ELECTROCHEMISTRY

1. How many kJ of energy is spent when a current of 4 amp passes for 200 second under a potential of 115 V?

   (A) 52 kJ (B) 72 kJ (C) 82 kJ (D) 92 kJ

Answer: [D]

$\Delta G = -nFE = -QE = -4 \times 200 \times 115 \times 10^{-3} kJ; \Delta G = -92 kJ$

The energy spent by the source = 92 kJ

2. The minimum equivalent conductance in fused state is shown by -

   (A) $MgCl_2$ (B) $BeCl_2$ (C) $CrCl_2$ (D) $SrCl_2$

Answer: [B]

$BeCl_2$ has higher covalent character hence it ionizes in least extent in the fused state.

3. In which of the following electrochemical cell overall cell reaction is :

$HgO(s) + H_2(g) \rightarrow Hg(\ell) + H_2O(\ell)$

(A) $Pt|H_2(g)|H^+(aq)||HgO(s)Hg(\ell)|Pt$

(B) $Pt|H_2(g)|NaOH(aq)|HgO(s)|Hg(\ell)|Pt$

(C) $Pt|H_2(g)|H^+||NaOH(aq)|HgO(s)Hg(\ell)|Pt$

(D) $Pt|H_2(g)|H^+|Hg|O(s)Hg(\ell)|Pt$

Answer: [B]

$Pt|H_2(g)|NaOH(aq)|HgO(s)|Hg(\ell)|Pt$

LHE reaction: $H_2(g) + 2OH^- \rightarrow 2H_2O(\ell) + 2e$

RHE reaction: $HgO(s) + H_2O(\ell) + 2e \rightarrow Hg(\ell) + 2OH^-$

Net cell reaction: $HgO(s) + H_2(g) \rightarrow Hg(\ell) + H_2O$

4. What would be the product of electrolysis if molten ICl_3 is electrolysed?

(A) I_2 is produced at cathode and Cl_2 is produced at anode

(B) Cl_2 is produced at cathode and I_2 is produced at anode

(C) Both I_2 and Cl_2 are liberated at both electrodes

(D) ICl_2 is produced at cathode and ICl_4 is produced at anode

Answer: [C]

In the molten state ICl_3 ionises as follows

$$2ICl_3 \rightleftharpoons ICl_2^+ + ICl_4^-$$

Hence both I_2 and Cl_2 are produced at both the electrodes.

5. The equilibrium constant (K_{eq}) of weak acid HA in the aqueous solution of concentration c (M) is given by

(A) $K_{eq} = \frac{c\pi_c^2}{\pi_\infty - \pi_c}$ (B) $K_{eq} = \frac{c\pi_c^2}{\pi_\infty(\pi_\infty - \pi_c)}$

(C) $K_{eq} = \frac{c\pi_c^2}{\pi_\infty + \pi_c}$ (D) $K_{eq} = \frac{c\pi_c^2}{\pi_\infty(\pi_\infty + \pi_c)}$

Answer: [B]

$$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$$

$c(1-\alpha)$ $\quad c\alpha \quad c\alpha$

$$K_{eq} = \frac{c\alpha^2}{(1-\alpha)}$$

If $\alpha < 1$, then $\alpha = \frac{\pi_c}{\pi_\infty}$

$$\therefore\ K_{eq} = \frac{c\pi^2 c}{\pi(\infty\pi\infty - \pi c)}$$

6. At 298 K , the conductivity of a saturated solution of AgCl in water is $2.6\times10^{-6}\,ohm^{-1}cm^{-1}$ (Given $\lambda_m^\infty(Ag^+) = 63\,ohm^{-1}cm^2 mol^{-1}$ & $\lambda_m^\infty(Cl^-) = 67\,ohm^{-1}cm^2 mol^{-1}$)

Therefore solubility product of AgCl is

(A) 2×10^{-5} (B) 4×10^{-10} (C) 4×10^{-16} (D) 2×10^{-8}

Answer: [B]

$$s = \frac{10^3 \times K}{\lambda_m^\infty(AgCl)} = \frac{10^3 \times 2.6\times10^{-6}}{(63+67)} = 2\times10^{-5}\,(M)$$

$$K_{sp} = (2\times10^{-5})^2 = 4\times10^{-10}.$$

7. Faraday's law of electrolysis are related to the

(A) atomic number of the cation

(B) atomic number of anion

(C) equivalent mass of the electrolyte

(D) speed of the cation

Answer: [C]

Faraday's law of electrolysis are related to the equivalent mass of the electrolyte.

8. It is impossible to measure the actual voltage of any half-cell by itself because

(A) Both half-cell reactions take place simultaneously

(B) Of resistance of the wire

(C) A reaction does not take place on its own

(D) None of the above

Answer: [C]

It is impossible to measure the actual voltage of any half cell by itself because A reaction does not take place on it's own. Thus, there must be a SOURCE and SINK (somewhere for the electrons to go) of electrons.

9. The reaction occurring at anode when the electrolysis of an aqueous solution containing Na_2SO_4 and $CuSO_4$ is done using Pt electrode is

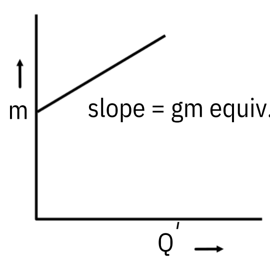
(A) $Cu \rightarrow Cu^{2+}+2e$ (B) $2SO_4^{2-}+2H_2O \rightarrow 2H_2SO_4+O_2+$

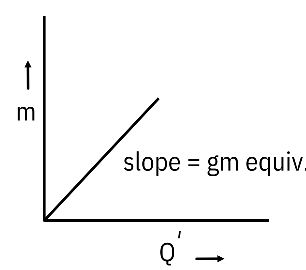
(C) $2H_2O \rightarrow O_2+4H^{+}+4e$ (D) $2Cl^{-} \rightarrow Cl_2+2e$

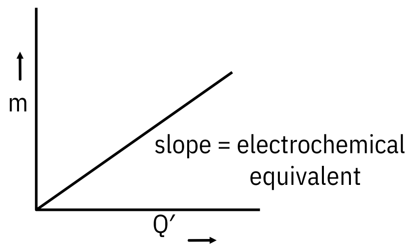
Answer: [C]

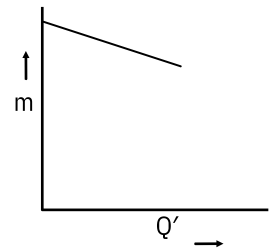
Anodic reaction is $2H_2O \rightarrow O_2+4H^{+}+4e$

10. During electrolysis, the gm of product produced (m) and when plotted against the electric charge Q′ in faradays which of the following plot is correct?

(A) 

(B) 

(C) 

(D) 

Answer: [B]

$m = \frac{EQ}{F}$; $\frac{Q}{F}$ represent the electric charge Q’

$\therefore$ $m = EQ'$

11. Which of the following is not correct for dry cell?
    a) Zinc is used as anode
    b) Manganese is reduced from + 4 to + 3 state
    c) It is a primary cell
    d) NH_3 gas is liberated out

Answer: [d]

Explanation:

Ammonia produced due to reaction

$MnO_2 + NH_4^+ + 1e^- \rightarrow MnO(OH) + NH_3$

Combines with Zn^{+2} ions to form $[Zn(NH_3)_4]^{2+}$

12. Two half cells have reduction potentials −0.76 V and −0.13 V respectively. A galvanic cell is made from these two half cells. Which of the following statements is correct?
    (A) Electrode of half-cell potential −0.76 V serve as cathode
    (B) Electrode of half-cell potential −0.76 V serve as anode

(C) Electrode of half-cell potential −0.13 V serve as anode

(D) Electrode of half-cell potential −0.76 V serve as positive electrode and −0.13 V as negative electrode

Answer: [B]

The electrode with more negative reduction potential constitute the anode.

13. Given that $E^o_{Fe^{3+}/Fe^{2+}} = -0.36V$ and $E^o_{Fe^{2}/Fe} = -0.439V$ respectively.

Therefore value of $E^o_{Fe^{3+}, Fe^{2+}/Pt}$ is:

(A) (−0.036 −0.439) V

(B) (−0.036 + 0.439) V

(C) [3(−0.036) + 2(−0.439)]V

(D) [3(−0.036)−2 (−0.439)] V

Answer: [D]

$Fe^{3+} + 3e^- \rightarrow Fe$	$\Delta G^0_1 = -3F(-0.36)V$
$Fe^{2+} + 2e^- \rightarrow Fe$	$\Delta G^0_2 = -2F(-0.439)V$
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	$\Delta G^0_{net} = \Delta G^0_1 - \Delta G^0_2$

or $\Delta G^0_{net} = -F[3(-0.36-2(-0.439)]V$

or $E^o = [3(-0.36)-2(-0.439)]V$

14. Beryllium is placed above magnesium in the II group. Beryllium dust, therefore, when added to $MgCl_2$ solution will –

(A) Have no effect

(B) Precipitate Mg metal

(C) Precipitate MgO

(D) Lead to dissolution of Be metal

Answer: [A]

No doubt Be is above Mg in periodic table but it is below Mg in electrochemical series.

15. 3 faraday of electricity are passed through molten Al2O3, aqueous solution of CuSO4and molten NaCl taken in three different electrolytic cells. The amount of Al, Cu and Na deposited at the cathodes will be in the ratio of

(A) 1 mole : 2 mole : 3 mole

(B) 3 mole : 2 mole : 1 mole

(C) 1 mole : 1.5 mole : 3 mole

(D) 1.5 mole : 2 mole : 3 mole

Answer: [C]

Equivalent of Al = Equivalent of Cu = Equivalent of Na

or $\frac{1}{3}$moleAl = $\frac{1}{2}$moleCu = 1 mole Na

or 2 : 3 : 6 or 1 : 1.5 : 3 mole ratio

16. Which statement is not correct for Kohlrausch law?

(A) The law is valid at infinite dilution

(B) The law is valid for weak electrolytes only

(C) The basis of law is independent migration of ions

(D) None of the above

Answer: [B]

The law is valid for weak and strong electrolytes both.

17. In the equation $\Lambda_m = \Lambda^\circ_m - A c^{1/2}$, the constant 'A' depends on

a) Nature of solvent
b) Temperature
c) Type of electrolyte
d) All of theses

Answer: [d]

Explanation: Constant 'A' depends on charges on cation and anion produced on the dissociation of electrolyte in the solution.

18. The standard reduction potentials at 298 K for the following half reactions are given against each

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ -0.762

$Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$ -0.740

$2H^+(aq) + 2e^- \rightarrow H_2(g)$ 0.00

$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$ 0.770

(a) Zn(s) (b) Cr(s) (c) $H_2(g)$ (d) $Fe^{2+}(aq)$

Answer: [a]

$Zn^{2+}(aq)$ is having minimum reduction potential. So oxidation potential of Zn(s) will be maximum and it will be strongest reducing agent.

19. The amount of electrical energy produced during the passage of 2 ampere current for 150 s under a potential of 125 volt is

a) 9.375 kJ
b) 37.5 kJ
c) 75 kJ
d) 3.75 kJ

Answer: [b]

Explanation:

Electrical energy = Charge × potential

Charge "Q" = current (i) × time (t)

Hence, Electrical energy = i × t × V = 2 ×150 × 125 = 37.5 kJ

20. If $E^0_{Fe^{2+}/Fe}$ is x_1, $E^0_{Fe^{3+}/Fe}$ is x_2 then $E^0_{Fe^{3+}/Fe^{2+}}$ will be

(a) $3x_2 - 2x_1$ (b) $x_2 - x_1$

(c) $x_2 + x_1$ (d) $2x_1 + 3x_2$

Answer: [a]

$$\Delta G_3 = \Delta G_2 - \Delta G_1$$

$$-n_3FE^0_3 = -n_2FE^0_2 + n_1FE^0_1$$

$$E^0_3 = \frac{n_2E^0_2 - n_1E^0_1}{n_3} = 3x_2 - 2x_1$$

21. $Zn|Zn^{2+}(c_1)||Zn^{2+}(c_2)|Zn$. For this cell ΔG is negative if

(a) $c_1 = c_2$ (b) $c_1 > c_2$ (c) $c_2 > c_1$ (d) none

Answer: [c]

$$E = E0 - \frac{0.0591}{2}\log\frac{c1}{c2} = \frac{0.0591}{2}\log\frac{c2}{c1}$$

to make $\Delta G = -ve$

$\Delta E = +ve$

hence c2>c1

22. Chromium plating of an article was carried out by treating it as cathode in an electrolysis. If during electrolysis the weight of the article increased by 2.6 g and 0.56 dm3 of oxgen (corrected up to STP) evolves at inert anode, the oxidation state of the chromium ions being discharged must be

(a) +1 (b) +2 (c) +3 (d) +6

Answer: [b]

Equivalent of chromium deposited = Equivalent of O2 evolved

$$\frac{2.6}{52} \times n = \frac{0.56}{22.4} \times 4$$

n = 2 i.e. $Cr^{2+} \rightarrow Cr$

23. Three faradays of charge is passed through an aqueous solution of $NaClO_3$ when $NaClO_4$ is formed according to the reaction

$$NaClO_3 + H_2O \rightarrow NaClO_4 + H_2$$

How many moles of $NaClO_4$ will be formed during the electrolysis?

(a) 0.75 (b) 1.0 (c) 1.5 (d) 3.0

Answer: [b]

Charge = Current(amp) ×Time(sec) = 1 ×193 = 193 C

No. of faradays $= \frac{193}{96500} = \frac{1}{500}$

1 F charge = 6.023×10^{23} electrons

$\frac{1}{500}$ F charge $= 6.023\times10^{23}\times\frac{1}{500} = 1.2\times10^{21}$

24. E°for the cell $Zn|Zn^{2+}(aq)||Cu^{2+}(aq)|Cu$ is 1.10 V at 25°C.The equilibrium constant for the reaction

$Zn + Cu^{2+}(aq) \rightarrow Cu + Zn^{2+}(aq)$ is of the order of

(a) 10^{-28} (b) 10^{37} (c) 10^{18} (d) 10^{17}

Answer: [b]

The nernst expression for the given cell is

$$E_{cell} = E^0_{cell} - \frac{0.059}{n}\log\frac{[Zn^{2+}]}{[Cu^{2+}]}$$

$$0 = 1.1 - \frac{0.059}{2}\log K_{eq}$$

$\therefore$

$$K_{eq} = 1.94\times10^{37}$$

25. In the electrolysis of $CuCl_2$ solution using Cu electrodes, the weight of Cu anode increased by 2 gram at cathode. In the anode.

(A) 0.2 mole of Cu^{2+} will go into solution.

(B) 560 ml O_2 liberate

(C) No loss in weight

(D) 2 gram of copper goes into solution as Cu^{2+}

Answer: [d]

During electrolysis of $CuSO_4$solution using Cu electrodes, the cell reaction is

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

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

The loss in weight of anode is equal to gain in weight of cathode.

26. Electrolysis of dil H_2SO_4liberates gases at anode and cathode

(A) O_2& SO_2respectively (B) SO_2& O_2respectively

(C) O_2& H_2respectively (D) H_2& O_2respectively

Answer: [c]

The reactions are $2H^+ + 2e^- \rightarrow H_2$ and $2H_2O \rightarrow 4H^+ + O_2 + 4e^-$

The volume of H_2liberated will be twice that of O_2.

27. A 500 ml of 0.2 M Cd^{+2} electrolysed with a current of 96.5

A. If the remaining concentration of Cd^{+2} ions is 0.1 M, then the duration of electrolysis is

a) 50 s
b) 75 s
c) 125 s
d) 100 s

Answer: [d]

Explanation:

Number of moles of Cd^{+2} before electrolysis

$= \frac{M \times V}{1000} = \frac{500 \times 0.2}{1000}$

Number of moles of Cd^{+2} after electrolysis

$= \frac{M \times V}{1000} = \frac{500 \times 0.1}{1000} = 0.05$

Moles of Cd^{+2} electrolysed = 0.1 – 0.05 = 0.05

Moles = $\frac{i \times t}{n \times 96500}$

$0.05 = \frac{96.5 \times t}{2 \times 96500}$ ⇒ t = 100 s

28. The specific conductance has the unit :

(A) $ohm^{-1}\ cm^{-1}$ (B) ohm.cm (C) $ohm\ cm^{-1}$ (D) $ohm^{-1}.cm$

Answer: [A]

Specific conductance = Observed conductance $\times \frac{l}{a}$

$= ohm^{-1} \times \frac{cm}{cm^2} = ohm^{-1}.cm^{-1}$

29. Electrolysis of a solution $MnSO_4$ in aqueous sulphuric acid is a method for the preparation of MnO_2. Passing a current of 27 A for 24 hours gives 1 kg of MnO_2. The current efficiency is -

(A) 100% (B) 95.185% (C) 80% (D) 82.951%

Answer: [B]

The electrode reactions are

Anode $Mn^{2+}(aq) + 2H_2O(l) \longrightarrow MnO_2(s) + 4H^+(aq) + 2e^-$

Cathode $2H^+(aq) + 2e^- \longrightarrow H_2(g)$

Overall reaction $Mn^{2+}(aq) + 2H_2O(l) \longrightarrow MnO_2(s) + 2H^+(aq) + H_2(g)$

The actual needed to deposit 1 kg MnO_2 is given by the expression

$$I = \left(\frac{m}{t}\right)\left(\frac{z}{M}\right) = \frac{(1000\,g)}{24\times60\times60}\left(96500\,C\,mol^{-1}\right)\left(\frac{2}{(54.9+2\times16)\,g\,mol^{-1}}\right) \quad 57$$

The current efficiency = $\left(\frac{25.70\,A}{27\,A}\right)(100) = 95.185\%$

30. At 25ºC,the standard emf of cell having reactions involving two electron change is found to be 0.295 V. The equilibrium constant of the reaction is-

(A) 29.5×10^{-2} (B) 10 (C) 10^{10} (D) 29.5×10^{10}

Answer: [C]

$$E^\circ = \frac{0.059}{n}\log K_c$$

$$\therefore \quad 0.295 = \frac{0.059}{n}\log K$$

$$\therefore \quad K = 10^{10}$$

31. The number of faradays required to produce one mole of water from a hydrogen - oxygen fuel cell containing aqueous alkali as electrolyte is-

(A) 1 (B) 3 (C) 2 (D) 4

Answer: [C]

This is based on the application of Faraday's laws of electrolysis. One mole of water can be obtained by the combination of a mole of H_2 and $\frac{1}{2}$ mole of oxygen which may be produced by the passage of 2 Faradays.

32. How much current is necessary to produce H_2 gas at the rate of 1 cm^3 per second under STP.

(A) 4.305 amp (B) 17.22 amp (C) 8.61 amp (D) 2.1525 amp

Answer: [c]

1 mol H_2 gas at STP = 2F = 2×96500 C

22400 cm^3 of H_2 gas at STP = 193×10^3 C

1 cm^3 of H_2 gas at STP per second = $\frac{193 \times 10^3}{22400}$ =8.61amp.

33. During discharge of a lead storage cell the density of sulphuric acid in the cell-

(A) Increases

(B) Decreases

(C) Remains unchanged

(D) Initially increases but decreases subsequently

Answer: [B]

During the discharge of lead storage cell, sulphuric acid is consumed. Its concentration decreases and therefore, density decreases.

34. F_2 gas can't be obtained by the electrolysis of any F^- salt because-

(A) Fluorine is the strongest reducing agent

(B) Fluorine is the strongest oxidising agent.

(C) Fluorine easily combine with atmospheric O_2

(D) All

Answer: [B]

Since Fluorine is the strongest oxidising agent so it will destroy the electrode employed.

35. A galvanic cell is composed of two hydrogen electrodes, one of which is a standard one. In which of the following solutions should the other electrode be immersed to get maximum emf.

(A) 0.1 M HCl
(B) 0.1 MCH_3COOH
(C) 0.1 MH_3PO_4
(D) 0.1 M H_2SO_4

Answer: [D]

H_2SO_4will furnish maximum H^+.

36. The solubility product of silver iodide is $8.3\times10^{-17}$and the standard potential (reduction) of Ag, Ag^+electrode is + 0.800 volts at 25ºC.The standard potential of Ag, AgI/I^- electrode (reduction) from these data is-

(A) −0.30 V (B) + 0.15 V (C) + 0.10 V (D) −0.15 V

Answer: [D]

Solubility product reaction is

$AgI \rightleftharpoons Ag^+ + I^-$

By calculating the EMF of this cell reaction from the given data and relating to Kspvia the ΔGºof the reaction, we can obtain Ksp.

37. A certain current liberates 0.504 g of hydrogen in 2 hours. How many gram of copper can be liberated by the same current flowing for the same time in aqueous $CuSO_4$ solution:

(A) 12.7 (B) 16 (C) 31.8 (D) 63.5

Answer: [B]

Eq. of H_2 = Eq. of Cu

$\therefore \frac{0.154}{1} = \frac{w}{63.5/2}$

$\therefore w_{Cu} = 16$ g

38. The calomel electrode is reversible with respect to-

(A) Mercury (B) H^+ (C) Hg^{2+} (D) Cl^-

Answer: [D]

Since half cell reaction; $2e + Hg_2Cl_2(aq.) \rightarrow 2Hg(l) + 2Cl^-$

39. Zn amalgam is prepared by electrolysis of aqueous $ZnCl_2$ using Hg cathode (9gm). How much current is to be passed through $ZnCl_2$ solution for 1000 seconds to prepare a Zn Amalgam with 25% Zn by wt. (Zn = 65.4)

(A) 5.6 amp (B) 7.2 amp (C) 8.85 amp (D) 11.2 amp

Answer: [C]

Let x gm of Zn deposit on 9 gm of Hg

% of Zn in Amalgam = $\frac{x \times 100}{9 + x} = 25 \therefore x = 3$gm

Eq. of Zn = $\frac{3 \times 2}{65.4}$; Current = $\frac{6}{65.4} \times \frac{96500}{1000} = 8.85$amp.

40. E^o for the reaction $Fe + Zn^{2+} = Zn + Fe^{2+}$ is -0.35 V. The given cell reaction is-

(A) Feasible (B) Not feasible

(C) In equilibrium (D) None

Answer: [B]

Since EMF of the cell is negative i.e. Free energy change will be positive so cell reaction will not be feasible.

41. A certain metal salt solutions is electrolysed in series with a silver coulometer. The weights of silver and the metal deposited are 0.5094 g and 0.2653 g. Calculate the valency of the metal if its atomic weight is nearly that of silver.

(A) 1 (B) 2 (C) 3 (D) 4

Answer: [B]

Evidently $\frac{0.5094}{0.2653} = \frac{\text{Equivalent weight of Ag}}{\text{Equivalent weight of metal}} \approx 0.2$

$\therefore$ valency ratio $= \frac{\text{Valency of metal}}{\text{Valency of Ag}} = 2$

42. In the electrochemical cell,

$Cd + Sn^{+4} \rightleftharpoons Cd^{+2} + Sn^{+2}$

Increase in the concentration of Sn^{+2}

a) Increasing the emf of cell

b) Decreasing the emf of cell

c) No change in emf of cell

d) Unpredictable

Answer: [b]

Explanation: Increase in concentration of cathodic product emf of cell decreases.

43. The reduction potential of hydrogen half-cell will be negative if

(A) $p(H_2)=1$ atm and $[H^+]=1$ M (B) $p(H_2)=2$ atm and $[H^+]=2$ M

(C) $p(H_2)=2$ atm and $[H^+]=1$ M (D) $p(H_2)=1$ atm and $[H^+]=2$ M

Answer: [C]

$$H^+ + e^- \rightarrow \frac{1}{2}H_2$$

Apply Nernst equation

$$E = 0 - \frac{0.059}{1}\log\left(\frac{P_{H_2}^{\frac{1}{2}}}{[H^+]}\right)$$

$$E = -\frac{0.059}{1}\left(\frac{2^{\frac{1}{2}}}{1}\right)$$

Therefore E is negative.

44. Na-amalgam is prepared by electrolysis of NaCl solution using liquid Hg as cathode. How long should the current of 10 amp. is passed to produce 10% Na –Hg on a cathode of 10 gm Hg. (atomic mass of Na = 23).

(A) 7.77 min. (B) 9.44 min. (C) 5.24 min. (D) 11.39 min.

Answer: [A]

90 gm Hg has 10 gm Na

$\therefore$ 10gm Hg= $\frac{10}{90}\times 10 = \frac{10}{9}$ gmNa

$\therefore$ weight of Na $= \frac{M}{n} \times \frac{i \times t}{96500}$

$\frac{10}{9} = \frac{23}{1} \times \frac{10 \times t}{96500}$ $[\therefore Na^{+} + e \rightarrow Na]$

$= \frac{10 \times 96500}{9 \times 10 \times 23} = 7.77 \text{min}$

45. Calculate the quantity of electricity that would be required to reduce 12.3g of nitrobenzene to aniline. If current efficient is 50%. If the potential drops across the cell is 3.0 volts

(A) 369000 coulomb
(B) 115800 coulomb
(C) 32100 coulomb
(D) 521900 coulomb

Answer: [B]

The reduction reaction is

$C_6H_5NO_2 + 3H_2 C_6H_5NH_2 + 2H_2O$

Hydrogen required for reduction of 12.3/123 or 0.1 mole of nitrobenzene = 0.1×3=0.3mole

Amount of charge required for liberation of 0.3 mole of hydrogen = 2×96500×0.3=57900 coulomb

Actual amount of charge required as efficiency is 50%

= 2 ×57900 = 115800 coulomb

Energy consumed = 115800 ×3.0 = 347400 J = 347.4 kJ