ELECTROCHEMISTRY

FACT/DEFINITION TYPE QUESTIONS

1. Which of the following statements is incorrect regarding electrochemistry?
- (a) It is the study of production of electricity from energy released during spontaneous chemical reactions.
- (b) NaOH, Cl_2, alkali and alkaline earth metals are prepared by electrochemical methods.
- (c) The demerit associated with electrochemical methods is that they are more polluting. Thus they are eco-destructive.
- (d) Electrochemical reactions are more energy efficient and less polluting.

2. What flows in the internal circuit of a galvanic cell?
- (a) Ions (b) Electrons
- (c) Electricity (d) Atoms

3. Which of the following statements about galvanic cell is incorrect
- (a) anode is positive
- (b) oxidation occurs at the electrode with lower reduction potential
- (c) cathode is positive
- (d) reduction occurs at cathode

4. Reaction that takes place at graphite anode in dry cell is
- (a) $Zn^{2+} + 2e^- \longrightarrow Zn(s)$
- (b) $Zn(s) \longrightarrow Zn^{2+} + 2e^-$
- (c) $Mn^{2+} + 2e^- \longrightarrow Mn(s)$
- (d) $Mn(s) \longrightarrow Mn^+ + e^- + 1.5V$

5. In which of the following conditions salt bridge is not required in a galvanic cell?
- (a) When galvanic cell is used in geyser.
- (b) When distance between oxidation half cell and reduction half cell is negligible.
- (c) Electrolytic solutions used in both the half cells are of same concentration.
- (d) When both the electrodes are dipped in the same electrolytic solution.

6. Which device converts chemical energy of a spontaneous redox reaction into electrical energy?
- (a) Galvanic cell (b) Electrolytic cell
- (c) Daniell cell (d) Both (a) and (c)

7. Which one is not called a anode reaction from the following?
- (a) $Cl^- \to \frac{1}{2}Cl_2 + e^-$ (b) $Cu \to Cu^{++} + 2e^-$
- (c) $Hg^+ \to Hg^{++} + e^-$ (d) $Zn^{2+} + 2e^- \to Zn$

8. The cell reaction $Cu + 2Ag^+ \to Cu^{+2} + Ag$ is best represented by
- (a) $Cu(s)\,|\,Cu^{+2}(aq)\,||\,Ag^+(aq)\,|\,Ag(s)$
- (b) $Pt\,|\,Cu^{+2}\,||\,Ag^+(aq)\,|\,Ag(s)$
- (c) $Cu^{+2}\,|\,Cu\,||\,Pt\,|\,Ag$
- (d) None of the above representations

9. $\underset{\text{(anode)}}{Zn(s)\,|\,Zn^{2+}(aq)}\,||\,\underset{\text{(cathode)}}{Cu^{2+}(aq)\,|\,Cu(s)}$ is
- (a) Weston cell (b) Daniel cell
- (c) Calomel cell (d) Faraday cell

10. The tendency of an electrode to lose electrons is known as
- (a) electrode potential (b) reduction potential
- (c) oxidation potential (d) e.m.f.

11. The chemical reaction,

$$2AgCl(s) + H_2(g) \longrightarrow 2HCl(aq) + 2Ag(s)$$

taking place in a galvanic cell is represented by the notation
- (a) $Pt(s)\,|\,H_2(g), 1\text{ bar}\,|\,1M\,KCl(aq)\,|\,AgCl(s)\,|\,Ag(s)$
- (b) $Pt(s)\,|\,H_2(g), 1\text{ bar}\,|\,1M\,HCl(aq)\,|\,1M\,Ag^+(aq)\,|\,Ag(s)$
- (c) $Pt(s)\,|\,H_2(g), 1\text{ bar}\,|\,1M\,HCl(aq)\,|\,AgCl(s)\,|\,Ag(s)$
- (d) $Pt(s)\,|\,H_2(g), 1\text{ bar}\,|\,1M\,HCl(aq)\,|\,Ag(s)\,|\,AgCl(s)$

12. Given that the standard reduction potentials for M^+/M and N^+/N electrodes at 298 K are 0.52 V and 0.25 V respectively. Which of the following is correct in respect of the following electrochemical cell ?

$$M/M^+\,||\,N^+/N$$

- (a) The overall cell reaction is a spontaneous reaction.
- (b) The standard EMF of the cell is – 0.27 V.
- (c) The standard EMF of the cell is 0.77 V.
- (d) The standard EMF of the cell is – 0.77 V.

13. The difference between the electrode potentials of two electrodes when no current is drawn through the cell is called _________.

(a) Cell potentials (b) Cell emf
(c) Potential difference (d) Cell voltage

14. For the given Nernst equation

$$E_{cell} = E^\circ_{cell} - \frac{RT}{2F}\ln\frac{[Mg^{2+}]}{[Ag^+]^2}$$

Which of the following representation is correct?

(a) $Ag^+|Ag||Mg^{2+}|Mg$ (b) $Mg^{2+}|Mg||Ag|Ag^+$
(c) $Mg|Mg^{2+}||Ag^+|Ag$ (d) $Mg|Mg^{2+}||Ag|Ag^+$

15. For cell representation:

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

Which of the following is correct?

(i) Cu is reducing agent.
(ii) Overall cell reaction is

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

(iii) Cu is cathode
(iv) Ag is anode

(a) (ii), (iii) and (iv) (b) (ii), (iii) and (iv)
(c) (iii) and (iv) (d) (i) and (ii)

16. The reference electrode is made by using

(a) $ZnCl_2$ (b) $CuSO_4$
(c) $HgCl_2$ (d) Hg_2Cl_2

17. The standard hydrogen electrode potential is zero, because

(a) hydrogen oxidized easily
(b) electrode potential is considered as zero
(c) hydrogen atom has only one electron
(d) hydrogen is a very light element

18. Without losing its concentration $ZnCl_2$ solution cannot be kept in contact with

(a) Au (b) Al
(c) Pb (d) Ag

19. On the basis of the following E° values, the strongest oxidizing agent is :

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

(a) $[Fe(CN)_6]^{4-}$ (b) Fe^{2+}
(c) Fe^{3+} (d) $[Fe(CN)_6]^{3-}$

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) X > Y > Z
(c) Z > X > Y (d) X > Y > Z

21. 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

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

$F_2(g) + 2e^- \rightarrow 2F^-(aq)$; $E^\circ = +2.85$ V
$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$; $E^\circ = +1.36$ V
$Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$; $E^\circ = +1.06$ V
$I_2(s) + 2e^- \rightarrow 2I^-(aq)$; $E^\circ = +0.53$ 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

23. A button cell used in watches functions 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) 0.42 V (b) 0.84 V
(c) 1.34 V (d) 1.10 V

24. The oxidation potentials of A and B are +2.37 and +1.66 V respectively. In chemical reactions

(a) A will be replaced by B
(b) A will replace B
(c) A will not replace B
(d) A and B will not replace each other

25. A smuggler could not carry gold by depositing iron on the gold surface since

(a) gold is denser
(b) iron rusts
(c) gold has higher reduction potential than iron
(d) gold has lower reduction potential than iron

26. Which cell will measure standard electrode potential of copper electrode ?

(a) $Pt(s)|H_2(g, 0.1\ bar)|H^+(aq., 1\ M)||Cu^{2+}(aq., 1\ M)|Cu$
(b) $Pt(s)|H_2(g, 1\ bar)|H^+(aq., 1\ M)||Cu^{2+}(aq., 2\ M)|Cu$
(c) $Pt(s)|H_2(g, 1\ bar)|H^+(aq., 1\ M)||Cu^{2+}(aq., 1\ M)|Cu$
(d) $Pt(s)|H_2(g, 1\ bar)|H^+(aq., 0.1\ M)||Cu^{2+}(aq., 1\ M)|Cu$

27. Which of the following statement is not correct about an inert electrode in a cell ?

(a) It does not participate in the cell reaction.
(b) It provides surface either for oxidation or for reduction reaction.
(c) It provides surface for conduction of electrons.
(d) It provides surface for redox reaction.

28. In the electrochemical reaction

$$2Fe^{3+} + Zn \longrightarrow Zn^{2+} + 2Fe^{2+},$$

on increasing the concentration of Fe^{2+}

(a) increases cell emf
(b) increases the current flow
(c) decreases the cell emf
(d) alters the pH of the solution

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
(Given $F = 96500\ C\ mol^{-1}; R = 8.314 JK^{-1} mol^{-1}$)

(a) 2.0×10^{11} (b) 4.0×10^{12}
(c) 1.0×10^{2} (d) 1.0×10^{10}

30. What will be the emf for the given cell
$Pt | H_2(P_1) | H^+(aq) || H_2(P_2) | Pt$

(a) $\frac{RT}{F} \log_e \frac{P_1}{P_2}$ (b) $\frac{RT}{2F} \log_e \frac{P_1}{P_2}$
(c) $\frac{RT}{F} \log_e \frac{P_2}{P_1}$ (d) None of these

31. The value of electrode potential $(10^{-4}\ M)\ H^+ | H_2(1\ atm) | Pt$ at 298 K would be

(a) −0.236 V (b) +0.404 V
(c) +0.236 V (d) −0.476 V

32. According to Nernst equation, which is not correct if $Q = K_c$:

(a) $E_{cell} = 0$ (b) $\frac{RT}{nF} \ln Q = E^\circ_{cell}$
(c) $K_c = e^{\frac{nFE^\circ_{cell}}{RT}}$ (d) $E_{cell} = E^\circ_{cell}$

33. The standard emf of a cell, involving one electron change is found to be 0.591 V at 25°C. The equilibrium constant of the reaction is $(F = 96500\ C\ mol^{-1})$

(a) 1.0×10^{1} (b) 1.0×10^{5}
(c) 1.0×10^{10} (d) 1.0×10^{30}

34. For the galvanic cell
$Zn | Zn^{2+}(0.1M) || Cu^{2+}(1.0M) | Cu$ the cell potential increase if:

(a) $[Zn^{2+}]$ is increased
(b) $[Cu^{2+}]$ is increased
(c) $[Cu^{2+}]$ is decreased
(d) surface area of anode is increased

35. Consider the following cell reaction:
$2Fe(s) + O_2(g) + 4H^+(aq) \rightarrow 2Fe^{2+}(aq) + 2H_2O(l); E^\circ = 1.67 V$
At $[Fe^{2+}] = 10^{-3}$ M, $p(O_2) = 0.1$ atm and pH = 3, the cell potential at 25°C is

(a) 1.47 V (b) 1.77 V
(c) 1.87 V (d) 1.57 V

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

$$Zn \left| \begin{matrix} ZnSO_4 \\ (0.01\ M) \end{matrix} \right|\left| \begin{matrix} CuSO_4 \\ (1.0\ M) \end{matrix} \right| Cu$$

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

(a) $E_2 = 0 \neq E_1$ (b) $E_1 > E_2$
(c) $E_1 < E_2$ (d) $E_1 = E_2$

37. The electrode potential $E_{(Zn^{2+}/Zn)}$ of a zinc electrode at 25°C with an aqueous solution of 0.1 M $ZnSO_4$ is
$[E^\circ_{(Zn^{2+}/Zn)} = -0.76$ V. Assume $\frac{2.303RT}{F} = 0.06$ at 298 K].

(a) +0.73 (b) −0.79
(c) −0.82 (d) −0.70

38. In the cell reaction
$Cu(s) + 2Ag^+(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$,
$E^o_{cell} = 0.46$ V. By doubling the concentration of Cu^{2+}, E^o_{cell} will become

(a) doubled
(b) halved
(c) increases but less than double
(d) decreases by a small fraction

39. E° of a cell $aA + bB \longrightarrow cC + dD$ is

(a) $E + RT \ln \frac{[a]^A [b]^B}{[c]^C [d]^D}$ (b) $E + \frac{RT}{nF} \ln \frac{[C]^c [D]^d}{[A]^a [B]^b}$
(c) $E + \frac{RT}{nF} \ln \frac{[C]^c [d]^D}{[A]^A [B]^B}$ (d) $E + \frac{RT}{nF} \ln \frac{[a]^A [B]^B}{[C]^C [d]^D}$

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

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

41. What is the standard cell potential E° for an electrochemical cell in which the following reaction takes place spontaneously ?
$Cl_2(g) + 2Br^- \rightarrow Br_2(aq) + 2Cl^-\ \Delta G^\circ = -50.6$ kJ

(a) 1.2 V (b) 0.53 V
(c) 0.26 V (d) –0.53 V

42. If 0.01 M solution of an electrolyte has a resistance of 40 ohms in a cell having a cell constant of 0.4 cm^{-1}, then its molar conductance in $ohm^{-1}\ cm^2\ mol^{-1}$ is

(a) 10^2 (b) 10^4
(c) 10 (d) 10^3

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

(a) 0.0616 cm^{-1} (b) 0.66 cm^{-1}
(c) 6.60 cm^{-1} (d) 660 cm^{-1}

44. The unit of equivalent conductivity is

(a) ohm cm
(b) $ohm^{-1}\ cm^2$ (g equivalent)$^{-1}$
(c) ohm cm^2 (g equivalent)
(d) S cm^{-2}

45. The resistance of 0.01 N solution of an electrolyte was found to be 220 ohm at 298 K using a conductivity cell with a cell constant of $0.88 cm^{-1}$. The value of equivalent conductance of solution is –

(a) 400 mho cm^2 g eq^{-1} (b) 295 mho cm^2 g eq^{-1}
(c) 419 mho cm^2 g eq^{-1} (d) 425 mho cm^2 g eq^{-1}

46. Specific conductance of 0.1 M HNO_3 is 6.3×10^{-2} ohm^{-1} cm^{-1}. The molar conductance of the solution is

(a) 100 ohm^{-1} cm^2 (b) 515 ohm^{-1} cm^2
(c) 630 ohm^{-1} cm^2 (d) 6300 ohm^{-1} cm^2

47. 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) 0.142 cm^{-1} (b) 0.66 cm^{-1}
(c) 0.918 cm^{-1} (d) 1.12 cm^{-1}

48. The unit of specific conductivity is

(a) ohm cm^{-1} (b) ohm cm^{-2}
(c) ohm^{-1} cm (d) ohm^{-1} cm^{-1}

49. Which of the following solutions of KCl will have the highest value of specific conductance?

(a) 1.0 N (b) 0.1 N
(c) 1.0×10^{-2}N (d) 1.0×10^{-3}N

50. The cell constant of a conductivity cell ___________.

(a) changes with change of electrolyte.
(b) changes with change of concentration of electrolyte.
(c) changes with temperature of electrolyte.
(d) remains constant for a cell.

51. Which of the following pair(s) is/are incorrectly matched?

(i) R (resistance) – ohm (Ω)
(ii) ρ (resistivity) – ohm metre (Ωm)
(iii) G (conductance) – seimens or ohm (S)
(iv) κ (conductivity) – seimens metre^{-1} (Sm^{-1})

(a) (i), (ii) and (iii) (b) (ii) and (iii)
(c) (i), (ii) and (iv) (d) (iii) only

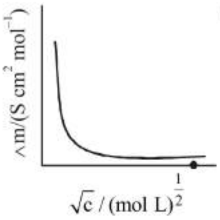
52. On which of the following magnitude of conductivity does not depends?

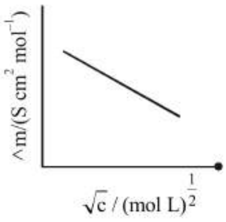
(a) Nature of material (b) Temperature
(c) Pressure (d) Mass of the material

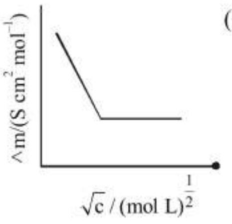
53. Which of the following expression correctly represents molar conductivity?

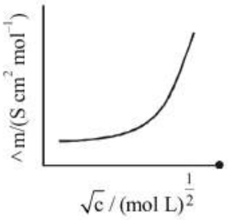
(a) $\wedge_m = \frac{K}{C}$ (b) $\wedge_m = \frac{KA}{1}$
(c) $\wedge_m = KV$ (d) All of these

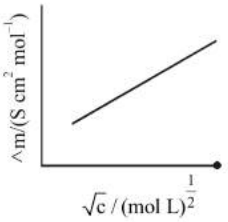
54. Which of the following represents variation of molar conductance of electrolyte with $(\text{concentration})^{1/2}$ respectively for weak and strong electrolyte ?

(i)


(ii)


(iii)


(iv)


(v)


	Weak acid	Strong acid
(a)	(iv)	(v)
(b)	(ii)	(iv)
(c)	(i)	(ii)
(d)	(iii)	(ii)

55.

Electrolyte:	KCl	KNO_3	HCl	NaOAc	NaCl
Λ^{∞} ($S\ cm^2\ mol^{-1}$) :	149.9	145	426.2	91	126.5

Calculate Λ^{∞}_{HOAc} using appropriate molar conductances of the electrolytes listed above at infinite dilution in H_2O at 25°C

(a) 217.5 (b) 390.7
(c) 552.7 (d) 517.2

56. Kohlrausch's law states that at

(a) finite dilution, each ion makes definite contribution to equivalent 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 depending on the nature of the other ion of the electrolyte.
(c) infinite dilution, each ion makes definite contribution to 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, whatever be the nature of the other ion of the electrolyte.

57. Which of the following expressions correctly represents the equivalent conductance at infinite dilution of $Al_2(SO_4)_3$, Given that $\Lambda^{\circ}_{Al^{3+}}$ and $\Lambda^{\circ}_{SO_4^{2-}}$ are the equivalent conductances at infinite dilution of the respective ions?

(a) $\frac{1}{3}\Lambda^{\circ}_{Al^{3+}} + \frac{1}{2}\Lambda^{\circ}_{SO_4^{2-}}$ (b) $2\Lambda^{\circ}_{Al^{3+}} + 3\Lambda^{\circ}_{SO_4^{2-}}$
(c) $\Lambda^{\circ}_{Al^{3+}} + \Lambda^{\circ}_{SO_4^{2-}}$ (d) $\left(\Lambda^{\circ}_{Al^{3+}} + \Lambda^{\circ}_{SO_4^{2-}}\right)\times 6$

58. Limiting molar conductivity of NH_4OH $\left(\text{i.e., } \Lambda^{\circ}_{m(NH_4OH)}\right)$ 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) $\Lambda^{\circ}_{m(NH_4Cl)} + \Lambda^{\circ}_{m(NaOH)} - \Lambda^{\circ}_{m(NaCl)}$

59. 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}

60. At 25°C, the molar conductance at infinite dilution for the strong electrolytes NaOH, NaCl and $BaCl_2$ are 248×10^{-4}, 126×10^{-4} and 280×10^{-4} Sm^2mol^{-1} respectively. Λ^0_m $Ba(OH)_2$ in S m^2 mol^{-1} is

(a) 52.4×10^{-4} (b) 524×10^{-4}

(c) 402×10^{-4} (d) 262×10^{-4}

61. $\Lambda_{ClCH_2COONa} = 224\ ohm^{-1}cm^2gm\,eq^{-1}$,

$\Lambda_{NaCl} = 38.2\ ohm^{-1}cm^2gm\,eq^{-1}$,

$\Lambda_{HCl} = 203\ ohm^{-1}cm^2gm\,eq^{-1}$,

What is the value of Λ_{ClCH_2COOH}

(a) 288.5 $ohm^{-1}cm^2gmeq^{-1}$

(b) 289.5 $ohm^{-1}cm^2gmeq^{-1}$

(c) 388.8 $ohm^{-1}cm^2gmeq^{-1}$

(d) 59.5 $ohm^{-1}cm^2gmeq^{-1}$

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

(a) 20.800% (b) 4.008%

(c) 40.800% (d) 2.080%

63. The electrical properties and their respective SI units are given below. Identify the wrongly matched pair.

	Electrical property	**SI unit**
(a)	Specific conductance	S m^{-1}
(b)	Conductance	S
(c)	Equivalent conductance	S m^2 $gequiv^{-1}$
(d)	Cell constant	m

64. The ion of least limiting molar conductivity among the following is

(a) SO_4^{2-} (b) H^+

(c) Ca^{2+} (d) CH_3COO^-

65. Molar ionic conductivities of a two-bivalent electrolytes x^{2+} and y^{2-} are 57 and 73 respectively. The molar conductivity of the solution formed by them will be

(a) 130 S cm^2 mol^{-1} (b) 65 S cm^2 mol^{-1}

(c) 260 S cm^2 mol^{-1} (d) 187 S cm^2 mol^{-1}

66. 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

67. If 0.5 amp current is passed through acidified silver nitrate solution for 100 minutes. The mass of silver deposited on cathode, is (eq.wt.of silver nitrate = 108)

(a) 2.3523 g (b) 3.3575 g

(c) 5.3578 g (d) 6.3575 g

68. Aluminium oxide may be electrolysed at 1000°C to furnish aluminium metal (At. Mass = 27 amu; 1 Faraday = 96,500 Coulombs). The cathode reaction is– $Al^{3+} + 3e^- \rightarrow Al$

To prepare 5.12 kg of aluminium metal by this method we require electricity of

(a) 5.49×10^1 C (b) 5.49×10^4 C

(c) 1.83×10^7 C (d) 5.49×10^7 C

69. Which of the following is the use of electrolysis?

(a) Electrorefining (b) Electroplating

(c) Both (a) & (b) (d) None of these

70. An electrolytic cell contains a solution of Ag_2SO_4 and has platinum electrodes. A current is passed until 1.6 gm of O_2 has been liberated at anode. The amount of silver deposited at cathode would be

(a) 107.88 gm (b) 1.6 gm

(c) 0.8 gm (d) 21.60 gm

71. When 9650 coulombs of electricity is passed through a solution of copper sulphate, the amount of copper deposited is (given at. wt. of Cu = 63.6)

(a) 0318g (b) 3.18 g

(c) 31.8g (d) 63.6g

72. Find the charge in coulombs required to convert 0.2 mole VO_3^{-2} into VO_4^{-3} –

(a) 1.93×10^4 (b) 9.65×10^4

(c) 1.93×10^5 (d) 9.65×10^5

73. A silver cup is plated with silver by passing 965 coulombs of electricity. The amount of Ag deposited is :

(a) 107.89 g (b) 9.89 g

(c) 1.0002 g (d) 1.08 g

74. The number of coulombs required to reduce 12.3 g of nitrobenzene to aniline is :

(a) 115800 C (b) 5790 C

(c) 28950 C (d) 57900 C

75. The amount of electricity that can deposit 108 g of Ag from $AgNO_3$ solution is:

(a) 1 F (b) 2 A

(c) 1 C (d) 1 A

76. To deposit one equivalent weight of silver at cathode, the charge required will be

(a) 9.65×10^4 C (b) 9.65×10^3 C
(c) 9.65×10^5 C (d) 9.65×10^7 C

77. The volume of oxygen gas liberated at NTP by passing a current of 9650 coulombs through acidified water is

(a) 1.12 litre (b) 2.24 litre
(c) 11.2 litre (d) 22.4 litre

78. Three faradays electricity was passed through an aqueous solution of iron (II) bromide. The weight of iron metal (at. wt = 65) deposited at the cathode (in gm) is

(a) 56 (b) 84
(c) 112 (d) 168

79. On passing C ampere of electricity through a electrolyte solution for t second. m gram metal deposits on cathode. The equivalent weight E of the metal is

(a) $E = \dfrac{C \times t}{m \times 96500}$ (b) $E = \dfrac{C \times m}{t \times 96500}$

(c) $E = \dfrac{96500 \times m}{C \times t}$ (d) $E = \dfrac{C \times t \times 96500}{m}$

80. The number of electrons passing per second through a cross-section of copper wire carrying 10^{-6} amperes of current per second is found to be

(a) 1.6×10^{-19} (b) 6×10^{-35}
(c) 6×10^{-16} (d) 6×10^{12}

81. Faraday's laws of electrolysis will fail when

(a) temperature is increased
(b) inert electrodes are used
(c) a mixture of electrolytes is used
(d) None of these cases

82. The electric charge for electrode decomposition of one gram equivalent of a substance is

(a) one ampere per second
(b) 96500 coulombs per second
(c) one ampere for one hour
(d) charge on one mole of electrons

83. In electrolysis of dilute H_2SO_4 using platinum electrodes

(a) H_2 is evolved at cathode
(b) NH_2 is produced at anode
(c) Cl_2 is obtained at cathode
(d) O_2 is produced

84. In electrolysis of NaCl when Pt electrode is taken then H_2 is liberated at cathode while with Hg cathode it forms sodium amalgam. This is because

(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

85. Electrolysis of fused NaCl will give

(a) Na (b) NaOH
(c) NaClO (d) None of these

86. How many moles of Pt may be deposited on the cathode when 0.80 F of electricity is passed through a 1.0 M solution of Pt^{4+}?

(a) 1.0 mol (b) 0.20 mol
(c) 0.40 mol (d) 0.80 mol

87. A current strength of 9.65 amperes is passed through excess fused $AlCl_3$ for 5 hours. How many litres of chlorine will be liberated at STP? (F = 96500 C)

(a) 2.016 (b) 1.008
(c) 11.2 (d) 20.16

88. A solution of copper sulphate ($CuSO_4$) is electrolysed for 10 minutes with a current of 1.5 amperes. The mass of copper deposited at the cathode (at. mass of Cu = 63u) is

(a) 0.3892g (b) 0.2938g
(c) 0.2398g (d) 0.3928g

89. 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

90. 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) 54.9 g (d) 108.0 g

91. Electrolysis of a salt solution was carried out, after some time solution turned yellow than salt can be

(i) NaCl (ii) KCl
(iii) RbCl (iv) KBr

(a) (i), (ii) and (iii) (b) (ii), (ii) and (iv)
(c) (i), (ii) and (iv) (d) (i), (iii) and (iv)

92. Which of the following statements is incorrect?

(a) Both electronic and electrolytic conductance depends on the nature of conducting material.
(b) Both electronic and electrolytic conductance varies similarly with temperature.
(c) Electronic conductance is independent but electrolytic conductance depends on the amount of the conducting substance.
(d) All the above statements are incorrect.

93. Which of the following statements is incorrect?

(a) Electrodes made up of gold participates in the chemical reaction.
(b) Electrolytic products of NaCl are Na and Cl_2 whereas of aqueous NaCl are NaOH, Cl_2 and H_2.
(c) During electrolysis at cathode, reaction with higher value of $E^{\oplus}$ is preferred.
(d) All of the above statements are incorrect.

94. During electrolysis of sulphuric acid, which of the following processes is possible at anode?

A. $2H_2O(l) \longrightarrow O_2(g) + 4H^+(aq) + 4e^-$ $E^{\ominus}_{\text{cell}} = +1.23\ V$

B. $2SO_4^{2-}(aq) \longrightarrow S_2O_8^{2-}(aq) + 2e^-$ $E^{\ominus}_{\text{cell}} = 1.96\ V$

Choose the correct option based on following statements.

(i) Process A is preferred at higher concentration of sulphuric acid.

(ii) Process B is preferred at higher concentration of sulphuric acid.

(iii) Process A is preferred for dilute sulphuric acid.

(iv) Process B is preferred for dilute sulphuric acid.

(v) Both A and B are equally possible at higher concentration.

(a) (v) and (iii) (b) (iii) and (ii)

(c) (i) and (iv) (d) (v) and (iv)

95. Which of the following metals is not produced by electrochemical reduction?

(a) Na (b) Fe

(c) Mg (d) Al

96. As lead storage battery is charged

(a) lead dioxide dissolves

(b) sulphuric acid is regenerated

(c) lead electrode becomes coated with lead sulphate

(d) the concentration of sulphuric acid decreases

97. During the charging of lead storage battery, the reaction at anode is represented by

(a) $Pb^{2+} + SO_4^{2-} \longrightarrow PbSO_4$

(b) $PbSO_4 + 2H_2O \longrightarrow PbO_2 + SO_4^{2-} + 4H^+ + 2e^-$

(c) $Pb \longrightarrow Pb^{2+} + 2e^-$

(d) $Pb^{2+} + 2e^- \longrightarrow Pb$

98. Which colourless gas evolves, when NH_4Cl reacts with zinc in a dry cell battery

(a) NH_4 (b) N_2

(c) H_2 (d) Cl_2

99. In a hydrogen-oxygen fuel cell, combustion of hydrogen occurs to

(a) produce high purity water

(b) create potential difference between two electrodes

(c) generate heat

(d) remove adsorbed oxygen from elctrode surfaces

100. Among the following cells:

Leclanche cell (i)

Nickel-Cadmium cell (ii)

Lead storage battery (iii)

Mercury cell (iv)

primary cells are

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

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

101. The electrolyte used in Leclanche cell is

(a) paste of KOH and ZnO

(b) 38% solution of H_2SO_4

(c) moist paste of NH_4Cl and $ZnCl_2$

(d) moist sodium hydroxide

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

(a) Electrolytic cell (b) Dynamo

(c) Ni-Cd cell (d) Fuel Cell

103. Which one of the following cells can convert chemical energy of H_2 and O_2 directly into electrical energy?

(a) Mercury cell (b) Daniell cell

(c) Fuel cell (d) Lead storage cell

104. Hydrogen-Oxygen fuel cells are used in space craft to supply

(a) power for heat and light

(b) power for pressure

(c) oxygen

(d) water

105. Prevention of corrosion of iron by zinc coating is called

(a) electrolysis (b) photoelectrolysis

(c) cathodic protection (d) galvanization

106. The best way to prevent rusting of iron is

(a) making it cathode (b) putting in saline water

(c) Both of these (d) None of these

107. Several blocks of magnesium are fixed to the bottom of a ship to

(a) make the ship lighter

(b) prevent action of water and salt

(c) prevent puncturing by under-sea rocks

(d) keep away the sharks

108. Which of the following batteries cannot be reused?

(a) Lead storage battery (b) Ni-Cd cell

(c) Mercury cell (d) Both (b) and (c)

109. Which of the following is a merit of Ni–Cd cell over lead storage battery?

(a) Ni–Cd cell can be re-used.

(b) Ni–Cd cell is comparatively economical to manufacture

(c) Ni–Cd cell has comparatively longer life

(d) All the above are the merits of Ni–Cd cell over lead storage battery.

110. Which of the following statements regarding fuel cell is incorrect?

(a) These cells are eco-friendly.

(b) These cells convert energy of combustion of fuels like H_2, CH_4, CH_3OH etc., directly into electrical energy.

(c) $H_2 - O_2$ fuel cell is used in Apollo space programme.

(d) Fuel cells produce electricity with an efficiency of about 100%.

STATEMENT TYPE QUESTIONS

111.

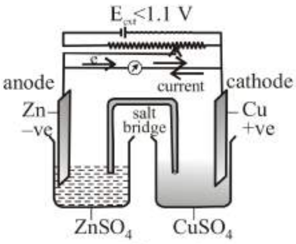


Figure -1

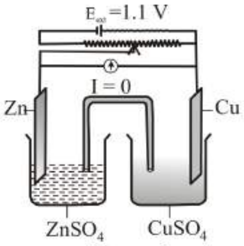


Figure -2

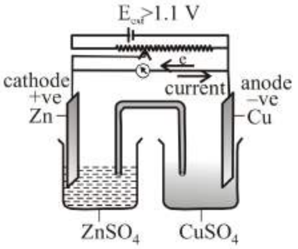


Figure -3

(i) Figure 1 represents electrochemical and Figure 3 represents electrolytic cell.

(ii) Figure 2 represents electrolytic and Figure 3 represents electrochemical cell.

(iii) Figure 2 represents a cell which is not working i.e. no current flows through the cell.

(iv) Energy conversion shown in Figure 1 is chemical to electrical whereas energy conversion shown in Figure 2 is electrical to chemical.

Which of the following is the correct coding for the statements above.

(a) TFTT (b) TTTT

(c) TFFT (d) FTFF

112. Which of the following statements regarding given cell representation is/are correct?

$Cd(s)/Cd^{2+}(aq)//Ag^{+}(aq)/Ag(s)$

(i) In the given cell Cd electrode act as an anode whereas Ag electrode acts as a cathode.

(ii) In the given cell Cd electrode acts as a cathode whereas Ag electrode acts as a annode.

(iii) $E_{cell} = E_{Ag^{+}/Ag} - E_{Cd^{2+}/Cd}$

(a) (i) and (ii) (b) Only (ii)

(c) Only (i) (d) (i) and (iii)

113. Which of the following is/are correct statement(s) for the addition of Li, K, Rb to the aqueous solution of Na^{+}.

(i) The correct order of metals in which they reduce the Na^{+} ion is $Rb < K < Li$.

(ii) Reduction of metal ions would not take place.

(a) Statement (i) and (ii) are correct.

(b) Statement (i) is correct only.

(c) Statement (ii) is correct only.

(d) Neither (i) nor (ii) is correct.

114. Read the following statements carefully.

(i) According to a convention cell potential of hydrogen electrode (S.H.E.) is considered to be zero at all temperatures.

(ii) e.m.f. of the cell $Pt(s)/H_2(g,\ 1\ bar)/H^{+}(aq,\ 1\ M) \,||\, Zn^{2+}(aq, 1M)/Zn$ is – 0.76. This negative value indicates that Zn^{2+} ion reduces less easily then H^{+} ions.

(iii) Copper does not dissolve in HCl but dissolves in HNO_3 as in nitric acid it gets oxidised by nitrate ion.

(iv) Inert metals like Pt or Au are used in certain electrodes i.e., these metals does not participate in reaction but provide surface for oxidation and reduction reactions.

(v) Fluorine has the highest electrode potential thereby making it strongest oxidising agent whereas lithium with lowest electrode potential is the weakest oxidising and strongest reducing agent.

Which of the following is the correct coding for the statements above.

(a) TTTTT (b) TTFTF

(c) FFTTT (d) FFFTT

115. Which of the following statement(s) is/are correct?

(i) Molar conductivity for strong electrolytes increases gradually and of weak electrolytes increases rapidly on dilution.

(ii) If α is the degree of dissociation of weak electrolytes.

Then, $\alpha = \frac{\Lambda_m}{\Lambda^{\circ}_m}$

(iii) Molar conductivity of CaX_2 increases rapidly on dilution.

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

(c) (i) only (d) (i) and (ii)

116. Read the following statements.

(i) According to Faraday's second law amounts of different substances liberated by same quantity of electricity passing through the electrolytic solution are proportional to their chemical equivalent weights.

(ii) $1\ F = 96487\ Cmol^{-1} \simeq 96500\ Cmol^{-1}$ (for more accurate calculation).

(iii) As per electrode reactions

$$K^{+} + e^{-} \longrightarrow K$$

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

one mole of K^{+} and Al^{3+} require 1(1F) and 3(3F) mol of electrons respectively.

Which of the following is the correct coding for the above statements?

(a) TTT (b) FFT

(c) TFT (d) FTF

117. Which of the following statement(s) is/ are incorrect for corrosion of iron?

(i) Reaction occurring at anode is

$$O_2(g) + 4H^{+}(aq) + 4e^{-} \longrightarrow 2H_2O(l)$$

(ii) Reaction occurring at cathode is

$$2Fe(s) \longrightarrow 2Fe^{2+} + 4e^{-}$$

(iii) Rust is $Fe_2O_3.xH_2O$

(iv) H^{+} involved in corrosion reaction is provided from H_2CO_3 which is formed due to dissolution of carbon dioxide from air in to water.

(a) (iv) only (b) (i) only

(c) (i) and (ii) (d) (i), (ii) and (iv)

MATCHING TYPE QUESTIONS

118. Match the Column-I (functioning of Daniel cell) with Column-II (value of E_{ext}) and choose the correct option.

Column-I	**Column-II**
(A) Flow of electrons from Cu to Zn and current flows from Zn to Cu	(p) $E = 1.1$ V
(B) No flow of electrons or current	(q) $E < 1.1$ V
(C) Zn dissolves at anode and copper deposits at cathode	(r) $E > 1.1$ V

(a) (A) – (q), (B) – (p), (C) – (r)
(b) (A) – (r), (B) – (p), (C) – (q)
(c) (A) – (p), (B) – (r), (C) – (q)
(d) (A) – (r), (B) – (q), (C) – (p)

119. Match the items of Column I and Column II on the basis of data given below :

$$E^{\Theta}_{F_2/F^-} = 2.87\ V,\ E^{\Theta}_{Li^+/Li} = -3.5\ V,\ E^{\Theta}_{Au^{3+}/Au} = 1.4\ V,$$

$$E^{\Theta}_{Br_2/Br^-} = 1.09\ V$$

Column-I	**Column-II**
(A) F_2	(p) metal is the strongest reducing agent.
(B) Li	(q) anion that can be oxidised by Au^{3+}
(C) Au^{3+}	(r) non metal which is the best oxidising agent
(D) Br^-	(s) metal ion which is an oxidising agent

(a) (A) – (r), (B) – (p), (C) – (s), (D) – (q)
(b) (A) – (p), (B) – (r), (C) – (s), (D) – (q)
(c) (A) – (q), (B) – (p), (C) – (s), (D) – (r)
(d) (A) – (r), (B) – (s), (C) – (p), (D) – (q)

120. Match the columns.

Column-I	**Column-II**
(A) Λ_m	(p) intensive property
(B) E^{Θ}_{cell}	(q) Depends on number of ions/ volume
(C) κ	(r) Extensive property
(D) $\Delta_r G_{cell}$	(s) Increases with dilution

(a) (A) – (p), (B) – (s), (C) – (q), (D) – (r)
(b) (A) – (s), (B) – (p), (C) – (q), (D) – (r)
(c) (A) – (s), (B) – (q), (C) – (p), (D) – (r)
(d) (A) – (s), (B) – (p), (C) – (r), (D) – (q)

121. Match the columns.

Column-I	**Column-II**
(A) κ	(p) $I \times t$
(B) Λ_m	(q) Λ_m / Λ_m^o
(C) α	(r) $\frac{\kappa}{c}$
(D) Q	(s) $\frac{G^*}{R}$

(a) (A) – (p), (B) – (r), (C) – (q), (D) – (s)
(b) (A) – (s), (B) – (q), (C) – (r), (D) – (p)
(c) (A) – (r), (B) – (s), (C) – (q), (D) – (p)
(d) (A) – (s), (B) – (r), (C) – (q), (D) – (p)

122. Match the columns

Column-I	**Column-II**
(A) Cell in which electrolyte is a paste of KOH and ZnO. This cell is used in low current devices like hearing aids, watches, etc.	(p) $H_2 - O_2$ fuel cell
(B) Cell in which 38% H_2SO_4 solution is used as an electrolyte.	(q) Mercury cell
(C) Cell in which vapours produced during electrochemical reaction were condensed and added to drinking water	(r) Lead storage battery
(D) Cell having longer life than lead storage cell and is expensive to manufacture	(s) Nickel – Cadmium cell

(a) (A) – (r), (B) – (q), (C) – (p), (D) – (s)
(b) (A) – (q), (B) – (r), (C) – (p), (D) – (s)
(c) (A) – (q), (B) – (r), (C) – (p), D) – (s)
(d) (A) – (q), (B) – (r), (C) – (s), (D) – (p)

ASSERTION-REASON TYPE QUESTIONS

Directions : Each of these questions contain two statements, Assertion and Reason. Each of these questions also has four alternative choices, only one of which is the correct answer. You have to select one of the codes (a), (b), (c) and (d) given below.

(a) Assertion is correct, reason is correct; reason is a correct explanation for assertion.
(b) Assertion is correct, reason is correct; reason is not a correct explanation for assertion
(c) Assertion is correct, reason is incorrect
(d) Assertion is incorrect, reason is correct.

123. Assertion : The resistivity for a substance is its resistance when it is one meter long and its area of cross section is one square meter.

Reason : The SI units of resistivity is ohm metre (Ωm).

124. **Assertion :** On increasing dilution, the specific conductance keep on increasing.

Reason : On increasing dilution, degree of ionisation of weak electrolyte increases and molality of ions also increases.

125. **Assertion :** Galvanised iron does not rust.

Reason : Zinc has a more negative electrode potential than iron.

CRITICAL THINKING TYPE QUESTIONS

126. If salt bridge is removed from two half-cells the voltage

(a) drops to zero (b) does not change

(c) increases gradually (d) increases rapidly

127. In the electrolytic cell, flow of electrons is from

(a) cathode to anode in solution

(b) cathode to anode through external supply

(c) cathode to anode through internal supply

(d) anode to cathode through internal supply

128. Standard potentials (E^o) for some half-reactions are given below :

(i) $Sn^{4+} + 2e^- \rightarrow Sn^{2+}$; $E^o = +0.15$ V

(ii) $2Hg^{2+} + 2e^- \rightarrow Hg_2^{2+}$; $E^o = +0.92$ V

(iii) $PbO_2 + 4H^+ + 2e^- \rightarrow Pb^{2+} + 2H_2O$; $E^o = +1.45$ V

Based on the above, which one of the following statements is correct ?

(a) Sn^{4+} is a stronger oxidising agent than Pb^{4+}

(b) Sn^{2+} is a stronger reducing agent than Hg_2^{2+}

(c) Hg^{2+} is a stronger oxidising agent than Pb^{4+}

(d) Pb^{2+} is a stronger reducing agent than Sn^{2+}

129. Consider the following relations for emf of a 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 = (Reduction 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) (ii) and (iv) (b) (iii) and (i)

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

130. The correct order of $E^\circ_{M^{2+}/M}$ values with negative sign for the four successive elements Cr, Mn, Fe and Co is

(a) Mn > Cr > Fe > Co (b) Cr < Fe > Mn > Co

(c) Fe > Mn > Cr > Co (d) Cr > Mn > Fe > Co

131. Consider the following four electrodes:

P = Cu^{2+} (0.0001 M)/Cu(s)

Q = Cu^{2+} (0.1 M)/Cu(s)

R = Cu^{2+} (0.01 M)/Cu(s)

S = Cu^{2+} (0.001 M)/Cu(s)

If the standard reduction potential of Cu^{2+}/Cu is +0.34 V, the reduction potentials in volts of the above electrodes follow the order.

(a) P > S > R > Q (b) S > R > Q > P

(c) R > S > Q > P (d) Q > R > S > P

132. At 298K the standard free energy of formation of $H_2O(\ell)$ is –237.20 kJ/mole while that of its ionisation into H^+ ion and hydroxyl ions is 80 kJ/mole, then the emf of the following cell at 298 K will be

[Take Faraday constant F = 96500C]

$H_2(g, 1\ bar) | H^+(1M) || OH^-(1M) | O_2(g, 1\ bar)$

(a) 0.40V (b) 0.81V

(c) 1.23 V (d) –0.40 V

133. If the following half cells have E° values as

$A^{3+} + e^- \longrightarrow A^{2+}$, $E^\circ = y_2$ V

$A^{2+} + 2e^- \longrightarrow A$, $E^\circ = -y_1$ V

The E° of the half cell $A^{3+} + 3e \longrightarrow A$ will be

(a) $\frac{2y_1 - y_2}{3}$ (b) $\frac{y_2 - 2y_1}{3}$

(c) $2y_1 - 3y_2$ (d) $y_2 - 2y_1$

134. $Cu^+(aq)$ is unstable in solution and undergoes simultaneous oxidation and reduction according to the reaction :

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

Choose correct E° for given reaction if E^o $Cu^{2+}/Cu = 0.34$ V and E^o $Cu^{2+}/Cu^+ = 0.15$ V

(a) –0.38 V (b) +0.49 V

(c) +0.38 V (d) –0.19 V

135. In a cell that utilises the reaction

$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$ addition of H_2SO_4 to cathode compartment, will

(a) increase the E and shift equilibrium to the right

(b) lower the E and shift equilibrium to the right

(c) lower the E and shift equlibrium to the left

(d) increase the E and shift equilibrium to the left

136. For a cell reaction involving two electron change, the standard EMF of the cell is 0.295 V at 2°C. The equilibrium constant of the reaction at 25°C will be:

(a) 29.5×10^{-2} (b) 10

(c) 1×10^{10} (d) 2.95×10^{-10}

137. Standard cell voltage for the cell $Pb | Pb^{2+} || Sn^{2+} | Sn$ is – 0.01 V. If the cell is to exhibit $E_{cell} = 0$, the value of $[Sn^{2+}] / [Pb^{2+}]$ should be antilog of –

(a) +0.3 (b) 0.5

(c) 1.5 (d) –0.5

138. The cell, $Zn | Zn^{2+}(1\ M) || Cu^{2+}(1\ M) | Cu$ ($E^\circ_{cell} = 1.10$ V) was allowed to be completely discharged at 298 K. The relative concentration of Zn^{2+} to Cu^{2+} $\left(\frac{[Zn^{2+}]}{[Cu^{2+}]}\right)$ is

(a) 9.65×10^4 (b) antilog (24.08)

(c) 37.3 (d) $10^{37.3}$.

139. What is the potential of half-cell consisting of zinc electrode in 0.01 M $ZnSO_4$ solution at 25°C $\left(E^\circ_{ox} = 0.763\ V\right)$

(a) 0.8221 V (b) 8.221 V
(c) 0.5282 V (d) 9.282 V

140. The oxidation potential of 0.05 M H_2SO_4 is

(a) -2×0.0591 (b) -0.01×0.0591
(c) -2.321×0.0591 (d) $+1 \times 0.0591$

141. For a relation

$$\Delta_r G = -nFE_{cell}$$

$E_{cell} = E^\circ_{cell}$ in which of the following condition?

(a) Concentration of any one of the reacting species should be unity
(b) Concentration of all the product species should be unity.
(c) Concentration of all the reacting species should be unity.
(d) Concentration of all reacting and product species should be unity.

142. A 0.5 M NaOH solution offers a resistance of 31.6 ohm in a conductivity cell at room temperature. What shall be the approximate molar conductance of this NaOH solution if cell constant of the cell is 0.367 cm^{-1}.

(a) 234 S cm^2 $mole^{-1}$ (b) 23.2 S cm^2 $mole^{-1}$
(c) 4645 S cm^2 $mole^{-1}$ (d) 5464 S cm^2 $mole^{-1}$

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

$$\Lambda^\circ_{CH_3COONa} = 91.0\ S\ cm^2 / equiv.$$

$$\Lambda^\circ_{HCl} = 426.2\ S\ cm^2 / equiv.$$

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

(a) Λ° of chloroacetic acid ($ClCH_2COOH$)
(b) Λ° of NaCl
(c) Λ° of CH_3COOK
(d) the limiting equivalent coductance of $H^+ (\lambda^\circ_{H^+})$.

144. Resistance of 0.2 M solution of an electrolyte is 50 Ω. The specific conductance of the solution is 1.3 S m^{-1}. If resistance of the 0.4 M solution of the same electrolyte is 260 Ω, its molar conductivity is :

(a) 6.25×10^{-4} S m^2 mol^{-1} (b) 625×10^{-4} S m^2 mol^{-1}
(c) 62.5 S m^2 mol^{-1} (d) 6250 S m^2 mol^{-1}

145. The limiting molar conductivities of HCl, CH_3COONa and NaCl are respectively 425, 90 and 125 mho cm^2 mol^{-1} at 25°C. The molar conductivity of 0.1 M CH_3COOH solutions is 7.8 mho cm^2 mol^{-1} at the same temperature. The degree of dissociation of 0.1 M acetic acid solution at the same temperature is

(a) 0.10 (b) 0.02
(c) 0.15 (d) 0.03

146. A weak electrolyte having the limiting equivalent conductance of 400 S cm^2. $equivalent^{-1}$ at 298 K is 2% ionized in its 0.1 N solution. The resistance of this solution (in ohms) in an electrolytic cell of cell constant 0.4 cm^{-1} at this temperature is

(a) 200 (b) 300
(c) 400 (d) 500

147. Conductivity κ, is equal to ________.

(i) $\frac{1}{R}\frac{l}{A}$ (ii) $\frac{G^*}{R}$
(iii) Λ_m (iv) $\frac{l}{A}$

(a) (i) and (iii) (b) (i) and (ii)
(c) (i), (ii) and (iii) (d) (ii), (iii) and (iv)

148. Arrange the following in increasing order of their conductivity Na^+ (A), K^+ (B), Ca^{2+} (C), Mg^{2+} (D)

(a) A, B, C, D (b) B, A, C, D
(c) C, A, D, B (d) A, B, D, C

149. The conductivity of electrolytic solutions depends upon which of the following?

(i) Size of ions produced
(ii) Viscosity of the solvent
(iii) Concentration of electrolyte
(iv) Solvation of ions produced

(a) (i) and (iii) (b) (i), (ii) and (iii)
(c) (i), (iii) and (iv) (d) All of these

150. Mark the false statement?

(a) A salt bridge is used to eliminate liquid junction potential
(b) The Gibbs free energy change, ΔG is related with electromotive force E as ΔG = –nFE
(c) Nernst equation for single electrode potential is

$$E = E^o - \frac{RT}{nF}\log_a M^{n+}$$

(d) The efficiency of a hydrogen-oxygen fuel cell is 23%

151. When electric current is passed through acidified water, 112 ml of hydrogen gas at STP collected at the cathode in 965 seconds. The current passed in amperes is

(a) 1.0 (b) 0.5
(c) 0.1 (d) 2.0

152. On passing current through two cells, connected in series containing solution of $AgNO_3$ and $CuSO_4$, 0.18 g of Ag is deposited. The amount of the Cu deposited is:

(a) 0.529 g (b) 10.623 g
(c) 0.0529 g (d) 1.2708 g

153. In the electrolysis of water, one faraday of electrical energy would liberate

(a) one mole of oxygen (b) one gram atom of oxygen
(c) 8 g oxygen (d) 22.4 lit. of oxygen

154. Electrolysis of dilute aqueous NaCl solution was carried out by passing 10 milli ampere current. The time required to liberate 0.01 mol of H_2 gas at the cathode is
(1 Faraday = 96500 C mol^{-1})

(a) 9.65×10^4 sec (b) 19.3×10^4 sec
(c) 28.95×10^4 sec (d) 38.6×10^4 sec

155. What is the amount of chlorine evolved when 2 amperes of current is passed for 30 minutes in an aqueous solution of NaCl ?

(a) 66 g (b) 1.32 g

(c) 33 g (d) 99 g

156. On passing a current of 1.0 ampere for 16 min and 5 sec through one litre solution of $CuCl_2$, all copper of the solution was deposited at cathode. The strength of $CuCl_2$ solution was (Molar mass of Cu= 63.5; Faraday constant = 96,500 $Cmol^{-1}$)

(a) 0.01 N (b) 0.01 M

(c) 0.02 M (d) 0.2 N

157. 0.2964 g of copper was deposited on passage of a current of 0.5 amp for 30 mins through a solution of copper sulphate. Calculate the oxidation state of Cu (At. mass 63.56).

(a) +1 (b) +2

(c) +3 (d) +4

158. One Faraday of electricity is passed through molten Al_2O_3, aqueous solution of $CuSO_4$ and molten NaCl taken in three different electrolytic cells connected in series. The mole ratio of Al, Cu and Na deposited at the respective cathode is

(a) 2 : 3 : 6 (b) 6 : 2 : 3

(c) 6 : 3 : 2 (d) 1 : 2 : 3

159. What will happen during the electrolysis of aqueous solution of $CuSO_4$ by using platinum electrodes ?

(i) Copper will deposit at cathode.

(ii) Copper will deposit at anode.

(iii) Oxygen will be released at anode.

(iv) Copper will dissolve at anode.

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

(c) (i) and (ii) (d) (ii) and (iii)

160. How much charge is required, when 1 mole of $Cr_2O_7^{2-}$ reduce to form 1 mole of Cr^{3+} ?

(a) 6F (b) 3F

(c) 1F (d) 2F

161. When a lead storage battery is discharged

(a) SO_2 is evolved

(b) Lead sulphate is consumed

(c) Lead is formed

(d) Sulphuric acid is consumed

162. The most durable metal plating on iron to protect against corrosion is

(a) nickel plating (b) copper plating

(c) tin plating (d) zinc plating

163. Which of the following statements is incorrect regarding dry (Leclanche) cell?

(a) Cathode used in the cell is coated by powdered manganese dioxide and carbon.

(b) Most common application of this cell is in our transistors and clocks.

(c) At cathode, Mn is oxidised from + 3 to + 4.

(d) At anode Zn is oxidised from 0 to + 2.

HINTS AND SOLUTIONS

FACT/DEFINITION TYPE QUESTIONS

1. **(c)** Electrochemical reactions are more energy efficient and less polluting. Thus study of electrochemistry is important to create new eco-friendly technologies.

2. **(a)** We know that in the internal circuit of a galvanic cell ions flow whereas in the external circuit, the electrons flow from one electrode to another.

3. **(a)** Anode has negative polarity.

4. **(b)**

5. **(d)** When both the electrodes are kept in the same solution there will be no requirement of salt bridge.

6. **(d)**

7. **(d)** $\overbrace{Zn^{2+} + 2e^- \longrightarrow Zn}^{\text{Reduction}}$ It shows reduction reaction.

8. **(a)**

9. **(b)** The cell in which Cu and Zn rods are dipped in its solution is called Daniel cell.

10. **(c)** The magnitude of the electrode potential of a metal is a measure of its relative tendency to loose or gain electrons. i.e., it is a measure of the relative tendency to undergo oxidation (loss of electrons) or reduction (gain of electrons).

$M \rightarrow M^{n+} + ne^-$ (oxidation potential)

$M^{n+} + ne^- \rightarrow M$ (reduction potential)

11. **(b)** $2AgCl(s) + H_2(g) \rightarrow 2HCl(aq) + 2Ag(s)$

The activities of solids and liquids are taken as unity and at low concentrations, the activity of a solute is approximated to its molarity.

The cell reaction will be

$Pt(s) | H_2(g), 1bar | H^+(aq)1M | AgCl(aq)1M | Ag(s)$

12. **(b)** $E^o_{cell} = E^o_{cathode} - E^o_{anode} = E^o_{right} - E^o_{left}$

$E^o_{cell} = 0.25 - 0.52 = -0.27V$

13. **(b)** 14. **(c)**

15. **(d)** Cu is anode and Ag^+ is cathode.

16. **(d)** Calomel electrode is used as reference electrode.

17. **(b)** Electrode potential is considered as zero.

18. **(b)** Without losing its concentration $ZnCl_2$ solution cannot kept in contact with Al because Al is more reactive than Zn due to its highly negative electrode reduction potential.

19. **(c)** From the given data we find Fe^{3+} is strongest oxidising agent. More the positive value of E°, more is the tendency to get oxidized. Thus correct option is (c).

20. **(c)** As the value of standard reduction potential decreases the reducing power increases i.e.,

$$\underset{(-3.0)}{Z} > \underset{(-1.2)}{X} > \underset{(+0.5)}{Y}$$

21. **(b)** Given $E_{Sn^{+4}/Sn^{+2}} = +0.15\,V$

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

$$\begin{aligned} E^\circ_{cell} &= E^\circ_{ox} + E^\circ_{red} \\ &= 0.74 + 0.15 \\ &= 0.89\,V \end{aligned}$$

22. **(a)** Higher the value of reduction potential higher will be the oxidising power whereas the lower the value of reduction potential higher will be the reducing power.

23. **(d)** $E^\circ_{Cell} = E^\circ_{OP} + E^\circ_{RP}$

$= 0.76 + 0.314 = 1.10\,V$

24. **(b)** Follow ECS, A will replace B.

25. **(c)** Gold having higher E^o_{Red} and oxidises $Fe \longrightarrow Fe^{++}$.

26. **(c)** 27. **(d)**

28. **(c)** Nernst equation $E_{cell} = E^o_{cell} - \frac{RT}{nF} \frac{\ln[Fe^{2+}]^2}{[Fe^{3+}]^2}\left[Zn^{2+}\right]$

increasing $[Fe^{2+}]$ will decrease the E_{cell}.

29. **(d)** $E^\circ = \frac{0.0591}{n} \log K$

Here, $n = 2, E^\circ = 0.295$

$\therefore \log K = \frac{2 \times 0.295}{0.0591} = 9.98 \approx 10$ or $K = 10^{10}$

30. **(b)** RHS : $2H^+ + 2e^- \rightleftharpoons H_2(P_2)$

LHS : $H_2(P_1) \rightleftharpoons 2H^+ + 2e^-$

overall reaction : $H_2(P_1) \rightleftharpoons H_2(P_2)$

$$E = E^\circ - \frac{RT}{nF} \ell n \frac{P_2}{P_1} = 0 - \frac{RT}{nF} \ell n \frac{P_2}{P_1} = \frac{RT}{nF} \ell n \frac{P_1}{P_2}$$

31. **(a)** $E = E^\circ - \frac{0.059}{n} \log \frac{1}{[H^+]}$

$= 0 - \frac{0.059}{1} \log \frac{1}{10^{-4}} = -0.236V$

32. **(d)** $E_{cell} = E^\circ_{cell} - \frac{RT}{nF} \ln Q$

At equilibrium,

$E_{cell} = 0$ and $Q = K_c$

$\therefore \quad E_{cell} \neq E^\circ_{cell}$

33. **(c)** The E°_{cell} is given by

$$E^\circ_{\text{cell}} = \frac{0.0591}{n}\log K_{eq}$$

$$\therefore\ 0.591 = \frac{0.0591}{1}\log K_{eq}$$

$$\text{or}\ \log K_{eq} = \frac{0.591}{0.0591} = 10$$

$$\text{or}\ K_{eq} = 1\times 10^{10}$$

34. **(b)** For the given cell

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.059\text{V}}{2}\log\frac{[\text{Zn}^{2+}(aq)]}{[\text{Cu}^{2+}(aq)]}$$

The cell potential will decreases with increase in $[\text{Zn}^{2+}\ (aq)]$ and will increases with increase in $[\text{Cu}^{2+}(aq)]$.

35. **(d)** Here $n = 4$, and $[\text{H}^+] = 10^{-3}$ (as pH = 3)

Applying Nernst equation

$$\text{E} = \text{E}^\circ - \frac{0.059}{n}\log\frac{[\text{Fe}^{2+}]^2}{[\text{H}^+]^4(p_{O_2})}$$

$$= 1.67 - \frac{0.059}{4}\log\frac{(10^{-3})^2}{(10^{-3})^4\times 0.1}$$

$$= 1.67 - \frac{0.059}{4}\log 10^7 = 1.67 - 0.103 = 1.567$$

36. **(b)** Cell reaction is, $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF}\ell\text{n}\frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

Greater the factor $\left[\frac{(\text{Zn}^{2+})}{(\text{Cu}^{2+})}\right]$, less is the EMF

Hence $E_1 > E_2$

37. **(b)** For $\text{Zn}^{2+} \rightarrow \text{Zn}$

$$\text{E}_{\text{Zn}^{2+}/\text{Zn}} = \text{E}^\circ_{\text{Zn}^{2+}/\text{Zn}} - \frac{2.303\text{RT}}{\text{nF}}\log\frac{[\text{Zn}]}{[\text{Zn}^{2+}]}$$

$$= -0.76 - \frac{0.06}{2}\log\frac{1}{[0.1]} = -0.76 - 0.03$$

$$\text{E}_{\text{Zn}^{2+}/\text{Zn}} = -0.79\text{V}$$

38. **(d)** emf will decrease.

39. **(b)** $\text{E}_{\text{cell}} = \text{E}^\circ_{\text{cell}} - \dfrac{\text{RT}}{\text{nF}}\ln\dfrac{[\text{C}]^c[\text{D}]^d}{[\text{A}]^a[\text{B}]^b}$.

Hence $\text{E}^\circ_{\text{cell}} = \text{E}_{\text{cell}} + \dfrac{\text{RT}}{\text{nF}}\ln\dfrac{[\text{C}]^c[\text{D}]^d}{[\text{A}]^a[\text{B}]^b}$

40. **(b)** $\text{E}^\circ_{\text{cell}} = \dfrac{0.059}{2}\log\text{K}_\text{C}$ or $\dfrac{1.10\times 2}{0.059} = \log\text{K}_\text{C}$

$$\therefore \text{K}_\text{C} = 1.9\times 10^{37}$$

41. **(c)** $\Delta\text{G}^\circ = -\text{nFE}^\circ;\ \text{E}^\circ = \dfrac{-\Delta\text{G}^\circ}{\text{nF}}$;

$$\text{E}^\circ = \frac{-(-50.61\text{kJ})}{2\times 96500\times 10^{-3}} = 0.26\text{V}$$

42. **(d)** Molarity = 0.01 M ; Resistance = 40 ohm;

Cell constant $\dfrac{l}{A} = 0.4\text{cm}^{-1}$.

Specific conductivity (κ)

$$= \frac{\text{cell constant}}{\text{resistance}} = \frac{0.4}{40} = 0.01\ \text{ohm}^{-1}\ \text{cm}^{-1}$$

$$\text{Molar conductance}\ (\wedge_\text{m}) = \frac{1000\kappa}{\text{Molarity}}$$

$$= \frac{1000\times 0.01}{.01} = 10^3\ \text{ohm}^{-1}\ \text{cm}^2\ \text{mol}^{-1}$$

43. **(b)** Specific conductance of the solution $(\kappa) = 0.012\ \text{ohm}^{-1}\ \text{cm}^{-1}$ and resistance (R) = 55 ohm.

Cell constant = Specific conductance × Observed resistance

$$= 0.012\times 55 = 0.66\ \text{cm}^{-1}.$$

44. **(b)** $\text{ohm}^{-1}\ \text{cm}^2\ (\text{geq})^{-1}$

45. **(a)** $\Lambda_{eq} = \kappa\times\dfrac{1000}{\text{N}} = \dfrac{1}{\text{R}}\times\dfrac{l}{\text{a}}\times\dfrac{1000}{\text{N}}$

$$= \frac{1}{\text{R}}\times\text{cell constant}\times\frac{1000}{\text{N}} = \frac{1}{220}\times 0.88\times\frac{1000}{0.01}$$

$$= 400\ \text{mho cm}^2\ \text{g eq}^{-1}$$

46. **(c)** Molar conductance of solution is related to specific conductance as follows :

$$\wedge_\text{m} = \kappa\times\frac{1000}{C} \qquad \text{....(a)}$$

where C is molar concentration.

Putting $\kappa = 6.3\times 10^{-2}\ \text{ohm}^{-1}\ \text{cm}^{-1}$ and $C = 0.1\text{M}$

$$\wedge_\text{m} = (6.3\times 10^{-2}\,\text{ohm}^{-1}\,\text{cm}^{-1})\times\frac{1000}{(0.1\text{mol}/\text{cm}^3)}$$

$$= 6.3\times 10^{-2}\times 10^4\ \text{ohm}^{-1}\text{cm}^2\ \text{mol}^{-1}$$

$$= 630\ \text{ohm}^{-1}\ \text{cm}^2\ \text{mol}^{-1}$$

47. **(b)** $\kappa = \dfrac{1}{\text{R}}\times\text{Cell constant}$

Cell constant $= \kappa\times\text{R};\ 0.012\times 55 = 0.66\ \text{cm}^{-1}$.

48. **(d)** $\text{ohm}^{-1}\ \text{cm}^{-1}$

49. **(a)** The specific conductance increases with concentration. The number of ions per cm^{-3} increase with increase of concentration.

50. **(d)**

51. **(c)** Correct matching for pair (iii) will be [G (conductance) – siemens or ohm^{-1}(S).]

52. **(d)** Conductivity does not depend upon mass or weight of material.

53. **(d)**

54. **(c)** (i) represents weak electrolyte
(ii) represents strong electrolyte.

55. **(b)** $\Lambda^{\infty}_{HCl} = 426.2 \quad \text{(i)}$

$\Lambda^{\infty}_{AcONa} = 91.0 \quad \text{(ii)}$

$\Lambda^{\infty}_{NaCl} = 126.5 \quad \text{(iii)}$

$\Lambda^{\infty}_{AcOH} = \text{(i)} + \text{(ii)} - \text{(iii)}$

$= [426.2 + 91.0 - 126.5] = 390.7$

56. **(d)** Kohlrausch's Law states that at infinite dilution, each ion migrates independent of its co-ion and contributes to the total equivalent conductance of an electrolyte a definite share which depends only on its own nature. From this definition we can see that option (d) is the correct answer.

57. **(c)** Conductivity of an electrolyte depends on the mobility of ions and concentration of ions. The motion of an ionic species in an electric field is retarded by the oppositely charged ions due to their interionic attraction. On dilution, concentration of electrolyte decreases and the retarding influence of oppositely charged ions decreases. Therefore mobility of ions increases.

58. **(d)** $\Lambda^{\circ}_{m(NH_4Cl)} = \Lambda^{\circ}_{m_{NH_4^+}} + \Lambda^{\circ}_{m_{Cl^-}}$

$\Lambda^{\circ}_{m(NaOH)} = \Lambda^{\circ}_{m_{Na^+}} + \Lambda^{\circ}_{m_{OH^-}}$

$\Lambda^{\circ}_{m(NaCl)} = \Lambda^{\circ}_{m_{Na^+}} + \Lambda^{\circ}_{m_{Cl^-}}$

$\therefore \ \Lambda^{\circ}_{m\left(NH_4^+\right)} + \Lambda^{\circ}_{m\left(OH^-\right)}$

$= \Lambda^{\circ}_{m\left(NH_4^+\right)} + \Lambda^{\circ}_{m\left(Cl^-\right)} + \Lambda^{\circ}_{m\left(Na^+\right)}$

$+ \Lambda^{\circ}_{m\left(OH^-\right)} - \left[\Lambda^{\circ}_{m\ Na^+} + \Lambda^{\circ}_{m\ Cl^-}\right]$

$\Lambda^{\circ}_{m(NH_4OH)} = \Lambda^{\circ}_{m\ NH_4Cl} + \Lambda^{\circ}_{m(NaOH)} - \Lambda^{\circ}_{m(NaCl)}$

59. **(d)** $\Lambda^{\circ}_{CH_3COOH} = \Lambda^{\circ}_{CH_3COONa} + \Lambda^{\circ}_{HCl} - \Lambda^{\circ}_{NaCl}$

$= 91 + 425.9 - 126.4 = 390.5 \text{ S cm}^2\text{mol}^{-1}$

60. **(b)** $\Lambda^{\circ}_{Na^+} + \Lambda^{\circ}_{OH^-} = 248 \times 10^{-4} \text{Sm}^2\text{mol}^{-1}$

$\Lambda^{\circ}_{Na^+} + \Lambda^{\circ}_{Cl^-} = 126 \times 10^{-4} \text{Sm}^2\text{mol}^{-1}$

$\Lambda^{\circ}_{Ba^{2+}} + \Lambda^{\circ}_{2Cl^-} = 280 \times 10^{-4} \text{ S m}^2\text{mol}^{-1}$

Now, $\Lambda^{\circ}_{Ba(OH)_2} = \Lambda^{\circ}_{BaCl_2} + 2\Lambda^{\circ}_{NaOH} - 2\Lambda^{\circ}_{NaCl}$

$\Lambda^{\circ}_{Ba(OH)_2} = 280 \times 10^{-4} + 2 \times 248 \times 10^{-4} - 2 \times 126 \times 10^{-4}$

$\Lambda^{\circ}_{Ba(OH)_2} = 524 \times 10^{-4} \text{Sm}^2\text{mol}^{-1}.$

61. **(c)** $ClCH_2COONa + HCl \rightarrow ClCH_2COOH + NaCl$

$\Lambda_{ClCH_2COONa} + \Lambda_{HCl} = \Lambda_{ClCH_2COOH} + \Lambda_{NaCl}$

$224 + 203 = \Lambda_{ClCH_2COOH} + 38.2$

$\Lambda_{ClCH_2COOH} = 427 - 38.2$

$= 388.8 \text{ ohm}^{-1}\text{cm}^2\text{gm eq}^{-1}$

62. **(b)** $\alpha = \dfrac{\Lambda_M}{\Lambda^{\infty}_M} = \dfrac{9.54}{238} = 0.04008 = 4.008\,\%.$

63. **(d)** Cell constant = l/a
Unit = $\text{m/m}^2 = \text{m}^{-1}$.

64. **(d)** Larger the size, lower the speed.

65. **(a)** $\Lambda^{\infty}_m = 57 + 73 = 130 \text{ Scm}^2\text{mol}^{-1}$

66. **(b)** At anode :

$$2HO^- \rightleftharpoons H_2O + \frac{1}{2}O_2$$

67. **(b)** Given current (i) = 0.5 amp;
Time (t) = 100 minutes × 60 = 6000 sec
Equivalent weight of silver nitrate (E) = 108.
According to Faraday's first law of electrolysis

$$W = \frac{Eit}{96500} = \frac{108 \times 0.5 \times 6000}{96500} = 3.3575 \text{ g.}$$

68. **(d)** 1 mole of $e^- = 1F = 96500$ C
27g of Al is deposited by 3 × 96500 C
5120 g of Al will be deposited by

$$= \frac{3 \times 96500 \times 5120}{27} = 5.49 \times 10^7 \text{C}$$

69. **(c)** Electrorefining and electroplating are done by electrolysis.

70. **(d)** $\dfrac{W_A}{E_A} = \dfrac{W_B}{E_B}; \ \dfrac{1.6}{8} = \dfrac{\text{Wt. of Ag}}{108}$

$\therefore$ Wt. of Ag = 21.6 g

71. **(b)**

$$\begin{array}{ccc} Cu^{+2} + & 2e^- & \rightarrow Cu(s) \\ & 2 \times 96500 \text{ C} & 63.6\text{g} \end{array}$$

$$9650 \text{ C will deposit} = \frac{63.6}{2 \times 96500} \times 9650 = 3.18 \text{ g}$$

72. **(a)** Charge $= 0.2 \times 1$ Faraday
$= 0.2 \times 96500$ coulombs
$= 19300 = 1.93 \times 10^4$ coulombs

73. **(d)** $Ag^+ + e^- \longrightarrow Ag$

96500 coulombs deposit = 108 g of Ag

$\therefore$ 965 coulombs deposit $= \frac{108}{96500} \times 965 = 1.08$ g Ag

74. **(d)** $C_6H_5NO_2 + 6H^+ + 6e^- \longrightarrow C_6H_5NH_2 + 2H_2O$

$E_{C_6H_5NO_2}(\text{eq.wt}) = \frac{123}{6} = 20.5$

Number of coulombs required $= \frac{w \times 96500}{\text{Eq. wt}}$

$= \frac{12.3 \times 96500}{20.5} = 57900$ C

75. **(a)** According to Faraday law's of electrolysis, amount of electricity required to deposit 1 mole of metal $= 96500\,C = 1\,F$

i.e., for deposition of 108g Ag electricity required = 1 F

76. **(a)** For deposition of silver, reaction is

$Ag^+ + e^- \longrightarrow Ag$

1 mol of Ag will be deposited by
$= 1\,F = 96500\,C = 9.65 \times 10^4\,C$

Since 1 equivalent weight of Ag is also equal to the weight of its 1 mol, hence 1 equivalent weight of Ag will be deposited by $= 9.65 \times 10^4$ C

77. **(b)** A current of 96500 coulombs liberate 1 mole of O_2.

$\Rightarrow$ 96500 C liberates = 22.4 L of O_2 at NTP

$\Rightarrow$ 9650 C liberates $= \frac{22.4}{96500} \times 9650$
$= 2.24$ L of O_2 at NTP

78. **(b)** $Fe^{2+} + 2e^- \rightarrow Fe;\ E_{Fe} = \frac{56}{2} = 28$

1 Faraday liberates = 28 g of Fe

3 Faraday liberates $= 3 \times 28 = 84$ gm

79. **(c)**

80. **(d)** Charge (Coulombs) pass per second $= 10^{-6}$

number of electrons passed per second

$= \frac{10^{-6}}{1.602 \times 10^{-19}} = 6.24 \times 10^{12}$

81. **(d)**

82. **(d)** Charge on one mole of electrons = 96500 C.

83. **(a)** When platinum electrodes are dipped in dilute solution H_2SO_4 than H_2 is evolved at cathode.

84. **(b)** In electrolysis of NaCl when Pt electrode is taken then H_2 liberated at cathode while with Hg cathode it forms sodium amalgam because more voltage is required to reduce H^+ at Hg than Pt.

85. **(a)** When molten or fused NaCl is electrolysed, it yields metallic sodium and gaseous chlorine. Reactions involved are as follows:

$NaCl\,(s) \longrightarrow Na^+\,(l) + Cl^-\,(l)$
$Na^+ + e^- \longrightarrow Na\,(l)$ (at cathode)
$Cl^-\,(l) \longrightarrow Cl\,(g) + e^-$ (at anode)
$Cl\,(g) + Cl(g) \longrightarrow Cl_2\,(g)$

86. **(b)** $Pt^{4+} + 4e^- \longrightarrow Pt$

4F electricity is required to deposit 1 mole of Pt.

$\therefore$ 0.80 F of electricity will deposit
$= 1/4 \times 0.80$ moles of Pt $= 0.20$ mol.

87. **(d)** $1F \rightarrow 11.2\,L\,Cl_2$ at STP

$\therefore$ No. of Faradays $= \frac{9.65 \times 5 \times 60 \times 60}{96500} = 1.8$

$\therefore$ Vol. of $Cl_2 = 1.8 \times 11.2\,L = 20.16$

88. **(b)** W = Zit

where Z = Electrochemical equivalent

Eq. wt. of copper $= \frac{63}{2} = 31.5$

$Z = \frac{31.5}{96500}$

$W = Zit = \frac{31.5}{96500} \times 1.5 \times 10 \times 60 = 0.2938g$

89. **(c)** $\underset{0.1\text{ mole}}{\overset{(+6)}{MnO_4^{2-}}} \longrightarrow \overset{(+7)}{MnO_4^-} + e^-$

Quantity of electricity required
$= 0.1F = 0.1 \times 96500 = 9650\,C$

90. **(d)** $w_{O_2} = n_{O_2} \times 32$

$w_{O_2} = \frac{5600}{22400} \times 32 = 8g$
$= 1$ equivalent of O_2
$= 1$ equivalent of $Ag = 108$

91. **(a)** Electrolysis of these (i), (ii) and (iii) salt release Chlorine which is yellowish in colour while Br_2 is reddish brown in colour

92. **(b)** Electronic conductance decreases with increase in temperature whereas electrolytic conductance increases with increase in temperature as no. of ions or charge carriers increases with increase in temperature.

93. **(a)** Gold is an inert metal. Electrodes made up of inert metals does not participate in chemical reaction.

94. **(b)** Reaction A is preferred for electrolysis of dilute sulphuric acid and B is preferred for electrolysis of concentrated sulphuric acid.

95. **(b)** Iron can be easily reduced with carbon it does not require electrochemical reduction.

96. **(b)** H_2SO_4 is regenerated.

97. **(b)** During charging, the lead storage battery behaves like an electrolytic cell. So, at anode the reaction is

$PbSO_4 + 2H_2O \longrightarrow PbO_2 + 4H^+ + SO_4^{2-} + 2e^-$

98. **(c)** $2NH_4Cl + Zn \rightarrow 2NH_3 + ZnCl_2 + H_2 \uparrow$.

99. **(b)** In $H_2 - O_2$ fuel cell, the combustion of H_2 occurs to create potential difference between the two electrodes.

100. **(d)** Primary cells are those cells, in which the reaction occurs only once and after use over a period of time, it becomes dead and cannot be reused again. e.g., Leclanche cell and mercury cell.

101. **(c)** The electrolyte used in Leclanche cell is moist paste of NH_4Cl and $ZnCl_2$.

102. **(d)** A device that converts energy of combustion of fuels, directly into electrical energy is known as fuel cell.

103. **(c)**

104. **(b)** $H_2 - O_2$ fuel cell supply power for pressure.

105. **(d)** Prevention of corrosion by zinc coating is called galvanization.

106. **(a)** Cathodic protection is best method to prevent iron from rusting. In this method iron is made cathode by application of external current.

Saline water is highly conducting and hence accelerates the formation of rust.

107. **(b)** Magnesium provides cathodic protection and prevent rusting or corrosion.

108. **(c)** Mercury cell being primary in nature can be used only once.

109. **(c)**

110. **(d)** Fuel cells produce electricity with an efficiency of about 70% compared to thermal plants whose efficiency is about 40%.

STATEMENT TYPE QUESTIONS

111. **(a)** Statement (ii) is false as in Figure 2 equal potential is applied from either side thereby making his cell non-functional. Figure 3 represents electrolytic cell as current flows from anode to cathode.

112. **(d)** According to an accepted convention anode is written on the left side and cathode on the right while representing the galvanic cell.

113. **(c)** Because reduction potential of water is higher than that of Na^+ so H_2 will be evolved and no reduction of metal ions occurs.

114. **(a)** All statements given above are correct.

115. **(d)** CaX_2 is a strong electrolyte thus its molar conductivity increases slowly on dilution.

116. **(c)** We consider,

$96487 Cmol^{-1} \simeq 96500 Cmol^{-1}$ for approximate calculations not for accurate calculations.

117. **(c)** Reaction occurring at anode

$$2Fe(s) \longrightarrow 2Fe^{2+} + 4e^-$$

Reaction occuring at cathode

$$O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O$$

MATCHING TYPE QUESTIONS

118. **(b)** **119.** **(a)** **120.** **(b)** **121.** **(d)** **122.** **(b)**

ASSERTION-REASON TYPE QUESTIONS

123. **(b)** We know, $R \propto \frac{\ell}{A}$ or $R = \rho\left(\frac{\ell}{A}\right)$, where proportionality constant ρ is called resistivity. If $\ell = 1m$ and $A = 1m^2$, then $R = \rho$ i.e., Resistance = Resistivity.

124. **(d)** The specific conductivity decreases while equivalent and molar conductivities increase with dilution.

125. **(a)** Zinc metal which has a more negative electrode potential than iron will provide electrons in preference of the iron, and therefore corrodes first. Only when all the zinc has been oxidised, the iron start to rust.

CRITICAL THINKING TYPE QUESTIONS

126. **(a)** Salt bridge allows the flow of current by completing circuit. No current will flow and voltage will drop to zero, if salt bridge is removed.

127. **(d)** In electrolytic cell the flow of electrons is from anode to cathode through internal supply.

128. **(b)** $E^\circ_{Sn^{2+}/Sn^{4+}} = -0.15V > E^\circ_{Hg_2^{2+}/Hg^{2+}} = -0.92$

Hence, Sn^{2+} is stronger reducing agent than Hg_2^{2+}.

129. **(a)** Option (ii) and (iv) are correct.

130. **(a)** The value of $E^\circ_{M^{2+}/M}$ for given metal ions are

$E^\circ_{Mn^{2+}/Mn} = -1.18$ V,

$E^\circ_{Cr^{2+}/Cr} = -0.9$ V,

$E^\circ_{Fe^{2+}/Fe} = -0.44$ V and

$E^\circ_{Co^{2+}/Co} = -0.28$ V.

The correct order of $E^\circ_{M^{2+}/M}$ values without considering negative sign would be

$$Mn^{2+} > Cr^{2+} > Fe^{2+} > Co^{2+}.$$

131. **(d)** $E_{red} = E^\circ_{red} + \frac{0.591}{n} \log[M^{n+}]$

Lower the concentration of M^{n+}, lower is the reduction potential.

Hence order of reduction potential is :

$$Q > R > S > P$$

132. (a) Cell reaction

cathode : $H_2O(\ell)+\frac{1}{2}O_2(g)+2e^- \longrightarrow 2OH^-(aq)$

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

$$H_2O(\ell)+\frac{1}{2}O_2(g)+H_2(g)\longrightarrow 2H^+(aq)+2OH^-(aq)$$

Also we have

$$H_2(g)+\frac{1}{2}O_2(g)\longrightarrow H_2O(\ell) \quad ;$$

$$\Delta G_f^\circ = -237.2\ \text{kJ/mole}$$

$$H_2O(\ell)\longrightarrow H^+(aq)+OH^-(aq);\ \Delta G^\circ = 80\ \text{kJ/mole}$$

Hence for cell reaction

$$\Delta G^\circ = -237.2+(2\times 80) = -77.20\ \text{kJ/mole}$$

$$\therefore\ E^\circ = -\frac{\Delta G^\circ}{nF} = \frac{77200}{2\times 96500} = 0.40\text{V}$$

133. (b) (i) $A^{3+}+e^-\longrightarrow A^{2+},\ \Delta G_1 = -1\,F\,y_2$

(ii) $A^{2+}+2e^-\longrightarrow A,\ \Delta G_2 = -2F(-y_1) = 2Fy_1$

Add, (i) and (ii), we get

$A^{3+}+3e^-\longrightarrow A;$

$\Delta G_3 = \Delta G_1+\Delta G_2$

$-3FE^\circ = -Fy_2+2Fy_1$

$-3FE^\circ = -F(y_2-2y_1)$

$$E^\circ = \frac{y_2-2y_1}{3}$$

134. (c) $2Cu^+\longrightarrow Cu^{+2}+Cu$

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

$e^-+Cu^{+2}\longrightarrow Cu^+\,;\ E_2^\circ = 0.15\text{V};$ ... (ii)

$Cu^++e^-\rightarrow Cu;\ E_3^\circ = ?$ ... (iii)

Now, $\Delta G_1^\circ = -nFE_1^\circ = -2\times 0.34F$

$\Delta G_2^\circ = -1\times 0.15F\,,\ \Delta G_3^\circ = -1\times E_3^\circ F$

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

$\Rightarrow -0.68F = -0.15F - E_3^\circ F$

$\Rightarrow E_3^\circ = 0.68-0.15 = 0.53\text{V}$

$E_{cell}^\circ = E_{cathode}^\circ(Cu^+/Cu) - E_{anode}^\circ(Cu^{+2}/Cu^+)$

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

135. (a) $Zn(s)+2H^+(aq)\rightleftharpoons Zn^{2+}(aq)+H_2(g)$

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

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

136. (c) Using the relation,

$$E^\circ_{cell} = \frac{2.303RT}{nF}\log K_c = \frac{0.0591}{n}\log K_c$$

$$\therefore\ 0.295\text{ V} = \frac{0.0591}{2}\log K_c$$

or $\log K_c = \dfrac{2\times 0.295}{0.0591} = 10$

or $K_c = 1\times 10^{10}$

137. (a) Apply Nernst equation to the reaction

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

or $E^\circ + \dfrac{0.059}{2}\log\dfrac{[Sn^{2+}]}{[Pb^{2+}]} = E_{cell}$

or $\log\dfrac{[Sn^{2+}]}{[Pb^{2+}]} = \dfrac{0.01\times 2}{0.059} = 0.3\ \ (\because E_{cell}=0)$

or $\dfrac{[Sn^{2+}]}{[Pb^{2+}]} = \text{antilog}\,(0.3)$

138. (d) $E_{cell} = 0$; when cell is completely discharged.

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

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

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

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

139. (a) The Half cell reaction is $Zn^{++}+2e^-\longrightarrow Zn$.

$$E_{cell} = E^\circ_{cell} - \frac{0.059}{2}\log\frac{1}{[Zn^{++}]}$$

$$= -0.763 - \frac{0.059}{2}\log\frac{1}{0.01} = -0.822\text{V}$$

$$E_{oxi} = 0.822\text{ V}$$

140. (d) $2H^++2e^-\longrightarrow H_2$

$$E_{Red} = E^\circ_{Red} - \frac{0.059}{n}\log\frac{1}{[H^+]^2}\ ;$$

$$E_{Red} = 0 - \frac{0.059}{2}\log\frac{1}{(0.1)^2}\ ;\ E_{Red} = -0.059\text{ V},$$

$E_{oxi} = 0.059$ V.

141. (c) When the concentration of all reacting species kept unity, then $E_{cell} = E^\circ_{cell}$ and the given relation will become

$\Delta_r G = -nFE^\circ_{cell}$

142. (b) Here, $R = 31.6$ ohm

$\therefore$ Conductance $= \frac{1}{R} = \frac{1}{31.6} \text{ohm}^{-1} = 0.0316 \text{ ohm}^{-1}$

Specific conductance = conductance × cell constant.
$= 0.0316 \text{ ohm}^{-1} \times 0.367 \text{ cm}^{-1}$
$= 0.0116 \text{ ohm}^{-1} \text{ cm}^{-1}$
Now, molar concentration $= 0.5$M (given)
$= 0.5 \times 10^{-3}$ mole cm^{-3}

$\therefore$ Molar conductance $= \frac{\kappa}{C} = \frac{0.0116}{0.5 \times 10^{-3}}$

$= 23.2 \text{ S cm}^2 \text{ mol}^{-1}$

143. (b) According to Kohlrausch's law, molar conductivity of weak electrolyte acetic acid (CH_3COOH) can be calculated as follows:

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

$\therefore$ Value of Λ°_{NaCl} should also be known for calculating value of $\Lambda^\circ_{CH_3COOH}$.

144. (a) $\kappa = \frac{1}{R} \times \frac{\ell}{A}$

$1.3 = \frac{1}{50} \times \frac{\ell}{A}$

$\frac{\ell}{A} = 65 \text{m}^{-1}$

$\Lambda = \frac{\kappa \times 1000}{\text{molarity}}$

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

$$= \frac{\left(\frac{1}{260} \times 65 \text{ m}^{-1}\right) \times 1000 \text{ cm}^3}{0.4 \text{ moles}}$$

$$= \frac{650 \text{ m}^{-1}}{260 \times 4 \text{ mol}} \times \frac{1}{1000} \text{ m}^3$$

$= 6.25 \times 10^{-4} \text{ S m}^2 \text{ mol}^{-1}$

145. (b) Λ° for CH_3COOH

$= \lambda^\circ_{CH_3COO^-} + \lambda^\circ_{Na^+} + \lambda^\circ_{H^+} + \lambda^\circ_{Cl^-} - \lambda^\circ_{Na^+} - \lambda^\circ_{Cl^-}$

$= \lambda^\circ_{CH_3COO^-} + \lambda^\circ_{H^+}$

$= 90 + 425 - 125 = 390 \text{ mho cm}^2 \text{ mol}^{-1}$.

Degree of dissociation $(\alpha) = \frac{\Lambda^c_m}{\Lambda^\circ_m} = \frac{7.8}{390} = 0.02$

146. (d) $\Lambda_C = \Lambda_0 \times \alpha = 8$

$\Lambda_C = \frac{1}{R} \times \frac{\ell}{A} \times \frac{1000}{N}$

$R = \frac{0.4 \times 1000}{8 \times 0.1} = 500$ Ohms

147. (b)

148. (d) Charge on Mg and Ca ion is greater than that of Na and K, so Mg and Ca ions possess higher conductivity, also solvation of metal ion decreases as we move down the group, hence conductivity increases)

149. (d) The conductivity of electrolytic solution depends upon all of the given factors.

150. (c) Correct Nernst equation is

$$E = E^o + \frac{2.303 RT}{nF} \log_a M^{n+}.$$

151. (a) $2H^+ + 2e^- \longrightarrow H_2$

E_H (Eq. wt) $= \frac{2}{2} = 1 \text{ g} = \frac{22400}{2} = 11200$ ml (STP)

Total charge passed $= \frac{96500 \times 112}{11200} = 965$

$Q = It = 965$

$I = \frac{965}{965} = 1$ amp.

152. (c) Using Faraday's second law of electrolysis,

$$\frac{\text{Weight of Cu deposited}}{\text{Weight of Ag deposited}} = \frac{\text{Equ. wt. of Cu}}{\text{Equ. wt. of Ag}}$$

$\Rightarrow \frac{w_{Cu}}{0.18} = \frac{63.5}{2} \times \frac{1}{108}$

$\Rightarrow w_{Cu} = \frac{63.5 \times 18}{2 \times 108 \times 100} = 0.0529$ g.

153. (c) According to the definition 1 F or 96500 C is the charge carried by 1 mol of electrons when water is electrolysed
$2H_2O \longrightarrow 4H^+ + O_2 + 4e^-$
So, 4 Faraday of electricity liberate = 32 g of O_2.
Thus 1 Faraday of electricity liberate
$= \frac{32}{4}$ g of $O_2 = 8$ g of O_2

154. (b) $H_2O \rightleftharpoons H^+ + OH^-$

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

$\therefore$ 0.5 mole of H_2 is liberated by 1 F = 96500 C
0.01 mole of H_2 will be liberated by

$= \frac{96500}{0.5} \times 0.01 = 1930$ C

$Q = I \times t$

$t = \frac{Q}{I} = \frac{1930 \text{ C}}{10 \times 10^{-3} \text{ A}} = 19.3 \times 10^4$ sec

155. (b) At Anode, $Cl^- \rightarrow \frac{1}{2}Cl_2 + e^-$

Equivalent wt. of chlorine $(E_{Cl_2}) = \frac{35.5 \times 2}{2} = 35.5$

$$W_{Cl_2} = \frac{E_{Cl_2} \times I \times t}{96500} = \frac{35.5 \times 2 \times 30 \times 60}{96500} = 1.32 \text{ gm.}$$

156. (a) By Faraday's 1st Law of electrolysis, $\frac{W}{E} = \frac{Q}{96500}$

(where $Q = it$ = charge of ion)

We know that no. of gram equivalent

$$= \frac{W}{E} = \frac{it}{96500} = \frac{1 \times 965}{96500} = \frac{1}{100}$$

(where $i = 1$ A, $t = 16 \times 60 + 5 = 965$ sec.)

Since, we know that

$$\text{Normality} = \frac{\text{No. of gram equivalent}}{\text{Volume (in litre)}} = \frac{\frac{1}{100}}{1} = 0.01 \text{ N}$$

157. (b) Quantity of charge passed $= 0.5 \times 30 \times 60 = 900$ coulomb

900 coulomb will deposit = 0.2964g of copper

$\therefore$ 96500 coulomb will deposit

$$= \frac{0.2964}{900} \times 96500 = 31.75 \text{ g of copper}$$

Thus, 31.75 is the eq. mass of copper

At. mass = Eq. mass × Valency

$63.56 = 31.75 \times x$

$x = +2.$

158. (a) The charge carried by 1 mole of electrons is one faraday. Thus for a reaction

$$M^{n+} + ne^- \longrightarrow M$$

nF = 1 mole of M

$$\underset{\substack{3F\\1F}}{Al^{3+}} + 3e^- \longrightarrow \underset{\substack{1\text{ mol}\\1/3\text{ mol}}}{Al}$$

$$\underset{\substack{2F\\1F}}{Cu^{2+}} + 2e^- \longrightarrow \underset{\substack{1\text{ mol}\\1/2\text{ mol}}}{Cu}$$

$$\underset{1F}{Na^{+}} + e^- \longrightarrow \underset{1\text{ mol}}{Na}$$

The mole ratio of Al, Cu and Na deposited at the respective cathode is $\frac{1}{3} : \frac{1}{2} : 1$ or $2 : 3 : 6$.

159. (a)

160. (b) Total of 6 electrons are required to form 2 moles of Cr^{3+} therefore to form 1 mole 3F of charge is required.

161. (d) $Pb + PbO_2 + 2H_2SO_4 \underset{\text{Recharge}}{\overset{\text{Discharge}}{\rightleftharpoons}} 2PbSO_4 + 2H_2O.$

Sulphuric acid is consumed on discharging.

162. (d) This is because zinc has higher oxidation potential than Ni, Cu and Sn. The process of coating of iron surface with zinc is known as galvanization. Galvanized iron sheets maintain their lustre due to the formation of protective layer of basic zinc carbonate.

163. (c) At cathode reduction occurs according to following reaction.

$$\overset{+4}{Mn}O_2 + NH_4^+ + e^- \longrightarrow \overset{+3}{Mn}O(OH) + NH_3$$