

Answer: A

Solution:

$$k = \frac{1}{R} \times \frac{l}{a}$$
$$1.3 = \frac{1}{50} \times \frac{l}{a}$$
$$\frac{l}{a} = 65 \mathrm{m}^{-1}$$
$$\Lambda_m = \frac{k \times 1000}{\text{molarity}}$$

[molarity is in moles/litre but 1000 is used to convert litre into cm^3]

$$= \frac{\left(\frac{1}{260} \times 65 \mathrm{m}^{-1}\right) \times 1000 \mathrm{cm}^3}{0.4 \mathrm{mol}}$$
$$= \frac{65 \mathrm{m}^{-1}}{260 \times 0.4 \mathrm{mol}} \times \frac{1}{1000} \mathrm{m}^3$$
$$= 6.25 \times 10^{-4} \mathrm{Sm}^2 \mathrm{mol}^{-1}$$

Question137

The Gibbs energy for the decomposition of Al_2O_3 at 500 °C is as follows:
$\frac{2}{3}Al_2O_3 \rightarrow \frac{4}{3}Al + O_2, \Delta_rG = +966 \mathrm{kJ\ mol}^{-1}$
The potential difference needed for electrolytic reduction of Al_2O_3 at 500°C is at least
[2010]

Options:

A. 4.5V

B. 3.0V

C. 2.5V

D. 5.0V

Answer: C

Solution:

$\frac{2}{3}Al_2O_3 \rightarrow \frac{4}{3}Al + O_2$
The ionic reactions are
$\frac{2}{3} \times (2Al^{3+}) + 4e^{-} \rightarrow \frac{4}{3}Al$
$\frac{2}{3} \times (3O^{2-}) \rightarrow O_2 + 4e^{-}$
Thus, no. of electron transferred $\Rightarrow n = 4$
$\Delta G = -nFE$
or $E = \frac{\Delta G}{-nF} = \frac{966 \times 10^3}{4 \times 96500} = -2.5V$
∴ The potential difference needed for the reduction $= 2.5V$.

Question138

Given:

$E_{Fe^{3+}/Fe}^{\circ} = -0.036V$
$E^{\circ}Fe^{2+}/Fe = -0.439V$
The value of standard electrode potential for the change,
$Fe^{3+}(aq) + e^{-} \longrightarrow Fe^{2+}(aq)$ will be:
[2009]

Options:

A. $0.385V$

B. $0.770V$

C. $-0.270V$

D. $-0.072V$

Answer: B

Solution:

Given
$Fe^{3+} + 3e^{-} \rightarrow Fe$
$E_{Fe^{3+}/Fe}^{\circ} = -0.036V$ ...(i)
$Fe^{2+} + 2e^{-} \rightarrow Fe$
$E_{Fe^{2+}/Fe}^{\circ} = -0.439V$ ...(ii)
we have to calculate
$Fe^{3+} + e^{-} \rightarrow Fe^{2+}, \Delta G^{\circ} = ?$
To obtain this equationsubtract equ. (ii) from (i) we get
$Fe^{3+} + e^{-} \rightarrow Fe^{2+}$ ...(iii)
As we know that $\Delta G^{\circ} = -nFE^{\circ}$
Thus for reaction (iii)
$\Delta G_3 = \Delta G_1 - \Delta G_2$
$-nFE_3^{\circ} = -nFE_1^{\circ} - (-nFE_2^{\circ})$
$-nFE_3^{\circ} = nFE_2^{\circ} - nFE_1^{\circ}$
$-1FE_3^{\circ} = 2 \times 0.439F - 3 \times 0.036F$
$-1FE_3^{\circ} = 0.770F$
$\therefore E_3^{\circ} = -0.770V$

Question139

In a fuel cell methanol is used as fuel and oxygen gas is used as an oxidizer. The reaction is
$CH_3OH(l) + 3/2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$
At 298K standard Gibb's energies of formation for $CH_3OH(l)$, $H_2O(l)$ and and $CO_2(g)$ are $-166.2 - 237.2$ and -394.4 kJ mol^{-1} respectively. If standard enthalpy of combustion of methonal is -726 kJ mol^{-1}, efficiency of the fuel cell will be:
[2009]

Options:

A. 87%

B. 90%

C. 97%

D. 80%

Answer: C

Solution:

$CH_3OH(l) + \frac{3}{2}O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$
$\Delta G_r = \Delta G_f(CO_2, g) + 2\,\Delta\,G_f(H_2O, 1) -$
$\Delta G_f(CH_3OH, 1) - \frac{3}{2}\,\Delta\,G_f(O_2, g)$
$= -394.4 + 2(-237.2) - (-166.2) - 0$
$= -394.4 - 474.4 + 166.2 = -702.6kJ$
% efficiency $= \frac{702.6}{726} \times 100 = 97\%$

Question140

Given $E^\circ Cr^{3+}/Cr = -0.72V$
$E^\circ Fe^{2+}/Fe = -0.42V$.
The potential for the cell
$Cr\,|\,Cr^{3+}(0.1M)\,|\,Fe^{2+}(0.01M)\,|\,Fe$ is
[2008]

Options:

A. 0.26V

B. 0.336V

C. -0.339

D. 0.26V

Answer: D

Solution:

Solution:
From the given representation of the cell, E_{cell} can be found as follows.

$$E_{cell} = E_{Fe^{2+}/Fe}^{o} - E_{Cr^{3+}/Cr}^{o} - \frac{0.059}{6}\log\frac{[Cr^{3+}]^2}{[Fe^{2+}]^3}$$

[Nernst -Equ.]

$$= -0.42 - (-0.72) - \frac{0.059}{6}\log\frac{(0.1)^2}{(0.01)^3}$$

$$= -0.42 + 0.72 - \frac{0.059}{6}\log\frac{0.1 \times 0.1}{0.01 \times 0.01 \times 0.01}$$

$$= 0.3 - \frac{0.059}{6}\log\frac{10^{-2}}{10^{-6}} = 0.3 - \frac{0.059}{6} \times 4$$

$$= 0.30 - 0.0393 = 0.26V$$

Question141

The equivalent conductances of two strong electrolytes at infinite dilution in H_2O (where ions move freely through a solution) at 25°C are given below:

$$\Lambda_{CH_3COONa} = 91.0 Scm^2/ \text{ equiv.}$$

$$\Lambda_{HCl}^{\circ} = 426.2 Scm^2/ \text{ equiv.}$$

What additional information\/ quantity one needs to calculate Λ° of an aqueous solution of acetic acid?
[2007]

Options:

A. Λ of chloroacetic acid ($ClCH_2COOH$)

B. Λ of $NaCl$

C. Λ of CH_3COOK

D. the limiting equivalent coductance of $H^+ \left(\lambda_{H^+}^{\circ}\right)$.

Answer: B

Solution:

Solution:
Note: According to Kohlrausch's law, equivalent conductivity of weak electrolyte, acetic acid (CH_3COOH) can be calculated as follows:

$$\Lambda_{CH_3COOH}^{\circ} = \left(\Lambda_{CH_3COONa}^{\circ} + \Lambda_{HCl}^{\circ}\right) - \Lambda_{NaCl}^{\circ}$$

∴ Value of Λ_{NaCl}° should also be known for calculating value of $\Lambda_{CH_3COOH}^{\circ}$

Question142

The cell, $Zn \mid Zn^{2+}(1M) \mid Cu^{2+}(1M) \mid Cu(E_{cell}{}^{0} = 1.10V)$ was allowed to be completely discharged at 298K.

The relative concentration of Zn^{2+} to Cu^{2+} $\left(\frac{[Zn^{2+}]}{[Cu^{2+}]}\right)$ is

[2007]

Options:

A. 9.65×10^4

B. antilog (24.08)

C. 37.3

D. 10^{373}.

Answer: D

Solution:

Solution:

$E_{cell} = 0$; when cell is completely discharged.

$$E_{cell} = E_{cell}{}^{\circ} - \frac{0.059}{2}\log\left(\frac{[Zn^{2+}]}{[Cu^{2+}]}\right)$$

$$\text{or } 0 = 1.1 - \frac{0.059}{2}\log\left(\frac{[Zn^{2+}]}{[Cu^{2+}]}\right)$$

$$\log\left(\frac{[Zn^{2+}]}{[Cu^{2+}]}\right) = \frac{2 \times 1.1}{0.059} = 37.3$$

$$\therefore \left(\frac{[Zn^{2+}]}{[Cu^{2+}]}\right) = 10^{37.3}$$

Question143

The molar conductivities $\Lambda_{NaOAc}{}^{o}$ and $\Lambda_{HCl}{}^{o}$ at infinite dilution in water at 25°C are 91.0 and 426.2Scm^2 / mol respectively. To calculate $\Lambda_{HOAc}{}^{\circ}$, the additional value required is
[2006]

Options:

A. $\Lambda_{NaOH}{}^{0}$

B. $\Lambda_{NaCl}{}^{o}$

C. $\Lambda_{H_2O}{}^{O}$

D. $\Lambda_{KCl}{}^{0}$

Answer: B

Solution:

$\Lambda_{CH_3COOH}{}^{\circ}$ is given by the following equation

$\Lambda_{CH_3COOH}{}^{\circ} = \left(\Lambda_{CH_3COONa}{}^{\circ} + \Lambda_{HCl}{}^{\circ}\right) - \left(\Lambda_{NaCl}{}^{\circ}\right)$

Hence $\Lambda_{NaCl}{}^{\circ}$ is required.

Question144

Resistance of a conductivity cell filled with a solution of an electrolyte of concentration $0.1M$ is 100Ω. The conductivity of this solution is $1.29Sm^{-1}$. Resistance of the same cell when filled with $0.2M$ of the same solution is 520Ω. The molar conductivity of $0.2M$ solution of electrolyte will be

[2006]

Options:

A. $1.24 \times 10^{-4}Sm^{2}mol^{-1}$

B. $12.4 \times 10^{-4}Sm^{2}mol^{-1}$

C. $124 \times 10^{-4}Sm^{2}mol^{-1}$

D. $1240 \times 10^{-4}Sm^{2}mol^{-1}$

Answer: B

Solution:

$R = 100\Omega, \kappa = \frac{1}{R}\left(\frac{l}{a}\right)$

$\frac{l}{a}\left(\text{ cell constant }\right) = 1.29 \times 100m^{-1}$

Given, $R = 520\Omega$

$C = 0.2M = 0.2 \times 1000mol\,m^{-3}$

Λ_m (molar conductivity) $= ?$

$\Lambda_m = K \times V$ (K can be calculated as cell constantis known.)

$K = \frac{1}{R}\left(\frac{l}{a}\right)$

Hence,

$K = \frac{129m^{-1}}{520\Omega} = 0.248Sm^{-1}$

$\Lambda_m = \frac{K}{C} = \frac{0.248Sm^{-1}}{0.2 \times 1000mol\,m^{-3}}$

$= 12.4 \times 10^{-4}Sm^{2}mol^{-1}$

Question145

The highest electrical conductivity of the following aqueous solutions is of

[2005]

Options:

A. 0.1M difluoroacetic acid

B. 0.1M fluoroacetic acid

C. 0.1M chloroacetic acid

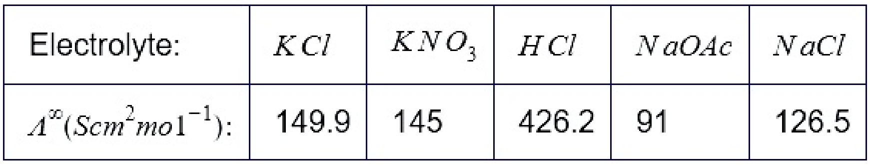
D. 0.1M acetic acid

Answer: A

Solution:

Solution:
Difluoro acetic acid being strongest acid will furnish maximum number of ions showing highest electrical conductivity.
The decreasing acidic strength of the carboxylic acids given is
difluoro acetic acid > fluoro acetic acid > chloro acetic acid> acetic acid.

Question146



Electrolyte:	KCl	KNO_3	HCl	$NaOAc$	$NaCl$
$\Lambda^{\infty}(Scm^2mol^{-1})$:	149.9	145	426.2	91	126.5

Calculate $\Lambda_{HOAc}{}^{\infty}$ using appropriate molar conductance's of the electrolytes listed above at infinite dilution in H_2O at 25°C [2005]

Options:

A. 217.5

B. 390.7

C. 552.7

D. 517.2

Answer: B

Solution:

Solution:
$\Lambda_{HCl}{}^{\infty} = 426.2 \ldots$ (i)
$\Lambda_{ACONa}{}^{\infty} = 91.0 \ldots$ (ii)
$\Lambda_{NaCl}{}^{\infty} = 126.5 \ldots$ (iii)
$\Lambda_{AcOH}{}^{\infty} = \text{(i)} + \text{(ii)} - \text{(iii)}$
$= [426.2 + 91.0 - 126.5] = 390.7$

Question147

Aluminium oxide may be electrolysed at 1000°C to furnish aluminium

metal (At. Mass $= 27$ amu; 1 Faraday $= 96, 500$ Coulombs). The cathode reaction is- $Al^{3+} + 3e^{-} \rightarrow Al^{\circ}$ Toprepare 5.12kg of aluminium metal by this methodwe require
[2005]

Options:

A. 5.49×10^{1}C of electricity

B. 5.49×10^{4}C of electricity

C. 1.83×10^{7}C of electricity

D. 5.49×10^{7}C of electricity

Answer: D

Solution:

Solution:
1 mole of $e^{-} = 1F = 96500C$
27g of Al is deposited by $3 \times 96500C$
5120g of Al will be deposited by
$= \frac{3 \times 96500 \times 5120}{27} = 5.49 \times 10^{7}C$
We know, $Q = \frac{mFz}{M}$
$\therefore Q = \frac{5120 \times 96500 \times 3}{27} = 5.49 \times 10^{7}C$

Question148

For a spontancous reaction the ΔG, equilibrium constant (K) and E_{cell}° will be respectively
[2005]

Options:

A. $-ve, >1, -ve$

B. $-ve, <1, -ve$

C. $+ve, >1, -ve$

D. $-ve, >1, +ve$

Answer: D

Solution:

Solution:
Note: For spontaneous reaction ΔG should be negative. Equilibrium constant should be more than one ($\Delta G = -2.303RT \log K_c$, If $K_c = 1$ then
$\Delta G = 0$; If $K_c > 1$ then $\Delta G = -ve$).
Again $\Delta G = -nFE_{cell}^{\circ}$
E_{cell}° must be + ve to have ΔG – ve.

Question149

The limiting molar conductivities Λ° for $NaCl$, KBr and KCl are 126,152 and 150Scm^2mol^{-1} respectively. The Λ for NaBr is [2004]

Options:

A. $278Scm^2mol^{-1}$

B. $176Scm^2mol^{-1}$

C. $128Scm^2mol^{-1}$

D. $302Scm^2mol^{-1}$

Answer: C

Solution:

Solution:
$\Lambda^{\circ}NaCl = \lambda^{\circ}Na^{+} + \lambda_{Cl}$ . . . (i)
$\Lambda_{KBr}^{\circ} = \lambda_{K^+}^{\circ} + \lambda_{Br^-}$ . . . (ii)
$\Lambda_{KCl}^{\circ} = \lambda_{K^+}^{\circ} + \lambda_{Cl^-}$ . . . (iii)
operating (i) + (ii) − (iii)
$\Lambda_{NaBr}^{\circ} = \lambda_{Na^+}^{\circ} + \lambda_{Br^-}^{-}$
$= 126 + 152 - 150 = 128Scm^2mol^{-1}$

Question150

Consider the following
E° values $E_{Fe^{3+}/Fe^{2+}} = +0.77V$; $E_{Sn^{2+}/Sn}^{\circ} = -0.14V$
Under standard conditions the potential for the reaction
$Sn(s) + 2Fe^{3+}(aq) \rightarrow 2Fe^{2+}(aq) + Sn^{2+}(aq)$is
[2004]

Options:

A. 0.91V

B. 1.40V

C. 1.68V

D. 0.63V

Answer: A

Solution:

$Fe^{3+} + e^- \rightarrow Fe^{2+}$; $Sn^{2+} + 2e^- \rightarrow Sn(s)$;
for $Sn(s) + 2Fe^{3+}(aq) \rightarrow 2Fe^{2+}(aq) + Sn^{2+}(aq)$
∴ Standard potential for the given reaction or
$E_{cell}^{\circ} = E_{Sn/Sn^{2+}}^{\circ} + E_{Fe^{3+}/Fe^{2+}}$
$= 0.14 + 0.77 = 0.91V$

Question151

The $E^0 M^{3+}/M^{2+}$ values for Cr, Mn, Fe and Coare $-0.4l$, $+1.57$, $+0.77$ and $+1.97V$ respectively. For which one of these metals the change in oxidation state from $+2$ to $+3$ is easiest?
[2004]

Options:

A. Fe

B. Mn

C. G

D. Co

Answer: C

Solution:

Solution:
The given values show that Cr has maximum oxidation potental, therefore its oxidation will be easiest. (Change the sign to get the oxidation values)

Question152

The standard e.m.f. of a cell involving one electron change is found to be $0.591V$ at $25^{\circ}C$. The equilibrium constant of the reaction is $(F = 96,500 C mol^{-1}; R = 8.314 J K^{-1} mol^{-1})$
[2004]

Options:

A. 1.0×10^{10}

B. 1.0×10^{5}

C. 1.0×10^{1}

D. 1.0×10^{30}

Answer: A

Solution:

(a) $E_{cell} = E_{cell}^{\circ} - \frac{0.059}{n} \log K_c$

or $0 = 0.591 - \frac{0.0591}{1} \log K_c$

or $\log K_c = \frac{0.591}{0.0591} = 10$ or $K_c = 1 \times 10^{10}$

Question153

In a cell that utilises the reaction $Zn(s) + 2H^{+}(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$ addition of H_2SO_4 to cathode compartment, will [2004]

Options:

A. increase the E and shift equilibrium to the right

B. lower the E and shift equilibrium to the right

C. lower the E and shift equilibrium to the left

D. increase the E and shift equilibrium to the left

Answer: A

Solution:

Solution:
$Zn(s) + 2H^{+}(aq) \rightleftharpoons Zn^{2+}(aq) + H_2(g)$

$$E_{cell} = E_{cell}^{\circ} - \frac{0.059}{2} \log \frac{[Zn^{2+}][H_2]}{[H^{+}]^2}$$

Addition of H_2SO_4 will increase $[H^{+}]$ and E_{cell} will also increase and thus the equilibrium shifts towards RHS.

Question154

In a hydrogen-oxygen fuel cell, combustion of hydrogen occurs to [2004]

Options:

A. produce high purity water

B. create potential difference between two electrodes

C. generate heat

D. remove adsorbed oxygen from elctrode surfaces

Answer: B

Solution:

In $H_2 - O_2$ fuel cell, the combustion of H_2 occurs to create potential difference between the two electrodes

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Question155

Standard reduction electrode potentials of three metals A, B & C are respectively $+0.5V$, $-3.0V$ $-1.2V$. The reducing powers of these metals are [2003]

Options:

A. $A > B > C$

B. $C > B > A$

C. $A > C > B$

D. $B > C > A$

Answer: D

Solution:

Solution:
Note: The higher the negative value of reduction potential, the more is the reducing power.
Hence $B > C > A$.

Question156

When during electrolysis of a solution of $AgNO_3$ 9650 coulombs of charge pass through the electroplating bath, the mass of silver deposited on the cathode will be [2003]

Options:

A. 10.8g

B. 21.6g

C. 108g

D. 1.08g

Answer: A

Solution:

Solution:
When 96500 coulomb of electricity is passed through the electroplating bath the amount of Ag deposited $= 108g$
$\therefore$ when 9650 coulomb of electricity is passed deposited Ag.

$= \frac{108}{96500} \times 9650 = 10.8g$

Question157

For a cell reaction involving a two-clectron change, the standard e.m.f. of the cell is found to be $0.295V$ at $25^{\circ}C$. The equilibrium constant of the reaction at $25^{\circ}C$ will be
[2003]

Options:

A. 29.5×10^{-2}

B. 10

C. 1×10^{10}

D. 1×10^{-10}

Answer: C

Solution:

Solution:
The equilibrium constant is related to the standard emf of cell by the expression

$\log K = E_{cell}^{\circ} \times \frac{n}{0.059} = 0.295 \times \frac{2}{0.059}$

$\log K = \frac{590}{59} = 10$ or $K = 1 \times 10^{10}$

Question158

For the redox reaction :
[2003]

$Zn(s) + Cu^{2+}(0.1M) \rightarrow Zn^{2+}(1M) + Cu(s)$
taking place in a cell, E_{cell}° is 1.10 volt. E_{cell} for the cell will be

$\left(2.303\frac{RT}{F} = 0.0591\right)$

Options:

A. 1.80 volt

B. 1.07 volt

C. 0.82 volt

D. 2.14 volt

Answer: B

Solution:

Solution:

$$E_{cell} = E_{cell}^{\circ} + \frac{0.059}{n}\log\frac{[Cu^{2+}]}{[Zn^{2+}]}$$
$$= 1.10 + \frac{0.059}{2}\log[0.1]$$
$$= 1.10 - 0.0295 = 1.07V$$

Question159

Conductivity (unit Siemen's S) is directly proportional to area of the vessel and the concentration of the solution in it and is inversely proportional to the length of the vessel then the unit of the constant of proportionality is [2002]

Options:

A. $Smmol^{-1}$

B. Sm^2mol^{-1}

C. $S^{-2}m^2mol$

D. $S^2m^2mol^{-2}$.

Answer: B

Solution:

Solution:

Given conductivity $\propto \frac{\text{area} \times \text{conc.}}{\text{length}}$

$$= \frac{K\,m^2mol}{m \times m^3}$$
$$\therefore K = Sm^2mol^{-1}$$

Question160

Which of the following reaction is possible at anode? [2002]

Options:

A. $2Cr^{3+} + 7H_2O \rightarrow Cr_2O_7^{2-} + 14H^+$ [2002]

B. $F_2 \rightarrow 2F$

C. $(1/2)O_2 + 2H^+ \rightarrow H_2O$

D. none of these.

Answer: A

Solution:

Solution:
$2Cr^{3+} + 7H_2O \rightarrow Cr_2O_7^{2-} + 14H^+$
O.S. of Cr changes from +3 to +6 by loss of electrons. At anode oxidation takes place.

Question161

When the sample of copper with zinc impurity is to be purified by electrolysis, the appropriate electrodes are Cathode Anode [2002]

Options:

A. Cathode - pure zinc
Anode - pure copper

B. Cathode - impure sample
Anode - pure copper

C. Cathode - impure zinc
Anode - impure sample

D. Cathode - pure copper
Anode - impure sample.

Answer: D

Solution:

Solution:
Pure metal always deposits at cathode.

Question162

EMF of a cell in terms of reduction potential of its left and right electrodes is
[2002]

Options:

A. $E = E_{left} - E_{right}$

B. $E = E_{left} + E_{right}$

C. $E = E_{right} - E_{left}$

D. $E = -(E_{righ} + E_{lef})$.

Answer: C

Solution:

Solution:

E_{cell} = Reduction potential of cathode (right)
− reduction potential of anode (left)
$= E_{right} - E_{left}$

Question163

What will be the emf for the given cell

$Pt \mid H_2(P_1) \mid H^+(aq) \mid\mid H_2(P_2) \mid Pt$ [2002]

Options:

A. $\frac{RT}{F}\log_e \frac{P_1}{P_2}$

B. $\frac{RT}{2F}\log_e \frac{P_2}{P_1}$

C. $\frac{RT}{F}\log_e \frac{P_2}{P_1}$

D. none of these.

Answer: C

Solution:

Oxidation half call:-
$H_2(g) \longrightarrow 2H^+(1M) + 2e^-$
P_1 Reduction half cell
$2H^+(1M) + 2e^- \longrightarrow H_2(g)$
P_2 The net cell reaction
$H_2(g) \longrightarrow H_2(g)$
P_t
$E_{cell}^\circ = 0.00V ; n = 2$
$\therefore E_{cell} = E_{cell}^\circ - \frac{RT}{nF}\log_e K$
$= 0 - \frac{RT}{nF}\log_e \frac{P_2}{P_1}$
or $E_{cell} = \frac{RT}{2F}\log_e \frac{P_1}{P_2}$
