CHAPTER 3

Electrochemistry

3.2 Galvanic Cells

1. The standard electrode potential ($E°$) values of Al^{3+}/Al, Ag^+/Ag, K^+/K and Cr^{3+}/Cr are –1.66 V, 0.80 V, –2.93 V and –0.74 V, respectively. The correct decreasing order of reducing power of the metal is

(a) Ag > Cr > Al > K (b) K > Al > Cr > Ag

(c) K > Al > Ag > Cr (d) Al > K > Ag > Cr

(Odisha NEET 2019)

2. A button cell used in watches function as following :

$Zn_{(s)} + Ag_2O_{(s)} + H_2O_{(l)} \square\ 2Ag_{(s)} + Zn^{2+}_{(aq)} + 2OH^-_{(aq)}$

If half cell potentials are

$Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn_{(s)}$; $E° = -0.76$ V

$Ag_2O_{(s)} + H_2O_{(l)} + 2e^- \rightarrow 2Ag_{(s)} + 2OH^-_{(aq)}$; $E° = 0.34$ V

The cell potential will be

(a) 0.84 V (b) 1.34 V

(c) 1.10 V (d) 0.42 V *(NEET 2013)*

3. Standard reduction potentials of the half reactions are given below :

$F_{2(g)} + 2e^- \rightarrow 2F^-_{(aq)}$; $E° = +2.85$ V

$Cl_{2(g)} + 2e^- \rightarrow 2Cl^-_{(aq)}$; $E° = +1.36$ V

$Br_{2(l)} + 2e^- \rightarrow 2Br^-_{(aq)}$; $E° = +1.06$ V

$I_{2(s)} + 2e^- \rightarrow 2I^-_{(aq)}$; $E° = +0.53$ V

The strongest oxidising and reducing agents respectively are

(a) F_2 and I^- (b) Br_2 and Cl^-

(c) Cl_2 and Br^- (d) Cl_2 and I_2

(Mains 2012)

4. Standard electrode potentials of three metals X, Y and Z are –1.2 V, + 0.5 V and – 3.0 V respectively. The reducing power of these metals will be

(a) $Y > Z > X$ (b) $Y > X > Z$

(c) $Z > X > Y$ (d) $X > Y > Z$ *(2011)*

5. Standard electrode potential for Sn^{4+}/Sn^{2+} couple is +0.15 V and that for the Cr^{3+}/Cr couple is –0.74 V. These two couples in their standard state are connected to make a cell. The cell potential will be

(a) + 1.19 V (b) + 0.89 V

(c) + 0.18 V (d) + 1.83 V *(2011)*

6. A solution contains Fe^{2+}, Fe^{3+} and I^- ions. This solution was treated with iodine at 35°C. $E°$ for Fe^{3+}/Fe^{2+} is + 0.77 V and $E°$ for $I_2/2I^- = 0.536$ V. The favourable redox reaction is

(a) I_2 will be reduced to I^-

(b) there will be no redox reaction

(c) I^- will be oxidised to I_2

(d) Fe^{2+} will be oxidised to Fe^{3+}. *(Mains 2011)*

7. Consider the following relations for emf of an electrochemical cell

(i) EMF of cell = (Oxidation potential of anode) – (Reduction potential of cathode)

(ii) EMF of cell = (Oxidation potential of anode) + (Reduction potential of cathode)

(iii) EMF of cell = (Reductional potential of anode) + (Reduction potential of cathode)

(iv) EMF of cell = (Oxidation potential of anode) – (Oxidation potential of cathode)

Which of the above relations are correct?

(a) (iii) and (i) (b) (i) and (ii)

(c) (iii) and (iv) (d) (ii) and (iv)

(Mains 2010)

8. On the basis of the following $E°$ values, the strongest oxidizing agent is

$[Fe(CN)_6]^{4-} \rightarrow [Fe(CN)_6]^{3-} + e^-$; $E° = -0.35$ V $Fe^{2+} \rightarrow Fe^{3+} + e^-$; $E° = -0.77$ V

(a) Fe^{3+} (b) $[Fe(CN)_6]^{3-}$

(c) $[Fe(CN)_6]^{4-}$ (d) Fe^{2+} *(2008)*

9. A hypothetical electrochemical cell is shown below :

$A\,/\,A^+\,(x\ M) \| B^+(y\ M) | B$

The emf measured is + 0.20 V. The cell reaction is

(a) $A + B^+ \rightarrow A^+ + B$

(c) $A^+ + e^- \rightarrow A$; $B^+ + e^- \rightarrow B$
(d) the cell reaction cannot be predicted. *(2006)*

10. $E^\circ_{Fe^{2+}/Fe} = -0.441$ V and $E^\circ_{Fe^{3+}/Fe^{2+}} = 0.771$ V, the standard EMF of the reaction $Fe + 2Fe^{3+} \rightarrow 3Fe^{2+}$ will be
(a) 0.111 V (b) 0.330 V
(c) 1.653 V (d) 1.212 V *(2006)*

11. Standard electrode potentials are Fe^{2+}/Fe; $E^\circ = -0.44$ and Fe^{3+}/Fe^{2+}; $E^\circ = 0.77$. Fe^{2+}, Fe^{3+} and Fe blocks are kept together, then
(a) Fe^{3+} increases (b) Fe^{3+} decreases
(c) Fe^{2+}/Fe^{3+} remains unchanged
(d) Fe^{2+} decreases. *(2001)*

12. Electrode potential for the following half-cell reactions are
$Zn \rightarrow Zn^{2+} + 2e^-$; $E^\circ = +0.76$ V;
$Fe \rightarrow Fe^{2+} + 2e^-$; $E^\circ = +0.44$ V.
The EMF for the cell reaction
$Fe^{2+} + Zn \rightarrow Zn^{2+} + Fe$ will be
(a) –0.32 V (b) + 1.20 V
(c) – 1.20 V (d) +0.32 V *(1996)*

13. An electrochemical cell is set up as : Pt; H_2 (1 atm)|HCl(0.1 M) || CH_3COOH (0.1 M) |H_2 (1 atm); Pt. The e.m.f. of this cell will not be zero, because
(a) acids used in two compartments are different
(b) e.m.f. depends on molarities of acids used
(c) the temperature is constant
(d) pH of 0.1 M HCl and 0.1 M CH_3COOH is not same. *(1995)*

14. Standard reduction potentials at 25°C of $Li^+|Li$, $Ba^{2+}|Ba$, $Na^+|Na$ and $Mg^{2+}|Mg$ are –3.05, –2.90, –2.71 and –2.37 volt respectively. Which one of the following is the strongest oxidising agent?
(a) Ba^{2+} (b) Mg^{2+}
(c) Na^+ (d) Li^+ *(1994)*

15. A solution of potassium bromide is treated with each of the following. Which one would liberate bromine?
(a) Hydrogen iodide (b) Sulphur dioxide
(c) Chlorine (d) Iodine *(1993)*

3.3 Nernst Equation

16. For the cell reaction :
$$2Fe^{3+}_{(aq)} + 2I^-_{(aq)} \longrightarrow 2Fe^{2+}_{(aq)} + I_{2(aq)}$$
$E^\circ_{cell} = 0.24$ V at 298 K. The standard Gibbs' energy ($\Delta_r G^\circ$) of the cell reaction is
[Given that Faraday constant, $F = 96500$ C mol^{-1}]
(a) 23.16 kJ mol^{-1} (b) –46.32 kJ mol^{-1}
(c) –23.16 kJ mol^{-1} (d) 46.32 kJ mol^{-1}
(NEET 2019)

17. For a cell involving one electron, $E^\circ_{cell} = 0.59$ V at 298 K, the equilibrium constant for the cell reaction is [Given that $\frac{2.303RT}{F} = 0.059$ V at $T = 298$ K]
(a) 1.0×10^{30} (b) 1.0×10^2
(c) 1.0×10^5 (d) 1.0×10^{10}
(NEET 2019)

18. In the electrochemical cell :
$Zn|ZnSO_4(0.01\ M)||CuSO_4(1.0\ M)|Cu$,
the emf of this Daniell cell is E_1. When the concentration of $ZnSO_4$ is changed to 1.0 M and that of $CuSO_4$ changed to 0.01 M, the emf changes to E_2. From the followings, which one is the relationship between E_1 and E_2? (Given, $RT/F = 0.059$)
(a) $E_1 < E_2$ (b) $E_1 > E_2$
(c) $E_2 = 0 \neq E_1$ (d) $E_1 = E_2$
(NEET 2017, 2003)

19. If the E°_{cell} for a given reaction has a negative value, which of the following gives the correct relationships for the values of ΔG° and K_{eq} ?
(a) $\Delta G^\circ > 0$; $K_{eq} < 1$ (b) $\Delta G^\circ > 0$; $K_{eq} > 1$
(c) $\Delta G^\circ < 0$; $K_{eq} > 1$ (d) $\Delta G^\circ < 0$; $K_{eq} < 1$
(NEET-II 2016, 2011)

20. The pressure of H_2 required to make the potential of H_2 electrode zero in pure water at 298 K is
(a) 10^{-10} atm (b) 10^{-4} atm
(c) 10^{-14} atm (d) 10^{-12} atm.
(NEET-I 2016)

21. A hydrogen gas electrode is made by dipping platinum wire in a solution of HCl of pH = 10 and by passing hydrogen gas around the platinum wire at one atm pressure. The oxidation potential of electrode would be
(a) 0.118 V (b) 1.18 V
(c) 0.059 V (d) 0.59 V *(NEET 2013)*

22. Consider the half-cell reduction reaction
$Mn^{2+} + 2e^- \rightarrow Mn$, $E^\circ = -1.18$ V
$Mn^{2+} \rightarrow Mn^{3+} + e^-$, $E^\circ = -1.51$ V
The E° for the reaction,
$3Mn^{2+} \rightarrow Mn^0 + 2Mn^{3+}$,
and possibility of the forward reaction are respectively
(a) – 4.18 V and yes (b) + 0.33 V and yes
(c) + 2.69 V and no (d) – 2.69 V and no.
(Karnataka NEET 2013)

23. The Gibbs' energy for the decomposition of Al_2O_3 at 500 °C is as follows

$\frac{2}{3}Al_2O_3 \rightarrow \frac{4}{3}Al + O_2,\ \Delta_rG = +960$ kJ mol^{-1}

The potential difference needed for the electrolytic reduction of aluminium oxide (Al_2O_3) at 500 °C is at least

(a) 4.5 V (b) 3.0 V

(c) 2.5 V (d) 5.0 V *(Mains 2012)*

24. The electrode potentials for, $Cu^{2+}_{(aq)} + e^- \rightarrow Cu^+_{(aq)}$ and $Cu^+_{(aq)} + e^- \rightarrow Cu_{(s)}$ are + 0.15 V and + 0.50 V respectively. The value of $E°_{Cu^{2+}/Cu}$ will be

(a) 0.500 V (b) 0.325 V

(c) 0.650 V (d) 0.150 V *(2011)*

25. For the reduction of silver ions with copper metal, the standard cell potential was found to be + 0.46 V at 25 °C. The value of standard Gibbs energy, $\Delta G°$ will be ($F = 96500$ C mol^{-1})

(a) –89.0 kJ (b) – 89.0 J

(c) –44.5 kJ (d) –98.0 kJ *(2010)*

26. Given :

(i) $Cu^{2+} + 2e^- \rightarrow Cu,\ E° = 0.337$ V

(ii) $Cu^{2+} + e^- \rightarrow Cu^+,\ E° = 0.153$ V

Electrode potential, $E°$ for the reaction,

$Cu^+ + e^- \rightarrow Cu$, will be

(a) 0.90 V (b) 0.30 V

(c) 0.38 V (d) 0.52 V *(2009)*

27. Standard free energies of formation (in kJ/mol) at 298 K are –237.2, –394.4 and –8.2 for $H_2O_{(l)}$, $CO_{2(g)}$ and pentane$_{(g)}$ respectively. The value of $E°_{cell}$ for the pentane-oxygen fuel cell is

(a) 1.0968 V (b) 0.0968 V

(c) 1.968 V (d) 2.0968 V *(2008)*

28. The equilibrium constant of the reaction :

$Cu_{(s)} + 2Ag^+_{(aq)} \rightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$; $E° = 0.46$ V at 298 K is

(a) 2.0×10^{10} (b) 4.0×10^{10}

(c) 4.0×10^{15} (d) 2.4×10^{10} *(2007)*

29. The standard e.m.f. of a galvanic cell involving cell reaction with $n = 2$ is found to be 0.295 V at 25°C. The equilibrium constant of the reaction would be

(a) 2.0×10^{11} (b) 4.0×10^{12}

(c) 1.0×10^2 (d) 1.0×10^{10}

(Given F = 96500 C mol^{-1}, $R = 8.314$ J K^{-1} mol^{-1})

(2004)

30. On the basis of the information available from the reaction,

$4/3Al + O_2 \rightarrow 2/3Al_2O_3,\ \Delta G = -827$ kJ mol^{-1} of O_2, the minimum e.m.f. required to carry out an electrolysis of Al_2O_3 is (F = 96500 C mol^{-1})

(a) 2.14 V (b) 4.28 V

(c) 6.42 V (d) 8.56 V *(2003)*

31. For the disproportionation of copper $2Cu^+ \rightarrow Cu^{2+} + Cu$, $E°$ is (Given : $E°$ for Cu^{2+}/Cu is 0.34 V and $E°$ for Cu^{2+}/Cu^+ is 0.15 V)

(a) 0.49 V (b) –0.19 V

(c) 0.38 V (d) –0.38 V *(2000)*

32. $E°$ for the cell, $Zn \mid Zn^{2+}_{(aq)} \| Cu^{2+}_{(aq)} \mid Cu$ 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

(a) 10^{+18} (b) 10^{+17}

(c) 10^{-28} (d) 10^{+37} *(1997)*

3.4 Conductance of Electrolytic Solutions

33. Following limiting molar conductivities are given as :

$\lambda°_{m(H_2SO_4)} = x$ S cm^2 mol^{-1}

$\lambda°_{m(K_2SO_4)} = y$ S cm^2 mol^{-1}

$\lambda°_{m(CH_3COOK)} = z$ S cm^2 mol^{-1}

$\lambda°_m$ (in S cm^2 mol^{-1}) for CH_3COOH will be

(a) $x - y + 2z$ (b) $x + y - z$

(c) $x - y + z$ (d) $\frac{(x-y)}{2} + z$

(Odisha NEET 2019)

34. The molar conductivity of a 0.5 mol/dm^3 solution of $AgNO_3$ with electrolytic conductivity of 5.76×10^{-3} S cm^{-1} at 298 K is

(a) 2.88 S cm^2/mol (b) 11.52 S cm^2/mol

(c) 0.086 S cm^2/mol (d) 28.8 S cm^2/mol

(NEET-II 2016)

35. At 25 °C molar conductance of 0.1 molar aqueous solution of ammonium hydroxide is 9.54 ohm^{-1} cm^2 mol^{-1} and at infinite dilution its molar conductance is 238 ohm^{-1} cm^2 mol^{-1}. The degree of ionisation of ammonium hydroxide at the same concentration and temperature is

(a) 4.008% (b) 40.800%

(c) 2.080% (d) 20.800% *(NEET 2013)*

36. Limiting molar conductivity of NH_4OH [*i.e.*, $\Lambda°_{m(NH_4OH)}$] is equal to

(a) $\Lambda°_{m(NH_4Cl)} + \Lambda°_{m(NaCl)} - \Lambda°_{m(NaOH)}$

(b) $\Lambda°_{m(NaOH)} + \Lambda°_{m(NaCl)} - \Lambda°_{m(NH_4Cl)}$

(c) $\Lambda°_{m(NH_4OH)} + \Lambda°_{m(NH_4Cl)} - \Lambda°_{m(HCl)}$

(d) $\Lambda°_{m(NH_4Cl)} + \Lambda°_{m(NaOH)} - \Lambda°_{m(NaCl)}$ *(2012)*

37. Molar conductivities ($\Lambda°_m$) at infinite dilution of NaCl, HCl and CH_3COONa are 126.4, 425.9 and 91.0 S cm^2 mol^{-1} respectively. ($\Lambda°_m$) for CH_3COOH will be

(a) 425.5 S cm^2 mol^{-1} (b) 180.5 S cm^2 mol^{-1}
(c) 290.8 S cm^2 mol^{-1} (d) 390.5 S cm^2 mol^{-1}
(Mains 2012, 1997)

38. An increase in equivalent conductance of a strong electrolyte with dilution is mainly due to
(a) increase in ionic mobility of ions
(b) 100% ionisation of electrolyte at normal dilution
(c) increase in both *i.e.*, number of ions and ionic mobility of ions
(d) increase in number of ions. *(2010)*

39. Which of the following expressions correctly represents the equivalent conductance at infinite dilution of $Al_2(SO_4)_3$? Given that $\overset{\circ}{\Lambda}_{Al^{3+}}$ and $\overset{\circ}{\Lambda}_{SO_4^{2-}}$ are the equivalent conductances at infinite dilution of the respective ions.
(a) $2\overset{\circ}{\Lambda}_{Al^{3+}} + 3\overset{\circ}{\Lambda}_{SO_4^{2-}}$
(b) $\overset{\circ}{\Lambda}_{Al^{3+}} + \overset{\circ}{\Lambda}_{SO_4^{2-}}$
(c) $(\overset{\circ}{\Lambda}_{Al^{3+}} + \overset{\circ}{\Lambda}_{SO_4^{2-}}) \times 6$
(d) $\frac{1}{3}\overset{\circ}{\Lambda}_{Al^{3+}} + \frac{1}{2}\overset{\circ}{\Lambda}_{SO_4^{2-}}$ *(Mains 2010)*

40. The equivalent conductance of $M/32$ solution of a weak monobasic acid is 8.0 mho cm^2 and at infinite dilution is 400 mho cm^2. The dissociation constant of this acid is
(a) 1.25×10^{-6} (b) 6.25×10^{-4}
(c) 1.25×10^{-4} (d) 1.25×10^{-5} *(2009)*

41. Kohlrausch's law states that at
(a) infinite dilution, each ion makes definite contribution to conductance of an electrolyte whatever be the nature of the other ion of the electrolyte
(b) infinite dilution, each ion makes definite contribution to equivalent conductance of an electrolyte, whatever be the nature of the other ion of the electrolyte
(c) finite dilution, each ion makes definite contribution to equivalent conductance of an electrolyte, whatever be the nature of the other ion of the electrolyte
(d) infinite dilution each ion makes definite contribution to equivalent conductance of an electrolyte depending on the nature of the other ion of the electrolyte. *(2008)*

42. Equivalent conductances of Ba^{2+} and Cl^- ions are 127 and 76 ohm^{-1} cm^{-1} eq^{-1} respectively. Equivalent conductance of $BaCl_2$ at infinite dilution is
(a) 139.5 (b) 101.5
(c) 203 (d) 279 *(2000)*

43. The specific conductance of a 0.1 N KCl solution at 23°C is 0.012 ohm^{-1} cm^{-1}. The resistance of cell containing the solution at the same temperature was found to be 55 ohm. The cell constant will be
(a) 0.918 cm^{-1} (b) 0.66 cm^{-1}
(c) 1.142 cm^{-1} (d) 1.12 cm^{-1} *(1999)*

44. On heating one end of a piece of a metal, the other end becomes hot because of
(a) energised electrons moving to the other end
(b) minor perturbation in the energy of atoms
(c) resistance of the metal
(d) mobility of atoms in the metal. *(1995)*

3.5 Electrolytic Cells and Electrolysis

45. On electrolysis of dil. sulphuric acid using platinum (Pt) electrode, the product obtained at anode will be
(a) hydrogen gas (b) oxygen gas
(c) H_2S gas (d) SO_2 gas. *(NEET 2020)*

46. The number of Faradays (F) required to produce 20 g of calcium from molten $CaCl_2$ (Atomic mass of Ca = 40 g mol^{-1}) is
(a) 1 (b) 2
(c) 3 (d) 4 *(NEET 2020)*

47. During the electrolysis of molten sodium chloride, the time required to produce 0.10 mol of chlorine gas using a current of 3 amperes is
(a) 55 minutes (b) 110 minutes
(c) 220 minutes (d) 330 minutes.
(NEET-II 2016)

48. The number of electrons delivered at the cathode during electrolysis by a current of 1 ampere in 60 seconds is (charge on electron = 1.60×10^{-19} C)
(a) 6×10^{23} (b) 6×10^{20}
(c) 3.75×10^{20} (d) 7.48×10^{23}
(NEET-II 2016)

49. When 0.1 mol MnO_4^{2-} is oxidised, the quantity of electricity required to completely oxidise MnO_4^{2-} to MnO_4^- is
(a) 96500 C (b) 2 × 96500 C
(c) 9650 C (d) 96.50 C *(2014)*

50. The weight of silver (at. wt. = 108) displaced by a quantity of electricity which displaces 5600 mL of O_2 at STP will be
(a) 5.4 g (b) 10.8 g
(c) 54.0 g (d) 108.0 g *(2014)*

51. How many grams of cobalt metal will be deposited when a solution of cobalt(II) chloride is electrolyzed with a current of 10 amperes for 109 minutes? (1 Faraday = 96,500 C; Atomic mass of Co = 59 u)
(a) 4.0 (b) 20.0
(c) 40.0 (d) 0.66
(Karnataka NEET 2013)

52. Al_2O_3 is reduced by electrolysis at low potentials and high currents. If 4.0×10^4 amperes of current is passed through molten Al_2O_3 for 6 hours, what mass of aluminium is produced? (Assume 100% current efficiency, at. mass of Al = 27 g mol^{-1})

(a) 8.1×10^4 g (b) 2.4×10^5 g
(c) 1.3×10^4 g (d) 9.0×10^3 g *(2009)*

53. 4.5 g of aluminium (at. mass 27 amu) is deposited at cathode from Al^{3+} solution by a certain quantity of electric charge. The volume of hydrogen produced at STP from H^+ ions in solution by the same quantity of electric charge will be

(a) 44.8 L (b) 22.4 L
(c) 11.2 L (d) 5.6 L *(2005)*

54. In electrolysis of NaCl when Pt electrode is taken then H_2 is liberated at cathode while with Hg cathode it forms sodium amalgam. The reason for this is

(a) Hg is more inert than Pt
(b) more voltage is required to reduce H^+ at Hg than at Pt
(c) Na is dissolved in Hg while it does not dissolve in Pt
(d) conc. of H^+ ions is larger when Pt electrode is taken. *(2002)*

55. A 5 ampere current is passed through a solution of zinc sulphate for 40 minutes. The amount of zinc deposited at the cathode is

(a) 0.4065 g (b) 65.04 g
(c) 40.65 g (d) 4.065 g *(1996)*

56. Sodium is made by the electrolysis of a molten mixture of about 40% NaCl and 60% $CaCl_2$ because

(a) Ca^{++} can reduce NaCl to Na
(b) Ca^{++} can displace Na from NaCl
(c) $CaCl_2$ helps in conduction of electricity
(d) this mixture has a lower melting point than NaCl. *(1995)*

57. When $CuSO_4$ is electrolysed using platinum electrodes,

(a) copper is liberated at cathode, sulphur at anode
(b) copper is liberated at cathode, oxygen at anode
(c) sulphur is liberated at cathode, oxygen at anode
(d) oxygen is liberated at cathode, copper at anode. *(1993)*

58. On electrolysis of dilute sulphuric acid using platinum electrodes, the product obtained at the anode will be

(a) hydrogen (b) oxygen
(c) hydrogen sulphide (d) sulphur dioxide. *(1992)*

3.7 Fuel Cells

59. A device that converts energy of combustion of fuels like hydrogen and methane, directly into electrical energy is known as

(a) dynamo (b) Ni-Cd cell
(c) fuel cell (d) electrolytic cell. *(2015, Cancelled)*

60. The efficiency of a fuel cell is given by

(a) $\Delta G/\Delta S$ (b) $\Delta G/\Delta H$
(c) $\Delta S/\Delta G$ (d) $\Delta H/\Delta G$ *(2007)*

3.8 Corrosion

61. Zinc can be coated on iron to produce galvanized iron but the reverse is not possible. It is because

(a) zinc is lighter than iron
(b) zinc has lower melting point than iron
(c) zinc has lower negative electrode potential than iron
(d) zinc has higher negative electrode potential than iron. *(NEET-II 2016)*

62. The most convenient method to protect the bottom of ship made of iron is

(a) coating it with red lead oxide
(b) white tin plating
(c) connecting it with Mg block
(d) connecting it with Pb block. *(2001)*

63. To protect iron against corrosion, the most durable metal plating on it, is

(a) copper plating (b) zinc plating
(c) nickel plating (d) tin plating. *(1994)*

ANSWER KEY

1.	(b)	**2.**	(c)	**3.**	(a)	**4.**	(c)	**5.**	(b)	**6.**	(c)	**7.**	(d)	**8.**	(a)	**9.**	(a)	**10.**	(d)
11.	(b)	**12.**	(d)	**13.**	(d)	**14.**	(b)	**15.**	(c)	**16.**	(b)	**17.**	(d)	**18.**	(b)	**19.**	(a)	**20.**	(c)
21.	(d)	**22.**	(d)	**23.**	(c)	**24.**	(b)	**25.**	(a)	**26.**	(d)	**27.**	(a)	**28.**	(c)	**29.**	(d)	**30.**	(a)
31.	(c)	**32.**	(d)	**33.**	(d)	**34.**	(b)	**35.**	(a)	**36.**	(d)	**37.**	(d)	**38.**	(a)	**39.**	(b)	**40.**	(d)
41.	(a)	**42.**	(a)	**43.**	(b)	**44.**	(a)	**45.**	(b)	**46.**	(a)	**47.**	(b)	**48.**	(c)	**49.**	(c)	**50.**	(d)
51.	(b)	**52.**	(a)	**53.**	(d)	**54.**	(b)	**55.**	(d)	**56.**	(d)	**57.**	(b)	**58.**	(b)	**59.**	(c)	**60.**	(b)
61.	(d)	**62.**	(b)	**63.**	(b)														

Hints & Explanations

1. (b) : Higher the value of $E°_{red}$, stronger is the oxidising power. Thus, the decreasing order of reducing power of the metal is K > Al > Cr > Ag.

2. (c) : $E°_{cell} = E°_{O.P.} + E°_{R.P.}$

$= 0.76 + 0.34 = 1.10$ V

3. (a) : Less positive the value of reduction potential, stronger will be the reducing agent thus, I^- is strongest reducing agent. More positive, the value of reduction potential shows good oxidising properties thus, strongest oxidising agent is F_2.

4. (c) : More negative the value of reduction potential, stronger will be the reducing agent.

So, Z (–3.0 V) > X (–1.2 V) > Y (+ 0.5 V)

5. (b) : $E°_{cell} = E°_{cathode} - E°_{anode}$

$= 0.15 - (-0.74) = 0.15 + 0.74 = 0.89$ V

6. (c) : Since the reduction potential of Fe^{3+}/Fe^{2+} is greater than that of I_2/I^-, Fe^{3+} will be reduced and I^- will be oxidised.

$2Fe^{3+} + 2I^- \ 2Fe^{2+} + I_2$

7. (d) : EMF of a cell = Reduction potential of cathode – Reduction potential of anode

= Reduction potential of cathode + Oxidation potential of anode

= Oxidation potential of anode – Oxidation potential of cathode.

8. (a) : $[Fe(CN)_6]^{3-} \rightarrow [Fe(CN)_6]^{4-}, E° = +0.35$ V

$Fe^{3+} \rightarrow Fe^{2+}; E° = +0.77$ V

Higher the +ve reduction potential, stronger will be the oxidising agent. Oxidising agent oxidises other compounds and gets itself reduced easily. Thus, Fe^{3+} is the strongest oxidising agent.

9. (a) : From the given expression :

At anode : $A \longrightarrow A^+ + e^-$ (oxidation)

At cathode : $B^+ + e^- \longrightarrow B$ (reduction)

Overall reaction is : $A + B^+ \longrightarrow A^+ + B$

10. (d) : $Fe^{2+} + 2e^- \longrightarrow Fe ; E° = -0.441$ V ... (i)

$Fe^{3+} + e^- \longrightarrow Fe^{2+} ; E° = 0.771$ V ... (ii)

$Fe + 2Fe^{3+} \longrightarrow 3Fe^{2+} ; E° = ?$

To get the above equation, (ii) × 2 – (i)

$$\begin{array}{l} 2Fe^{3+} + 2e^- \longrightarrow 2Fe^{2+} ; E° = 0.771 \text{ V} \\ -Fe^{2+} - 2e^- \longrightarrow -Fe ; E° = +0.441 \text{ V} \\ \hline 2Fe^{3+} + Fe \longrightarrow 3Fe^{2+} ; E° = 1.212 \text{ V} \\ \hline \end{array}$$

11. (b) : $E°_{Fe^{2+}/Fe} = -0.44$ V

$E°_{Fe^{3+}/Fe^{2+}} = +0.77$ V

If a cell is constructed combining these two electrodes oxidation occurs at Fe^{2+}/Fe electrode.

At anode : $Fe \rightarrow Fe^{2+} + 2e^-$

At cathode : $[Fe^{3+} + e^- \rightarrow Fe^{2+}] \times 2$

Cell reaction : $Fe + 2Fe^{3+} \rightarrow 3Fe^{2+}$

If Fe^{2+}, Fe^{3+} and Fe blocks are kept together then Fe^{3+} reacts with Fe to yield Fe^{2+} *i.e.*, concentration of Fe^{3+} is decreased and that of Fe^{2+} is increased.

12. (d) : $E°_{Zn/Zn^{2+}} = +0.76$ V

$E°_{Fe/Fe^{2+}} = 0.44 \text{ V} \Rightarrow E°_{Fe^{2+}/Fe} = -0.44$ V

$E°_{cell} = E°_{O.P.} + E°_{R.P.} = +0.76 - 0.44 = +0.32$ V

13. (d) : Since it is a concentration cell and the concentration of H^+ ions in two electrolyte solutions (HCl and CH_3COOH) are different *i.e.*, pH of 0.1 M HCl and 0.1 M CH_3COOH is not same, therefore e.m.f. of this cell will not be zero.

14. (b) : More positive or less negative the reduction potential value, the stronger is the oxidising agent.

15. (c) : A stronger oxidising agent (Cl_2) displaces a weaker oxidising agent (Br_2) from its salt solution.

$2KBr + Cl_2 \rightarrow 2KCl + Br_2$

16. (b) : The standard Gibbs' energy,

$(\Delta G°) = -nFE°_{cell}$. Value of $n = 2$

$\therefore \quad \Delta G° = -2 \times 96500 \times 0.24 = -46320$ J

$= -46.32$ kJ/mol

17. (d) : According to Nernst equation,

$$E_{cell} = E°_{cell} - \frac{0.059}{n} \log Q_c$$

At equilibrium $E_{cell} = 0$, $\therefore \quad Q_c = K_c$

$$E°_{cell} = \frac{0.059}{n} \log K_c \Rightarrow \quad 0.59 = \frac{0.059}{1} \log K_c$$

K_c = antilog 10 $\Rightarrow K_c = 1 \times 10^{10}$

18. (b) : $E_{cell} = E°_{cell} - \dfrac{0.059}{n} \log \dfrac{[Zn^{2+}]}{[Cu^{2+}]}$

$$E_1 = E° - \frac{0.059}{2} \log \frac{0.01}{1}$$

$$E_1 = E° - \frac{0.059}{2} (-2) = E° + 0.059$$

$$E_2 = E° - \frac{0.059}{2} \log \frac{1}{0.01} = E° - 0.059$$

Hence, $E_1 > E_2$.

19. (a) : $\Delta G° = -nFE°_{cell}$

If $E°_{cell}$ = – ve then $\Delta G°$ = +ve *i.e.*; $\Delta G° > 0$.

$\Delta G^\circ = -nRT \log K_{eq}$

For ΔG° = +ve, K_{eq} = –ve *i.e.*, $K_{eq} < 1$.

20. (c) : pH = 7 for water.

$-\log[H^+] = 7 \Rightarrow [H^+] = 10^{-7}$

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

$$E_{cell} = E^\circ_{cell} - \frac{0.0591}{2} \log \frac{p_{H_2}}{[H^+]^2}$$

$$0 = 0 - \frac{0.0591}{2} \log \frac{p_{H_2}}{(10^{-7})^2}$$

$$\log \frac{p_{H_2}}{(10^{-7})^2} = 0 \Rightarrow \frac{p_{H_2}}{(10^{-7})^2} = 1 \qquad [\because \log 1 = 0]$$

$p_{H_2} = 10^{-14}$ atm

21. (d) : $\underset{1\text{ atm}}{H_2} \longrightarrow \underset{10^{-10}}{2H^+} + 2e^-$

$$E_{H_2/H^+} = 0 - \frac{0.059}{2} \log \frac{(10^{-10})^2}{1}$$

$E_{H_2/H^+} = +\ 0.59$ V

22. (d) : $Mn^{2+} + 2e^- \longrightarrow Mn;\ E^\circ = -1.18$ V ... (i)

$2Mn^{2+} \longrightarrow 2Mn^{3+} + 2e^-;\ E^\circ = -1.51$ V ... (ii)

By adding equation (i) and (ii), we get equation for the cell,

$3Mn^{2+} \longrightarrow Mn + 2Mn^{3+};\ E^\circ = -2.69$ V

Since the E° value is negative, so the process is non-spontaneous as ΔG° is positive.

23. (c) : $\Delta G^\circ = -nFE^\circ$

$F = 96500$, $\Delta G^\circ = +\ 960 \times 10^3$ J/mol

$$\frac{2}{3} Al_2O_3 \rightarrow \frac{4}{3} Al + O_2$$

Total number of Al atoms in $Al_2O_3 = \frac{2}{3} \times 2 = \frac{4}{3}$

$Al^{3+} + 3e^- \longrightarrow Al$

As $3e^-$ change occur for each Al-atom

$\therefore$ total $n = \frac{4}{3} \times 3 = 4$

$$E^\circ = -\frac{\Delta G^\circ}{nF} = -\frac{960 \times 1000}{4 \times 96500}$$

$\Rightarrow E^\circ = -2.48 \approx -2.5$ V

24. (b) : $Cu^{2+}_{(aq)} + e^- \longrightarrow Cu^+_{(aq)};\ E^\circ_1 = 0.15$ V

$Cu^+_{(aq)} + e^- \longrightarrow Cu_{(s)};\ E_2^\circ = 0.50$ V

$Cu^{2+} + 2e^- \longrightarrow Cu;\ E^\circ = ?$

Now, $\Delta G^\circ = \Delta G_1^\circ + \Delta G_2^\circ$

or, $-nFE^\circ = -n_1FE_1^\circ - n_2FE_2^\circ$

$$\text{or, } E^\circ = \frac{n_1E_1^\circ + n_2E_2^\circ}{n} = \frac{1 \times 0.15 + 1 \times 0.50}{2} = 0.325 \text{ V}$$

25. (a) : The cell reaction can be written as

$Cu + 2Ag^+ \longrightarrow Cu^{2+} + 2Ag$

We know, $\Delta G^\circ = -nFE^\circ_{cell}$

$= -\ 2 \times 96500 \times 0.46 = -\ 88780$ J

$= -\ 88.78$ kJ $\approx -\ 89$ kJ

26. (d) : Given,

$Cu^{2+} + 2e^- \longrightarrow Cu;\ E_1^\circ = 0.337$ V

$Cu^{2+} + e^- \longrightarrow Cu^+;\ E_2^\circ = 0.153$ V

The required reaction is

$Cu^+ + e^- \longrightarrow Cu;\quad E_3^\circ = ?$

Applying, $\Delta G^\circ = -nFE^\circ$, $\Delta G_3^\circ = \Delta G_1^\circ - \Delta G_2^\circ$

$-(n_3FE_3^\circ) = -(n_1FE_1^\circ) - (-n_2FE_2^\circ)$

or $E_3^\circ = 2 \times E_1^\circ - E_2^\circ$

or $E_3^\circ = (2 \times 0.337) - 0.153 = 0.52$ V

27. (a) : $C_5H_{12(g)} + 8O_{2(g)} \longrightarrow 5CO_{2(g)} + 6H_2O_{(l)}$

$\Delta G^\circ = [(-394.4 \times 5) + (-237.2 \times 6)] - [(-8.2) + (8 \times 0)]$

$= -3387$ kJ

Note that the standard free energy change of elementary substances is taken as zero.

For the fuel cell, the complete cell reaction is :

$C_5H_{12(g)} + 8O_{2(g)} \longrightarrow 5CO_{2(g)} + 6H_2O_{(l)}$

which is the combination of the following two half reactions :

$C_5H_{12(g)} + 10H_2O_{(l)} \longrightarrow 5CO_{2(g)} + 32H^+ + 32e^-$ and

$8O_{2(g)} + 32H^+ + 32e^- \longrightarrow 16H_2O_{(l)}$

Therefore, the number of electrons exchanged is 32 here, *i.e.*, $n = 32$.

$\Delta G^\circ = -nFE^\circ = -\ 3387 \times 10^3$ J

$= -32 \times 96500$ J/Volt $\times E^\circ$

Thus, $E^\circ = 1.0968$ V

28. (c) : For a cell reaction in equilibrium at 298 K,

$$E^\circ_{cell} = \frac{0.0591}{n} \log K_c$$

where, K_c = equilibrium constant, n = number of electrons involved in the electrochemical cell reaction.

Given, $E^\circ_{cell} = 0.46$ V, $n = 2$

$$\therefore 0.46 = \frac{0.0591}{2} \times \log K_c \text{ or, } \log K_c = \frac{2 \times 0.46}{0.0591} = 15.57$$

or, $K_c = 3.7 \times 10^{15} \approx 4 \times 10^{15}$

29. (d) : $E = E^\circ - \frac{0.0591}{n} \log_{10} Q$ at 25°C

At equilibrium, $E = 0$, $Q = K$

$$0 = E^\circ - \frac{0.0591}{n} \log_{10} K$$

$$\text{or, } K = \text{antilog}\left[\frac{nE^\circ}{0.0591}\right]$$

or, $K = \text{antilog}\left[\frac{2\times 0.295}{0.0591}\right] = \text{antilog}\left[\frac{0.590}{0.0591}\right]$

$= \text{antilog } 10 = 1 \times 10^{10}$

30. (a) : $\Delta G^\circ = -nFE^\circ$

$$E^\circ = \frac{\Delta G^\circ}{-nF} = \frac{-827000}{-4\times 96500} = 2.14\text{ V}$$

$$\left(\because 1\text{ Al} \equiv 3e^-, \frac{4}{3}\text{Al} = \frac{4}{3}\times 3e^- = 4e^-\right)$$

31. (c) : For the reaction, $2Cu^+ \longrightarrow Cu^{2+} + Cu$ the cathode is Cu^+/Cu and anode is Cu^+/Cu^{2+}.

Given, $Cu^{2+} + 2e^- \longrightarrow Cu;\ E_1^\circ = 0.34\text{ V}$ ...(1)

$Cu^{2+} + e^- \longrightarrow Cu^+;\ E_2^\circ = 0\cdot 15\text{ V}$ ...(2)

$Cu^+ + e^- \longrightarrow Cu;\quad E_3^\circ = ?$ ...(3)

Now $\Delta G_1^\circ = -nFE_1^\circ = -2\times 0\cdot 34\times F = -0.68\text{ V}$

$\Delta G_2^\circ = -1\times 0\cdot 15\times F,\ \Delta G_3^\circ = -1\times E_3^\circ \times F,$

Again $\Delta G_1^\circ = \Delta G_2^\circ + \Delta G_3^\circ$

$\Rightarrow -0.68\,F = -0.15\,F - E^\circ_3 \times F$

$\Rightarrow E^\circ_3 = 0\cdot 68 - 0\cdot 15 = 0\cdot 53\text{V}$

As, $E^\circ_{cell} = E^\circ_{cathode(Cu^+/Cu)} - E^\circ_{anode\,(Cu^{2+}/Cu^+)}$

$= 0.53 - 0.15 = 0.38\text{ V}$

32. (d) : Nernst equation is

$$E = E^\circ - \frac{0.059}{2}\log K$$

$\Rightarrow E^\circ = \frac{0.059}{2}\log K$ ($E = 0$ at equilibrium condition)

$\Rightarrow 1.1 = \frac{0.059}{2}\log K \Rightarrow K = 1.9 \times 10^{+37}$

33. (d) : According to Kohlrausch's law,

λ°_m for $CH_3COOH = \lambda^\circ_{CH_3COO^-} + \lambda^\circ_{H^+}$

λ° for $H_2SO_4 = 2\lambda^\circ_{H^+} + \lambda^\circ_{SO_4^{2-}} = x\text{ S cm}^2\text{mol}^{-1}$ ...(i)

λ° for $K_2SO_4 = 2\lambda^\circ_{K^+} + \lambda^\circ_{SO_4^{2-}} = y\text{ S cm}^2\text{mol}^{-1}$ ...(ii)

λ° for $CH_3COOK = \lambda^\circ_{CH_3COO^-} + \lambda^\circ_{K^+} = z\text{ S cm}^2\text{mol}^{-1}$ ...(iii)

On adding equation (i) and 2 × (iii) and subtracting (ii), we get

$$2\lambda^\circ_{H^+} + \lambda^\circ_{SO_4^{2-}} + 2\lambda^\circ_{CH_3COO^-} + 2\lambda^\circ_{K^+} - 2\lambda^\circ_{K^+} - \lambda^\circ_{SO_4^{2-}} = x + 2z - y$$

$$2\lambda^\circ_{H^+} + 2\lambda^\circ_{CH_3COO^-} = x + 2z - y$$

$$\lambda^\circ_{H^+} + \lambda^\circ_{CH_3COO^-} = \frac{(x-y)}{2} + z$$

34. (b) : $\Lambda_m = \frac{\kappa\times 1000}{\text{Molarity}(M)}$

$= \frac{5.76\times 10^{-3}\text{ S cm}^{-1}\times 1000}{0.5\text{ mol cm}^{-3}} = 11.52\text{ S cm}^2\text{ mol}^{-1}$

35. (a) : Degree of dissociation

$$(\alpha) = \frac{\text{Molar conductivity at conc. } C(\Lambda_m^c)}{\text{Molar conductivity at infinite dilution } (\Lambda_m^\infty)}$$

$$\alpha = \frac{9.54\ \Omega^{-1}\text{ cm}^2\text{ mol}^{-1}}{238\ \Omega^{-1}\text{ cm}^2\text{ mol}^{-1}} = 0.04008 = 4.008\%$$

36. (d)

37. (d) : $\Lambda^\circ_{NaCl} = 126.4\text{ S cm}^2\text{ mol}^{-1}$

$\Lambda^\circ_{HCl} = 425.9\text{ S cm}^2\text{ mol}^{-1}$

$\Lambda^\circ_{CH_3COONa} = 91.0\text{ S cm}^2\text{ mol}^{-1}$

$\Lambda^\circ_{CH_3COOH} = \Lambda^\circ_{CH_3COONa} + \Lambda^\circ_{HCl} - \Lambda^\circ_{NaCl}$

$= 91.0 + 425.9 - 126.4$

$= 390.5\text{ S cm}^2\text{ mol}^{-1}$

38. (a) : Strong electrolytes are completely ionised at all concentrations. On increasing dilution, the no. of ions remains the same but the ionic mobility increases and the equivalent conductance increases.

39. (b) : At infinite dilution, when dissociation is complete, each ion makes a definite contribution towards molar conductance of the electrolyte irrespective of the nature of the other ion with which it is associated.

Hence, $\overset{\circ}{\Lambda}_{Al_2(SO_4)_3} = \overset{\circ}{\Lambda}_{Al^{3+}} + \overset{\circ}{\Lambda}_{SO_4^{2-}}$

40. (d) : Given, $\Lambda = 8\text{ mho cm}^2$, $\Lambda_\infty = 400\text{ mho cm}^2$

Degree of dissociation, $\alpha = \frac{\Lambda}{\Lambda_\infty}$

$\Rightarrow \alpha = \frac{8}{400} = 2\times 10^{-2}$

Dissociation constant, $K = C\alpha$

Given, $C = M/32$

$\therefore K = \frac{1}{32}\times 2\times 10^{-2}\times 2\times 10^{-2} = 1.25\times 10^{-5}$

41. (a) : At infinite dilution, when dissociation is complete each ion makes a definite contribution towards molar conductance of the electrolyte irrespective of the nature of the other ion with which it is associated and that the molar conductance of any electrolyte at infinite dilution is given by the sum of the contributions of two ions. This is called Kohlrausch's law.

$\Lambda_m^\infty = \Lambda_+^\infty + \Lambda_-^\infty$,

where, Λ_+^∞ and Λ_-^∞ are molar ionic conductance at infinite dilution for cation and anion, respectively.

42. (a) : $\lambda_{\infty} = \dfrac{1}{n_+}\lambda_+^{\infty} + \dfrac{1}{n_-}\lambda_-^{\infty}$

So, $\lambda_{\infty}(\mathrm{BaCl_2}) = \dfrac{1}{2}\times\lambda^{\infty}_{\mathrm{Ba^{2+}}} + \dfrac{1}{1}\times\lambda^{\infty}_{\mathrm{Cl^-}}$

$= \dfrac{1}{2}\times 127 + 76 = 63{\cdot}5 + 76 = 139{\cdot}5$

43. (b) : $\kappa = 0.012\ \mathrm{ohm^{-1}\,cm^{-1}}$

$R = 55$ ohm $\Rightarrow C = \dfrac{1}{R} = \dfrac{1}{55}\ \mathrm{ohm^{-1}}$

Cell Constant $\left(\dfrac{l}{a}\right) = \dfrac{\text{Specific Conductance}}{\text{Conductance}}$

$= \dfrac{0.012}{1/55} = 55 \times 0.012 = 0.66\ \mathrm{cm^{-1}}$

44. (a) : Conductivity of heat in metals is due to the presence of free electrons, which move due to increase in temperature.

45. (b) : During electrolysis of dilute sulphuric acid the following reaction takes place at anode.

$2H_2O_{(l)} \rightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^-;\ E^\circ_{cell} = +1.23$ V

i.e., $O_{2(g)}$ will be liberated at anode.

46. (a) : $\underset{(aq)}{Ca^{2+}} + \underset{\substack{2\,F\\1\,F}}{2e^-} \longrightarrow \underset{\substack{(s)\\(1\text{ mole} = 40\text{ g})\\20\text{ g}}}{Ca}$

Thus, one Faraday is required to produce 20 g of calcium from molten $CaCl_2$.

47. (b) : During the electrolysis of molten sodium chloride,

At cathode : $2Na^+ + 2e^- \longrightarrow 2Na$

At anode : $2Cl^- \longrightarrow Cl_2 + 2e^-$

Net reaction : $2Na^+ + 2Cl^- \longrightarrow 2Na + Cl_2$

According to Faraday's first law of electrolysis,

$w = Z \times I \times t$

$w = \dfrac{E}{96500} \times I \times t$

No. of moles of Cl_2 gas × Mol. wt. of Cl_2 gas

$= \dfrac{\text{Eq. wt. of } Cl_2 \text{ gas} \times I \times t}{96500}$

$0.10 \times 71 = \dfrac{35.5 \times 3 \times t}{96500}$

$t = \dfrac{0.10 \times 71 \times 96500}{35.5 \times 3} = 6433.33$ sec

$t = \dfrac{6433.33}{60}$ min $= 107.22$ min ≈ 110 min

48. (c) : $Q = I \times t$

$Q = 1 \times 60 = 60$ C

Now, 1.60×10^{-19} C $\equiv 1$ electron

$\therefore\quad 60\ \mathrm{C} \equiv \dfrac{\square 60}{1.6 \times 10^{-19}} = 3.75 \times 10^{20}$ electrons

49. (c) : The oxidation reaction is

$\underset{0.1\text{ mol}}{\overset{+6}{Mn}O_4^{2-}} \longrightarrow \overset{+7}{Mn}O_4^- + \underset{0.1\text{ mol}}{e^-}$

$Q = 0.1 \times F = 0.1 \times 96500$ C $= 9650$ C

50. (d) : According to Faraday's second law,

$\dfrac{W_{Ag}}{E_{Ag}} = \dfrac{W_{O_2}}{E_{O_2}}$ or $\dfrac{W_{Ag}}{108} = \dfrac{\frac{5600}{22400} \times 32}{8}$

or $\dfrac{W_{Ag}}{108} = \dfrac{8}{8} \Rightarrow W_{Ag} = 108$ g

51. (b) : $w = \dfrac{ItE}{96500}$

$= \dfrac{10 \times 109 \times 60 \times 59}{96500 \times 2} = 19.99 \approx 20$ g

52. (a) : Applying $E = Z \times 96500$

$\dfrac{27}{3} = Z \times 96500 \Rightarrow Z = \dfrac{9}{96500}$

Now applying the formula, $w = Z \times I \times t$

$w = \dfrac{9}{96500} \times 4 \times 10^4 \times 6 \times 60 \times 60 = 8.1 \times 10^4$ g

53. (d) : We know that,

1 Faraday charge liberates 1 eq. of substance.

This is the Faraday law.

eq. wt. of Al $= \dfrac{27}{3} = 9$

No. of eq. of Al $= \dfrac{\text{wt. of Al}}{\text{eq. wt.}} = \dfrac{4.5}{9} = 0.5$

No. of Faradays required = 0.5

$\Rightarrow$ No. of eq. of H_2 produced = 0.5 eq.

Volume occupied by 1 eq. of $H_2 = \dfrac{22.4}{2} = 11.2$ L

$\Rightarrow$ Volume occupied by 0.5 eq. of $H_2 = 11.2 \times 0.5$

$= 5.6$ L at STP

54. (b) : When sodium chloride is dissolved in water, it ionises as $NaCl \rightleftharpoons Na^+ + Cl^-$.

Water also dissociates as : $H_2O \rightleftharpoons H^+ + OH^-$

During passing of electric current through this solution using platinum electrode, Na^+ and H^+ ions move towards cathode. However, only H^+ ions are discharged more readily than Na^+ ions because of their low discharge potential (in the electromotive series hydrogen is lower than sodium). These H^+ ions gain electrons and change into neutral atoms.

At cathode $H^+ + e^- \text{——} H, H + H \longrightarrow H_2$

Cl^- and OH^- ions move towards anode. Cl^- ions lose electrons and change into neutral atom.

At anode, $Cl^- - e^- Cl, Cl + Cl\ Cl_2 \longrightarrow$

If mercury is used as cathode, H^+ ions are not discharged at mercury cathode because mercury has a high hydrogen overvoltage. Na^+ ions are discharged at the cathode in preference to H^+ ions, yielding sodium, which dissolves in mercury to form sodium amalgam.

At cathode : $Na^+ + e^- Na$

55. (d) : Current (I) = 5 ampere and

time (t) = 40 minutes = 2400 seconds.

Amount of electricity passed (Q) = $I \times t$

$$= 5 \times 2400 = 12000 \text{ C}$$

Now, $Zn^{2+} + 2e^- \text{——} Zn$ (1 mole = 65.39 g)

Since, two charges (*i.e.*, 2 × 96500 C) deposits 65.39 g of zinc, therefore 12000 C will deposit

$$= \frac{65.39 \times 12000}{2 \times 96500} = 4.065 \text{ g of zinc}$$

56. (d) : Sodium is obtained by electrolytic reduction of its chloride. Melting point of chloride of sodium is high so in order to lower its melting point, calcium chloride is added to it.

57. (b) : $CuSO_4 \rightleftharpoons Cu^{2+} + SO_4^{2-}$

$H_2O \; H^+ \rightleftharpoons + \; OH^-$

At cathode : $Cu^{2+} + 2e^- \rightarrow Cu$

At anode : $4OH^- \rightarrow 2H_2O + O_2 + 4e^-$

58. (b) : During electrolysis of dilute sulphuric acid, product obtained at anode will be oxygen.

At anode : $4OH^- \rightleftharpoons 2H_2O + 2O_2 + 4e^-$

59. (c)

60. (b) : The thermal efficiency, η of a fuel conversion device is the amount of useful energy produced relative to the change in enthalpy, ΔH between the product and feed streams.

$$\eta = \frac{\text{useful energy}}{\Delta H}$$

In an ideal case of an electrochemical converter, such as a fuel cell, the change in Gibb's free energy, ΔG of the reaction is available as useful electric energy at that temperature of the conversion.

Hence, $\eta_{\text{ideal}} = \dfrac{\Delta G}{\Delta H}$

61. (d) : Reduction potential values of $E°_{Zn^{2+}/Zn} = -0.76$ V and $E°_{Fe^{2+}/Fe} = -0.44$ V

Thus, due to higher negative electrode potential value of zinc than iron, iron cannot be coated on zinc.

62. (b) : The most convenient method to protect the bottom of the ship made of iron is white tin plating preventing the build up of barnacles.

63. (b)