Chapter 17. Electrochemistry

1. 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) $E_2 = 0^1 E_1$ (d) $E_1 = E_2$
(NEET 2017)

2. 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)

3. 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)

4. 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)

5. 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)

6. 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)*

7. 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)

8. 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)

9. 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)*

10. 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)*

11. 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)

12. A button cell used in watches function as following.

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

If half cell potentials are

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

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

The cell potential will be
(a) 0.84 V (b) 1.34 V
(c) 1.10 V (d) 0.42 V
(NEET 2013)

13. 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)

14. Consider the half-cell reduction reaction

$Mn^{2+} + 2e^- \rightarrow Mn,\ E° = -1.18\ V$

$Mn^{2+} \rightarrow Mn^{3+} + e^-,\ E° = -1.51\ V$

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)

15. How many gram 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)

16. Limiting molar conductivity of NH_4OH [*i.e.* $\Lambda^{\circ}_{m(NH_4OH)}$] is equal to

(a) $\Lambda^{\circ}_{m(NH_4Cl)} + \Lambda^{\circ}_{m(NaCl)} - \Lambda^{\circ}_{m(NaOH)}$
(b) $\Lambda^{\circ}_{m(NaOH)} + \Lambda^{\circ}_{m(NaCl)} - \Lambda^{\circ}_{m(NH_4Cl)}$
(c) $\Lambda^{\circ}_{m(NH_4OH)} + \Lambda^{\circ}_{m(NH_4Cl)} - \Lambda^{\circ}_{m(HCl)}$
(d) $\Lambda^{\circ}_{m(NH_4Cl)} + \Lambda^{\circ}_{m(NaOH)} - \Lambda^{\circ}_{m(NaCl)}$ *(2012)*

17. 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 23 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)

18. Molar conductivities (Λ°_m) at infinite dilution of NaCl, HCl and CH_3COONa are 126.4, 425.9 and 91.0 S cm^2 mol^{-1} respectively. (Λ°_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)

19. The Gibb's energy for the decomposition of Al_2O_3 at 500°C is as follows

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

$\Delta_r G = +960\ kJ\ mol^{-1}$

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)

20. Standard electrode potential 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)

21. 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)*

22. 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)*

23. 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)

24. 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 Gibb's 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)*

25. 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)*

26. 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)*

27. 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)

28. 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)*

29. 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)

30. 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)

31. On the basis of the following $E°$ values, the strongest oxidizing agent is
$[Fe(CN)_6]^{4-} \rightarrow [Fe(CN)_6]^{3-} + e^{-1}$; $E° = -0.35$ V
$Fe^{2+} \rightarrow Fe^{3+} + e^{-1}$; $E° = -0.77$ V
(a) Fe^{3+} (b) $[Fe(CN)_6]^{3-}$
(c) $[Fe(CN)_6]^{4-}$ (d) Fe^{2+} *(2008)*

32. 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)

33. 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)*

34. 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)

35. 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)*

36. A hypothetical electrochemical cell is shown below.

$A \mid A^+ (x\mathrm{M}) \parallel B^+ (y\mathrm{M}) \mid B$

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

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

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

(c) $A^+ + e^- \rightarrow A$; $B^+ + e^- \rightarrow B$

(d) the cell reaction cannot be predicted.

(2006)

37. $E^\circ_{\mathrm{Fe^{2+}/Fe}} = -0.441$ V and $E^\circ_{\mathrm{Fe^{3+}/Fe^{2+}}} = 0.771$ V, the standard EMF of the reaction $\mathrm{Fe + 2Fe^{3+} \rightarrow 3Fe^{2+}}$ will be

(a) 0.111 V (b) 0.330 V

(c) 1.653 V (d) 1.212 V *(2006)*

38. 4.5 g of aluminium (at. mass 27 amu) is deposited at cathode from $\mathrm{Al^{3+}}$ solution by a certain quantity of electric charge. The volume of hydrogen produced at STP from $\mathrm{H^+}$ 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)*

39. The mass of carbon anode consumed (giving only carbon dioxide) in the production of 270 kg of aluminium metal from bauxite by the Hall process is

(a) 270 kg (b) 540 kg

(c) 90 kg (d) 180 kg.

(Atomic mass : Al = 27) *(2005)*

40. 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 $F = 96500\ \mathrm{C\ mol^{-1}}$, $R = 8.314\ \mathrm{J\ K^{-1}\ mol^{-1}}$)

(2004)

41. The e.m.f. of a Daniell cell at 298 K is E_1.

$$\mathrm{Zn}\,\Big|\begin{matrix}\mathrm{ZnSO_4}\\ (0.01\ \mathrm{M})\end{matrix}\Big\|\begin{matrix}\mathrm{CuSO_4}\\ (1.0\ \mathrm{M})\end{matrix}\Big|\,\mathrm{Cu}$$

When the concentration of $\mathrm{ZnSO_4}$ is 1.0 M and that of $\mathrm{CuSO_4}$ is 0.01 M, the e.m.f. changed to E_2. What is the relationship between E_1 and E_2?

(a) $E_1 > E_2$ (b) $E_1 < E_2$

(c) $E_1 = E_2$ (d) $E_2 = 0 \neq E_1$

(2003)

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

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

(a) 2.14 V (b) 4.28 V

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

43. In electrolysis of NaCl when Pt electrode is taken then $\mathrm{H_2}$ is liberated at cathode while with Hg cathode it forms sodium amalgam

(a) Hg is more inert than Pt

(b) More voltage is required to reduce $\mathrm{H^+}$ at Hg than at Pt

(c) Na is dissolved in Hg while it does not dissolve in Pt

(d) Conc. of $\mathrm{H^+}$ ions is larger when Pt electrode is taken. *(2002)*

44. Standard electrode potentials are $\mathrm{Fe^{2+}/Fe}$; $E^\circ = -0.44$ and $\mathrm{Fe^{3+}/Fe^{2+}}$; $E^\circ = 0.77$ $\mathrm{Fe^{2+}}$, $\mathrm{Fe^{3+}}$ and Fe blocks are kept together, then

(a) $\mathrm{Fe^{3+}}$ increases

(b) $\mathrm{Fe^{3+}}$ decreases

(c) $\mathrm{Fe^{2+}/Fe^{3+}}$ remains unchanged

(d) $\mathrm{Fe^{2+}}$ decreases. *(2001)*

45. Equivalent conductances of $\mathrm{Ba^{2+}}$ and $\mathrm{Cl^-}$ ions are 127 and 76 $\mathrm{ohm^{-1}\ cm^{-1}\ eq^{-1}}$ respectively. Equivalent conductance of $\mathrm{BaCl_2}$ at infinite dilution is

(a) 139.5 (b) 101.5

(c) 203 (d) 279 *(2000)*

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

(a) 0.49 V (b) –0.19 V

(c) 0.38 V (d) –0.38 V *(2000)*

47. The specific conductance of a 0.1 N KCl solution at 23°C is 0.012 $\mathrm{ohm^{-1}\ cm^{-1}}$. The resistance of cell containing the solution at the same temperature was found to be 55 ohm. The cell constant will be

(a) $0.918\ \mathrm{cm^{-1}}$ (b) $0.66\ \mathrm{cm^{-1}}$

(c) $1.142\ \mathrm{cm^{-1}}$ (d) $1.12\ \mathrm{cm^{-1}}$

(1999)

48. For the cell reaction,

$\mathrm{Cu^{2+}}\ (C_1.aq) + \mathrm{Zn}_{(s)} = \mathrm{Zn^{2+}}(C_2.aq) + \mathrm{Cu}_{(s)}$

of an electrochemical cell, the change in free energy ΔG at a given temperature is a function of

(a) $\ln (C_2)$ (b) $\ln (C_2/C_1)$

(c) $\ln (C_1)$ (d) $\ln (C_1 + C_2)$

(1998)

49. E° for the cell, $\mathrm{Zn} \mid \mathrm{Zn^{2+}}_{(aq)} \parallel \mathrm{Cu^{2+}}_{(aq)} \mid \mathrm{Cu}$ is 1.10V at 25°C, the equilibrium constant for the reaction $\mathrm{Zn + Cu^{2+}}_{(aq)} \rightleftharpoons \mathrm{Cu + Zn^{2+}}_{(aq)}$ is the order of

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

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

50. The molar conductances of NaCl, HCl and CH_3COONa at infinite dilution are 126.45, 426.16 and 91 ohm^{-1} cm^2 mol^{-1} respectively. The molar conductance of CH_3COOH at infinite dilution (Λ_m^∞) is

(a) 698.28 ohm^{-1} cm^2 mol^{-1}
(b) 540.48 ohm^{-1} cm^2 mol^{-1}
(c) 201.28 ohm^{-1} cm^2 mol^{-1}
(d) 390.71 ohm^{-1} cm^2 mol^{-1} *(1997)*

51. 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)*

52. Reduction potential for the following half-cell reactions are

$Zn = Zn^{2+} + 2e^-$; $E° = +0.76$ V;
$Fe = Fe^{2+} + 2e^-$; $E° = +0.44$ 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)*

53. 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)*

54. 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)*

55. 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)*

56. 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)*

Answer Key

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

EXPLANATIONS

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

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

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

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

Hence, $E_1 > E_2$

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

$$= \frac{5.76\times10^{-3}\,\text{S cm}^{-1}\times1000}{0.5\text{ mol cm}^{-3}} = 11.52\text{ S cm}^2\text{ mol}^{-1}$$

3. **(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, $w = Z \times I \times t$

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

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

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

$$0.10\times71 = \frac{35.5\times3\times t}{96500}$$

$$t = \frac{0.10\times71\times96500}{35.5\times3} = 6433.33\text{ sec}$$

$$t = \frac{6433.33}{60}\text{ min} = 107.22\text{ min} \approx 110\text{ min}$$

4. **(a)** : $\Delta G° = -nFE°_{cell}$

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

$\Delta G° = -nRT\log K_{eq}$

For $\Delta G° = +$ve, $K_{eq} = -$ve *i.e.*, $K_{eq} < 1$.

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

$Q = 1 \times 60 = 60$ C

Now, 1.60×10^{-19} C ≡ 1 electron

$$\therefore\ 60\text{ C} \equiv \frac{60}{1.6\times10^{-19}} = 37.5\times10^{19}$$

$$= 3.75\times10^{20}\text{ electrons}$$

6. **(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.

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

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

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

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

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

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

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

8. **(c)**

9. **(c)** : The oxidation reaction is

$$\underset{0.1\text{ mol}}{\overset{+6}{MnO_4^{2-}}} \longrightarrow \overset{+7}{MnO_4^-} + \underset{0.1\text{ mol}}{e^-}$$

$Q = 0.1 \times F = 0.1 \times 96500\text{ C} = 9650\text{ C}$

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

$$\frac{W_{Ag}}{E_{Ag}} = \frac{W_{O_2}}{E_{O_2}} \quad\text{or}\quad \frac{W_{Ag}}{108} = \frac{\frac{5600}{22400}\times32}{8}$$

$$\text{or}\quad \frac{W_{Ag}}{108} = \frac{8}{8} \Rightarrow W_{Ag} = 108\text{ g}$$

11. **(a)** : Degree of dissociation

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

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

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

$= 0.76 + 0.34 = 1.10$ V

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

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

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

14. (d) : $Mn^{2+} + 2e^- \rightarrow Mn$; $E° = -1.18$ V

$2Mn^{2+} \rightarrow 2Mn^{3+} + 2e^-$; $E° = -1.51$ V

For the cell,

$3Mn^{2+} \rightarrow Mn + 2Mn^{3+}$; $E° = -2.69$ V

Since the $E°$ value is negative, so the process is non-spontaneous.

15. (b) : $W = \frac{ItE}{96500}$

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

16. (d)

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

18. (d) : $\Lambda°_{NaCl} = 126.4$ S cm^2 mol^{-1}

$\Lambda°_{HCl} = 425.9$ S cm^2 mol^{-1}

$\Lambda°_{CH_3COONa} = 91.0$ S cm^2 mol^{-1}

$\Lambda°_{CH_3COOH} = \Lambda°_{CH_3COONa} + \Lambda°_{HCl} - \Lambda°_{NaCl}$

$= 91.0 + 425.9 - 126.4 = 390.5$ S cm^2 mol^{-1}

19. (c) : $\Delta G° = -nFE°$

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

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

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

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

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

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

$$E° = -\frac{\Delta G°}{nF} = -\frac{960 \times 1000}{4 \times 96500} \Rightarrow E° = -2.48 \approx -2.5 \text{ V}$$

20. (c) : As the electrode potential drops, reducing power increases.

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

21. (b) : $Cu^{2+}_{(aq)} + e^- \rightarrow Cu^+_{(aq)}$; $E_1° = 0.15$ V

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

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

Now, $\Delta G° = \Delta G_1° + \Delta G_2°$

or, $-nFE° = -n_1FE_1° - n_2FE_2°$

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

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

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

23. (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.

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

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

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

We know, $\Delta G° = -nFE°_{cell}$

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

$= -88.780$ kJ ≈ -89 kJ

25. (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 conduction increases.

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

27. (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.

28. (d) : Given,

$Cu^{2+} + 2e^- \rightarrow Cu$; $E_1° = 0.337$ V

$Cu^{2+} + e^- \rightarrow Cu^+$; $E_2° = 0.153$ V

The required reaction is

$Cu^+ + e^- \rightarrow Cu$; $E_3° = ?$

Applying , $\Delta G° = -nFE°$,

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

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

or $E_3° = 2 \times E_1° - E_2°$

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

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

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

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

$$W = \frac{9}{96500} \times 4 \times 10^4 \times 6 \times 60 \times 60 = 8.1 \times 10^4 \text{ g}$$

30. (d) : Given, $\Lambda = 8$ mho cm^2

$\Lambda_\infty = 400$ mho cm^2

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

$$\Rightarrow \quad \alpha = \frac{8}{400} = 2\times10^{-2}$$

Dissociation constant, $K = C\alpha^2$
Given, $C = M/32$

$$\therefore \quad K = \frac{1}{32}\times 2\times10^{-2}\times 2\times10^{-2}$$

$K = 1.25 \times 10^{-5}$

31. (a) : $[Fe(CN)_6]^{3-} \rightarrow [Fe(CN)_6]^{4-}$, $E° = +0.35V$
$Fe^{3+} \rightarrow Fe^{2+}$; $E° = +0.77V$
Higher the +ve reduction potential, stronger will be the oxidising agent. Oxidising agent oxidises other compounds and gets itself reduced easily.

32. (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.
$\Lambda_m^\infty = \Lambda_+^\infty + \Lambda_-^\infty$, Λ_+^∞ and Λ_-^∞ are molar ionic conductance at infinite dilution for cations and anions, respectively.

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

$$\Delta G° = [(-394.4\times5) + (-237.2\times6)] - [(-8.2) + (8\times0)]$$
$$= -3387.5 \text{ kJ}$$

Note that the standard free energy change of elementary substances is taken as zero.
For the fuel cell, the complete cell reaction is :
$C_5H_{12(g)} + 8O_{2(g)} \rightarrow 5CO_{2(g)} + 6H_2O_{(l)}$
Which is the combination of the following two half reactions :
$C_5H_{12(g)} + 10H_2O_{(l)} \rightarrow 5CO_{2(g)} + 32H^+ + 32e$
and $8O_{2(g)} + 32H^+ + 32e \rightarrow 16H_2O_{(l)}$
Therefore, the number of electrons exchanged is 32 here, means $n = 32$. This is the trickiest part of the problem.

$$\Delta G° = -nFE° = -3387.5\times10^3 \text{ J}$$
$$= -32\times96500 \text{ J/Volt}\times E°$$

Thus $E° = 1.09698$ V

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

$$E°_{cell} = \frac{0.0591}{n}\log K_C$$

where K_C = equilibrium constant, n = number of electrons involved in the electrochemical cell reaction.
Given, $E°_{cell} = 0.46$ V, $n = 2$

$$\therefore \quad 0.46 = \frac{0.0591}{2}\times\log K_C$$

or, $\log K_C = \dfrac{2\times0.46}{0.0591} = 15.57$

or, $K_C = 3.7\times10^{15} \approx 4\times10^{15}$

35. (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 convertor, 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.

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

36. (a) : From the given expression:
At anode : $A \rightarrow A^+ + e$
At cathode : $B^+ + e \rightarrow B$
Overall reaction is : $A + B^+ \rightarrow A^+ + B$

37. (d) : $Fe^{2+} + 2e^- \rightarrow Fe$; $E° = -0.441$ V ... (i)
$Fe^{3+} + e^- \rightarrow Fe^{2+}$; $E° = 0.771$ V ... (ii)
$Fe + 2Fe^{3+} \rightarrow 3Fe^{2+}$; $E° = ?$
To get the above equation, (ii) × 2 − (i)

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

38. (d) : We know that,
1 Faraday charge liberates 1 eq. of substance.
This is the Faraday Law.

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

$$\text{no. of eq. of Al} = \frac{\text{wt. of Al}}{\text{eq. wt.}} = \frac{4.5}{9} = 0.5$$

no. of Faraday required = 0.5
⇒ no. of eq. of H_2 produced = 0.5 eq.

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

⇒ Volume occupied by 0.5 = 11.2 × 0.5
= 5.6 L at STP

39. (c) : $\underset{\text{(from bauxite)}}{3C + 2Al_2O_3} \longrightarrow 4Al + 3CO_2$
4 moles of Al is produced by 3 moles of C

1 mole of Al is produced by $\frac{3}{4}$ mole of C

10^4 moles of Al is produced by $\frac{3}{4}\times10^4$ moles of C

Amount of carbon used = $\frac{3}{4} \times 10^4 \times 12$ g

$= \frac{3}{4} \times 10 \times 12$ kg = 90 kg

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

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

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

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

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

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

41. (a) : Cell reaction can be represented as $Zn + Cu^{2+} \rightarrow Cu + Zn^{2+}$

Applying in both cases,

$$E^\circ = \frac{-0.0591}{2}\log\frac{Zn^{2+}}{Cu^{2+}}$$

42. (a) : For O_2, $\Delta G = -nFE^\circ$

$$E^\circ = \frac{\Delta G}{-nF} = \frac{-827000}{-2\times 96500} = 4.28$$

Minimum EMF required to carry out electrolysis of $Al_2O_3 = \frac{4.28}{2} = 2.14$ V

43. (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^+ ion 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 $H^+ + e \rightarrow H$, $H + H \rightarrow H_2$

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

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

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 perference to H^+ ions, yielding sodium, which dissolves in mercury to form sodium amalgam. At cathode : $Na^+ + e = Na$

44. (b) : The metals having higher negative values of their electrode potential can displace metals having lower values from their salt solutions.

45. (a) : $\lambda_\infty = \frac{1}{n^+}\lambda_+^\infty + \frac{1}{n^-}\lambda_-^\infty$

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

$= \frac{1}{2}\times 127 + 76 = 63\cdot5 + 76 = 139\cdot5$

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

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

$Cu^{2+} + e \rightarrow Cu^+;\ E_2^\circ = 0\cdot15$ V ...(2)

$Cu^+ + e \rightarrow Cu;\ E^\circ_3 = ?$ ...(3)

Now $\Delta G^\circ_1 - nFE^\circ_1 = -2\times 0\cdot34\times F$

$$\Delta G^\circ_2 = -1\times 0\cdot15\times F,\ \Delta G^\circ_3 = -1\times E^\circ_3\times F$$

Again $\Delta G_1^\circ = \Delta G_2^\circ + \Delta G_3^\circ \Rightarrow -0\cdot68F$

$= -0.15\,F - E_3^\circ \times F$

$\Rightarrow E_3^\circ = 0\cdot68 - 0\cdot15 = 0\cdot53$ V

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

$= 0\cdot53 - 0\cdot15 = 0\cdot38$ V

47. (b) : $K = 0.012$ ohm^{-1}cm^{-1}

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

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

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

48. (b) : We know $\Delta G = nFE^\circ$

Again according to Nernst equation,

$$E^\circ = \frac{0.059}{n}\log\frac{[Zn^{2+}]}{[Cu^{2+}]},\ \text{at } 25°C$$

So ΔG (free energy change) in an electrochemical cell at a given temperature is a function of

$$\ln\frac{[Zn^{2+}]}{[Cu^{2+}]} = \ln\left(\frac{C_2}{C_1}\right)$$

49. (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)

$\Rightarrow 1.1 = \frac{0.059}{2}\log K \Rightarrow K = 10^{-37}$

50. (d) : Molar conductance $\left(\Lambda_m^{\infty}\right)$ NaCl

$= 126.45\ \text{ohm}^{-1}\text{cm}^2\text{mol}^{-1}$;

HCl $= 426.16\ \text{ohm}^{-1}\text{cm}^2\text{mol}^{-1}$ and

$CH_3COONa = 91\ \text{ohm}^{-1}\text{cm}^2\text{mol}^{-1}$.

We know that molar conductance of CH_3COOH

$$= \Lambda_m^{\infty}(CH_3COONa) + \Lambda_m^{\infty}(HCl) - \Lambda_m^{\infty}(NaCl)$$

$$= 91 + 426.16 - 126.45 = 390.71\ \text{ohm}^{-1}\text{cm}^2\text{mol}^{-1}$$

51. (d) : Current (I) = 5 ampere and time (t) = 40 minutes = 2400 seconds.

Amount of electricity passed

$(Q) = I \times t = 5 \times 2400 = 12000$ C

Now $Zn^{2+} + 2e^- \rightarrow Zn$ (1 mole = 65.39)

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

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

52. (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.M.F. $= +0.76 - 0.44 = +0.32$ V

53. (d) : Since it is a concentration cell and the concentration of ions in two electrolyte solutions (HCl and CH_3COOH) are different, therefore e.m.f. of this cell will not be zero.

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

55. (b) : A cation with maximum value of standard reduction potential, has the strongest oxidising power.

56. (b) : Product obtained at anode will be oxygen.

At anode : $2OH^- \rightleftharpoons H_2O + 1/2O_2$

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