

Electro Chemistry Part-II

- Faraday's laws of electrolysis are related to the**
 - Molar mass of the electrolyte
 - Equivalent weight of the cation or anion
 - Molecular mass of the electrolyte
 - Atomic mass of the cation or anion
- The unit of electrochemical equivalent is**
 - Gram. Coulomb
 - Gram. Ampere
 - Gram / Coulomb
 - Coulomb / Gram
- The mass deposited by the passage of 1 amp current for one second is called as**
 - A mole
 - Gram equivalent weight
 - Electro-chemical equivalent
 - Gram atomic weight.
- A copper voltameter, a silver voltameter and a water voltameter are connected in series and current is passed for some time. The ratio of the number of moles of copper, silver and hydrogen formed at the cathode is**
 - 2:1:1
 - 1:1:1
 - 1:2:1
 - 1:2:2

Hint: when same quantity of electricity is passed through different electrolyte solutions then

no. of moles of elements deposited are in the ratio $1/Z_1:1/Z_2:1/Z_3:\dots$ where Z_1, Z_2, Z_3, \dots are valency of elements.

$$\therefore n_{\text{Cu}}: n_{\text{Ag}}: n_{\text{H}_2} = 1/2: 1/1: 1/2 = 1:2:1$$

- When one faraday of current is passed, which of the following would deposit one gram atomic weight of the metal?**
 - BaCl₂
 - NaCl
 - AlCl₃
 - CuCl₂

Hint: For Na⁺, GEW=GAW

Hint: EMF of cell = $E_{Ag^+/Ag}^0 - E_{Cu^{2+}/Cu}^0$

18. The significance of using saturated solution of KNO₃ as electrolyte in the Salt Bridge is

- 1) Velocity of K⁺ is greater than that of NO₃⁻.
- 2) Velocity of NO₃⁻ is greater than that of K⁺.
- 3) Velocities of both K⁺ and NO₃⁻ are nearly the same.
- 4) KNO₃ is highly soluble in water.

19' The standard electrode potential of the two half cells are given below



The voltage of cell formed by combining the two half cells would be

- 1) -1.02 2) +0.52 volt 3) +1.02 volt 4) -0.52 volt

Hint: in this Zn acts as anode (lower SRP) and Ni acts as cathode (higher SRP).

$$Emf = E^0_{\text{cathode}} - E^0_{\text{anode}}$$

20) The hydrogen electrode is dipped in a solution of pH = 3 at 25⁰ C. The potential of the cell would be (the value of 2.303 RT/F is 0.059 V)

- 1) 0.177 V 2) 0.087 V 3) -0.177 V 4) 0.059 V

Hint: R.P of Hydrogen electrode = $-0.059 \times \text{pH}$

21. When 965 amp current is passed through aqueous solution of salt X using platinum electrodes for 10 sec, the volume of gasses liberated at the respective electrodes is in 1:1 ratio. Then X is

- 1) MgSO₄ 2) AgCl 3) MgCl₂ 4) KNO₃

Hint: aqueous MgCl₂ contains Mg⁺², H⁺, Cl⁻ and OH⁻ ions. On electrolysis it gives H₂ at cathode and Cl₂ at anode. As same quantity of electricity is passed volumes of both are in the ratio 1:1

22. Three faradays of electricity are passed through molten Al_2O_3 , aqueous solution of CuSO_4 and 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

- 1) 1 mole :2 mole :3 mole 2) 1 mole:1.5 mole :3 mole
3) 3 mole :2 mole :1 mole 4) 1 mole :1.5 mole :2 mole

Hint; Three faradays will deposit three equivalents of each electrolyte.

23. Number of coulombs of current required to convert completely 1 mole of MnO_4^- ions in acid medium in to Mn^{2+} ions is

- 1) 96500 2) 5×96500
3) 96500×2 4) 96500×6

Hint: in acid medium $\text{MnO}_4^- + 5e \rightarrow \text{Mn}^{2+}$

No. of electrons gained per ion = 5, \therefore Current required = 5F

24. How long will a current of 2 ampere take for complete deposition of copper from 0.5 litre of 1N CuSO_4 . Solution ?

- 1) 96500 sec 2) 2×96500 sec 3) 48250 sec 4) 24125 sec

Hint: no. of equivalents of $\text{Cu}^{+2} = N \times V$ in lit = $1 \times 0.5 = 0.5$.

1 Equivalent requires 1 faraday i.e. 96500 coulombs. Thus 0.5 equivalents need 48250 coulombs.

i.e. $Q = c t$ or $t = Q/c = 48250/2 = 24125$ sec

25. The number of electrons required for deposition of one mole of copper on the cathode during the electrolysis of CuSO_4 solution is

- 1) 6.0×10^{23} 2) 1.2×10^{24} 3) 4.8×10^{24} 4) $3 \times 10^{23} E$

Hint: 1 mole of Cu^{2+} requires 2F (or) 2 'N' e⁻s i.e. $2 \times 6 \times 10^{23} = 12 \times 10^{23} = 1.2 \times 10^{24}$

26. 1.2 gm of a metal is deposited in 965 seconds by passing a current of 10 amperes through its respective electrolytic solution. The equivalent mass of metal is

- 1) 12 2) 24 3) 60 4) 6

Hint; $m = Ect/F$ or $E = mF/ct = 1.2 \times 96500 / 965 \times 10 = 12$

- 27. On passing a current through a molten aluminium chloride for some time, produced 11.2lit of Cl_2 at NTP at anode, the quantity of aluminium deposited at cathode is**

- 1) 27grams 2) 18gram 3) 9gram 4) 36 gram

Hint: $W_{Al} / W_{Cl_2} = E_{Al} / E_{Cl_2}$, 11.2 lit Cl_2 at STP = 35.5 gm = 1 GEW

1 GEW of 'Al' will be deposited which is '9' gm

- 28. In the electrolysis of acidulated H_2O , 11.2 litres of H_2 is liberated at the cathode. The volume of O_2 liberated at the anode is**

- 1) 11.2 lit 2) 5.6 lit 3) 2.8lit 4) 22.4 lt

Hint; 1 faraday gives 11.2 lit of H_2 or 5.6lit of O_2 at STP' i.e 11.2 lit of H_2 = 5.6lit of O_2

- 29. How much time is required for a current of 2 amp to decompose 18g of water**

- 1) 96500 hours 2) 9650 hours 3) 26.8 hours 4) 2.68 hours

Hint: for water, GEW=9gm. 9gm water decomposed by 96500c

\therefore 18gm water decomposed by $2 \times 96500c$,

$\therefore t = Q/c = 2 \times 96500 / 2 = 96500 \text{ sec} = 96500 / 3600 = 26.8 \text{ hours}$

- 30. When electric current is passed through molten NaCl for 1930 sec, 1120 ml of Cl_2 is liberated at anode at STP. The current passed in amp is**

- 1) 0.05 2) 0.5 3) 5 4) 50

Hint: i.e 96500 coulombs = 11200ml of Cl_2 . \therefore 1120ml of Cl_2 is give by 9650 coulombs.

$C = Q/t = 9650 / 1930 = 5$

31. One faraday of electricity is passed separately through one litre of one molar aqueous solution of I) AgNO_3 , ii) SnCl_4 and iii) CuSO_4 . The number of moles of Ag, Sn and Cu deposited at cathode are respectively

- 1) 1.0, 0.25, 0.5 2) 1.0, 0.5, 0.25 3) 0.5, 1.0, 0.25 4) 0.25, 0.5, 1.0

Hint: moles of $\text{Ag}^+ : \text{Sn}^{+4} : \text{Cu}^{+2} = \frac{1}{1} : \frac{1}{4} : \frac{1}{2} = 1.0, 0.25, 0.5$

32. A current of 2amps passing for 5 hours deposits 22.2g of tin (at.wt. = 119), the oxidation state of tin is

- 1) Zero 2) Three 3) Two 4) Four

Hint; $m = Mct/ZF \therefore Z = Mct/mF = 119 \times 2 \times 5 \times 3600 / 22, 2 \times 96500 = 1.999$ ie; 2

33. 1 ampere current is passed for 60 seconds into an electrolytic cell. Number of electrons that pass through the solution is.

- 1) 6.0×10^{23} 2) 1.2×10^{24} 3) 3.75×10^{20} 4) 7.48×10^{21}

Hint: $96,500\text{C} = 6.023 \times 10^{23} \text{ e}^- \text{s}$

$60\text{C} = 60 \times 6.023 \times 10^{23} \text{ e}^- \text{s} / 96500 = 3.75 \times 10^{20}$

34. The STP volume of oxygen liberated by 2 ampere of current when passed through acidulated water for 3 minutes and 13 seconds, is

- 1) 120cc 2) 22.4cc 3) 11.2cc 4) 44.8 cc

Hint: $Q = ct = 2 \times 193 = 386 \text{ coulombs}$.

96500C liberates 5600cc of oxygen.

$\therefore 386\text{C}$ liberates $386 \times 5600 / 96500 = 22.4$ of oxygen

35. Total volume of gases evolved at STP when 36g of H_2O are completely electrolysed between platinum electrodes

- 1) 22.4lit 2) 44.8 lit 3) 33.6lit 4) 67.2 lit

Hint: $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$

i.e. 36gm of water gives 3 moles of gases.

\therefore Total volume of gases evolved at STP = $3 \times 22.4 \text{ lit} = 67.2 \text{ lit}$

36. For a cell the cell reaction is $\text{Mg(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Cu(s)} + \text{Mg}^{2+}(\text{aq})$.

If the S.R.P. values of Mg and Cu are -2.37v and $+0.34\text{v}$ respectively, the e.m.f. of the cell is

- 1) $+2.03\text{V}$ 2) -2.03V 3) $+2.71\text{V}$ 4) -2.71V

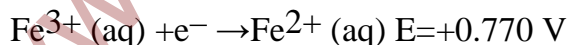
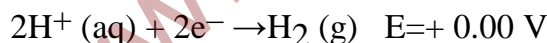
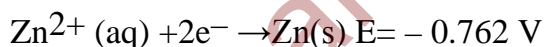
Hint; $E_{\text{Cell}}^0 = E_{\text{Sn}^{2+}/\text{Sn}}^0 - E_{\text{Zn}^{2+}/\text{Zn}}^0$

37. The standard reduction potentials of Ag, Cu, Co and Zn are $+0.799, +0.337, -0.277$ and -0.762V respectively. Which of the following cells will have maximum cell e.m.f.?

- 1) $\text{Zn} | \text{Zn}^{2+}(1\text{M}) || \text{Cu}^{2+}(1\text{M}) | \text{Cu}$
 2) $\text{Zn} | \text{Zn}^{2+}(1\text{M}) || \text{Ag}^+(1\text{M}) | \text{Ag}$
 3) $\text{Zn} | \text{Zn}^{2+}(1\text{M}) || \text{Co}^{2+}(1\text{M}) | \text{Co}$
 4) $\text{Cu} | \text{Cu}^{2+}(1\text{M}) || \text{Ag}^+(1\text{M}) | \text{Ag}$

Hint: emf of the cell is maximum when cathode has highest and anode has lowest SRP values.

38. The standard reduction potentials at 298K for the following half-cell reaction are given below:



Which one is the strongest reducing agent?

- 1) Zn(s) 2) Cr(s) 3) $\text{H}_2(\text{g})$ 4) $\text{Fe}^{+2}(\text{aq})$

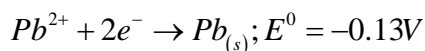
Hint: lower the SRP value higher is the reducing power

39. The standard reduction potential of three metals X, Y, Z are 0.52, -3.03 and -1.18V respectively. The order of reducing power of the corresponding metals is

- 1) $Y > Z > X$ 2) $X > Y > Z$ 3) $Z > Y > X$ 4) $Z > X > Y$

Hint: lower the SRP value higher is the reducing power

40. **Given:** $Fe^{+2} + 2e^{-} \rightarrow Fe; E^0 = -0.44V$



Which of the following metals will oxidise iron?

- a) Ag b) Cu c) Pb d) None of these

- 1) Only a 2) a,b 3) a,b,c 4) Only b

Hint: higher SRP metal oxidizes lower SRP metals. SRP values of pb, Ag, Cu are higher than Fe.

\therefore Pb, Ag, Cu oxidises the Fe

41. When electrons are used in the electrolysis of a metallic salt, 1.9 gm of the metal is deposited at the cathode. The atomic weight of that metal is 57. So oxidation state of the metal in the salt is

- 1) +2 2) +3 3) +1 4) +4

Hint; $Z = Mct/mF$

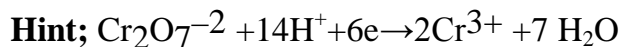
42. The electrochemical equivalent of an element is 0.001118 gm/coulomb. Its equivalent weight is

- 1) 10.7 2) 53.5 3) 111.8 4) 107

Hint; $GEW = eX 96500$

43. The charge required to reduce 1mole $Cr_2O_7^{-2}$ to Cr^{+3} ions is

- 1) 3F 2) 2F 3) 6F 4) 12F



No. of electrons gained per ion = 6

\therefore Charge required = 6F

44. One ampere of current is passed for 9650 seconds through molten AlCl_3 . What is the weight in grams of Al deposited at cathode? (Atomic weight of Al=27).

1) 0.9 2) 2.7 3) 9 4) 27

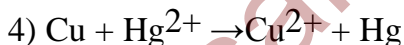
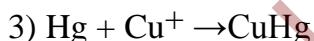
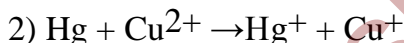
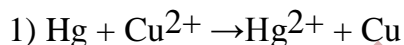
Hint; $m = \frac{Ect}{F}$, for Al $E = \frac{27}{3} = 9$, $m = \frac{9 \times 1 \times 9650}{96500} = 0.9\text{gm}$

45. The $E^0_{\text{M}^{3+}/\text{M}^{2+}}$ values for Cr, Mn, Fe and Co are -0.14, +1.57, +0.77 and +1.97 V respectively. For which one of these metals the change in oxidation state from +2 to +3 is easiest?

1) Cr 2) Co 3) Fe 4) Mn

Hint: lower SRP metal is strong reducing agent i.e. easily oxidized.

46. The cell reaction of the galvanic cell, $\text{Cu(s)} / \text{Cu}^{2+}(\text{aq}) // \text{Hg}^{2+}(\text{aq}) / \text{Hg(l)}$ is



47. (A): The Daniel cell becomes dead after some time.

(R): Oxidation potential of zinc anode increases and that of copper cathode decreases.

The correct answer is

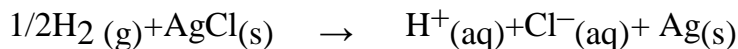
1) Both A and R are true, and R is the correct explanation of A.

2) Both A and R are true, and R is not the correct explanation of A.

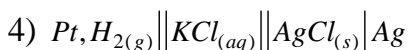
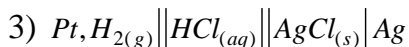
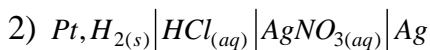
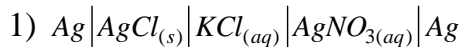
3) A is true, but R is false.

4) Both A and R are false.

48. The reaction,



occurs in the galvanic cell :



49. Which metal will dissolve if the cell works $Cu|Cu^{2+}||Ag^+|Ag$

- 1) Cu 2) Ag 3) Both (1) and (2) 4) None of these

Hint: at anode metal is dissolved.

50. Given standard electrode potentials



The standard electrode potential E^0 for $Fe^{3+} + e^- \rightarrow Fe^{2+}$ is

- 1) 0.476V 2) -0.404V 3) 0.40V 4) +0.772V

Hint: required equations obtained by subtracting eq 2 from eq 1. $\therefore E_3^0 = \frac{n_1 E_1^0 - n_2 E_2^0}{n_3}$

$$= 3[-0.036] - 2[-0.44]/1 = +0.772V$$

Key

1) 2 2) 3 3) 3 4) 3 5) 2 6) 2 7) 2 8) 3 9) 3 10) 2

11) 4 12) 3 13) 1 14) 4 15) 3 16) 2 17) 1 18) 3 19) 2 20) 3

21) 3 22) 2 23) 2 24) 4 25) 2 26) 1 27) 3 28) 2 29) 3 30) 3

31) 1 32) 3 33) 3 34) 2 35) 4 36) 3 37) 2 38) 1 39) 1 40) 3

41) 2 42) 4 43) 3 44) 1 45) 1 46) 4 47) 3 48) 3 49) 1 50) 4

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