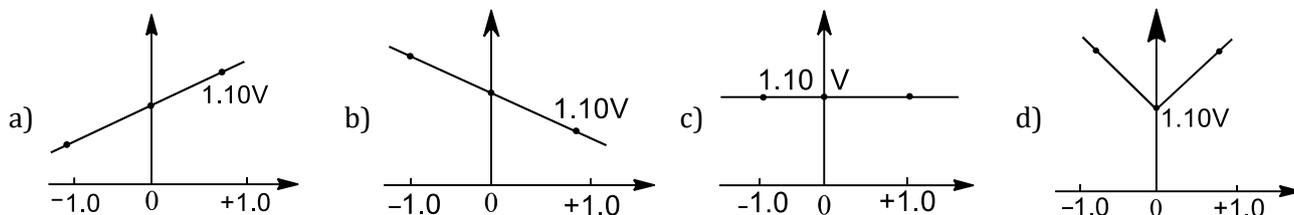


10. Molten NaCl conducts electricity due to the presence of :
 a) Free electrons b) Free molecules c) Free ions d) Atoms of Na and Cl
11. The emf of the cell, ($E_{\text{Zn}^{2+}/\text{Zn}} = -0.76 \text{ V}$)
 $\text{Zn} / \text{Zn}^{2+} (1 \text{ M}) \parallel \text{Cu}^{2+} (1 \text{ M}) | \text{Cu}$
 ($E_{\text{Cu}^{2+}/\text{Cu}} = +0.34 \text{ V}$) will be
 a) +1.10 V b) -1.10 V c) +0.42 V d) -0.42 V
12. Which represents a concentration cell?
 a) $\text{PtH}_2 | \text{HCl} (c_1) || \text{HCl} (c_2) | \text{PtH}_2$ b) $\text{PtH}_2 | \text{HCl} (c_1) || \text{Cl}_2 | \text{Pt}$ c) $\text{Zn} | \text{Zn}^{2+} || \text{Cu}^{2+} | \text{Cu}$ d) $\text{Fe} | \text{Fe}^{2+} || \text{Cu}^{2+} | \text{Cu}$
13. In electrolysis of aqueous copper sulphate, the gas at anode and cathode are
 a) O_2 and H_2 b) H_2 and O_2 c) SO_2 and H_2 d) SO_3 and O_2
14. Consider the reaction, $M^{n+}(aq) + ne \rightarrow M^0(s)$. The standard reduction potential values of the metals M_1, M_2 and M_3 are $-0.34 \text{ V}, -3.05 \text{ V}$ and -1.66 V respectively. The order of their reducing power will be :
 a) $M_1 > M_2 > M_3$ b) $M_3 > M_2 > M_1$ c) $M_1 > M_3 > M_2$ d) $M_2 > M_3 > M_1$
15. The charge required to liberate one gram equivalent of an element is
 a) 96500 F b) 1 F c) 1 C d) None of these
16. What will be pH of aqueous solution of electrolyte in electrolytic cell during electrolysis of $\text{CuSO}_4(aq)$ between graphite electrodes?
 a) pH = 14.0 b) pH > 7.0 c) pH < 7.0 d) pH = 7.0
17. In an electrolytic cell, the anode and cathode are respectively represented as :
 a) Positive electrode, negative electrode
 b) Negative electrode, positive electrode
 c) Both positive and negative electrode
 d) None of the above
18. The cell reaction is spontaneous, when
 a) E_{red}° is negative b) E_{red}° is positive c) ΔG° is negative d) ΔG° is positive
19. The emf of the cell $\text{Mg} | \text{Mg}^{2+} (0.01 \text{ M}) || \text{Sn}^{2+} (0.1 \text{ M}) | \text{Sn}$ at 298 K is (Given, $E_{\text{Mg}^{2+}, \text{Mg}}^\circ = -2.34 \text{ V}, -2.34 \text{ V}, E_{\text{Sn}^{2+}, \text{Sn}}^\circ = -0.14 \text{ V}$)
 a) 2.23 V b) 1.86 V c) 1.56 V d) 3.26 V
20. When an aqueous solution of lithium chloride is electrolysed using graphite electrodes :
 a) pH of the resulting solution increases
 b) pH of the resulting solution decreases
 c) As the current flows, pH of the solution around the cathode increases
 d) None of the above
21. In electrolytic purification, which of the following is made of impure metal?
 a) Anode b) Cathode c) Both (a) and (b) d) None of these
22. The specific conductivity of 0.1 N KCl solution is $0.0129 \Omega^{-1} \text{cm}^{-1}$. The resistance of the solution in the cell 100Ω . The cell constant of the cell will be
 a) 1.10 b) 1.29 c) 0.56 d) 2.80
23. Which graph correctly correlates E_{Cell} as a function of concentrations for the cell (for different values of M and M')?
 $\text{Zn}(s) + \text{Cu}^{2+}(M) \rightarrow \text{Zn}^{2+}(M') + \text{Cu}(s);$
 $E_{\text{Cell}}^\circ = 1.10 \text{ V}$
 $X - \text{axis} : \log_{10} \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}, Y - \text{axis} : E_{\text{Cell}}$



24. In acidic medium MnO_4^- is converted to Mn^{2+} . The quantity of electricity in faraday required to reduce 0.5 mole of MnO_4^- to Mn^{2+} would be
 a) 2.5 b) 5 c) 1 d) 0.5
25. In electrolysis, oxidation takes place at:
 a) Anode
 b) Cathode
 c) The anode as well as cathode
 d) The surface of electrolyte solution
26. A depolariser used in dry cell batteries is :
 a) Ammonium chloride b) Manganese dioxide c) Potassium hydroxide d) Sodium phosphate
27. The $E^\circ_{M^{3+}/M^{2+}}$ values for Cr, Mn, Fe and Co are $-0.41, +1.57, +0.77$ and $+1.97$ V respectively. For which one of these metals, the change in oxidation state from +2 to +3 is easiest?
 a) Fe b) Mn c) Co d) Cr
28. The standard reduction electrode potential values of the elements A, B and C are $+0.68, -2.50$ and -0.50 V respectively. The order of their reducing power is :
 a) $A > B > C$ b) $A > C > B$ c) $C > B > A$ d) $B > C > A$
29. The number of electrons involved in the reaction when a faraday of electricity is passed through an electrolyte in solution is :
 a) 12×10^{46} b) 96500 c) 8×10^{16} d) 6.02×10^{23}
30. The electrolysis of a solution resulted in the formation of H_2 at the cathode and Cl_2 at the anode. The liquid is:
 a) Pure water
 b) H_2SO_4 solution
 c) NaCl solution in water
 d) CuCl_2 solution in water
31. The passage of electricity in the Daniell cell when Zn and Cu electrodes are connected:
 a) From Cu to Zn inside the cell
 b) From Cu to Zn outside the cell
 c) From Zn to Cu outside the cell
 d) None of the above
32. $\text{Ni} / \text{Ni}^{2+} [1.0 \text{ M}] || \text{Au}^{3+} [1.0 \text{ M}] / \text{Au}$ where E° for $\text{Ni}^{2+} / \text{Ni}$ is -0.250 V; and E° for $\text{Au}^{3+} / \text{Au}$ is 0.150 V. The emf of the cell is
 a) $+1.25$ V b) -1.75 V c) $+1.75$ V d) $+0.4$ V
33. The product obtained at anode when 50% H_2SO_4 aqueous solution is electrolysed using platinum electrodes is
 a) H_2SO_3 b) $\text{H}_2\text{S}_2\text{O}_8$ c) O_2 d) H_2
34. The approximate e.m.f. of a dry cell is :
 a) 2.0 V b) 1.2 V c) 6 V d) 1.5 V
35. $E_1, E_2,$ and E_3 are the emfs of the following three galvanic cells respectively
 I. $\text{Zn} (\text{s}) | \text{Zn}^{2+} (0.1 \text{ M}) || \text{Cu}^{2+} (1 \text{ M}) | \text{Cu} (\text{s})$
 II. $\text{Zn} (\text{s}) | \text{Zn}^{2+} (1 \text{ M}) || \text{Cu}^{2+} (1 \text{ M}) | \text{Cu} (\text{s})$
 III. $\text{Zn} (\text{s}) | \text{Zn}^{2+} (1 \text{ M}) || \text{Cu}^{2+} (0.1 \text{ M}) | \text{Cu} (\text{s})$
 Which one of the following is true?

- a) $E_2 > E_1 > E_3$ b) $E_1 > E_2 > E_3$ c) $E_3 > E_1 > E_2$ d) $E_3 > E_2 > E_1$
36. The fraction of the total current carried by an ion is known as:
 a) Transport number of that ion
 b) Conductance of that ion
 c) Both(a) and (b)
 d) None of the above
37. In a galvanic cell, which is wrong?
 a) Anode has negative polarity
 b) Cathode has positive polarity
 c) Reduction takes place at anode
 d) Reduction takes place at cathode
38. The rusting of iron takes place as follows

$$2\text{H}^+ + 2\text{e}^- + \frac{1}{2}\text{O}_2 \rightarrow \text{H}_2\text{O}(\text{l});$$

$$E^\circ = +1.23\text{ V}$$

$$\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}(\text{s}); \quad E^\circ = -0.44\text{ V}$$
 Calculate ΔG° for the net process.
 a) -322 kJ mol^{-1} b) -161 kJ mol^{-1} c) -152 kJ mol^{-1} d) -76 kJ mol^{-1}
39. What weight of copper will be deposited by passing 2 faraday of electricity through a solution of Cu(II) salt?
 a) 35.6 g b) 63.5 g c) 6.35 g d) 3.56 g
40. Chlorine cannot displace :
 a) Fluorine from NaF b) Iodine from NaI c) Bromine from NaBr d) None of these
41. For Acell reaction involving Atwo-electron change, the standard emf of the cell is found to be 0.295 V at 25°C. The equilibrium constant of the reaction at 25 °C will be
 a) 1×10^{-10} b) 29.5×10^{-2} c) 10 d) 1×10^{10}
42. The resistance of a decinormal solution of a salt occupying a volume between two platinum electrodes 1.80 cm apart and 5.4 cm² in area was formed to be 32 ohm. The specific and equivalent conductivity respectively in their proper units are :
 a) 104.1 and 0.0104 b) 208.2 and 0.0208 c) 0.0104 and 104.0 d) None of these
43. The value of equilibrium constant for a feasible cell reaction is :
 a) < 1 b) Zero c) = 1 d) > 1
44. At 25°C, the standard e.m.f. of cell having reactions involving a two electron change is found to be 0.295 V. The equilibrium constant of the reaction is :
 a) 29.5×10^{-2} b) 10 c) 10^{10} d) 29.5×10^{10}
45. E° for $\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}$ is -0.44 V and E° for $\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$ is -0.76 V thus
 a) Zn is more electropositive than Fe b) Zn is more electronegative than Fe
 c) Fe is more electropositive than Zn d) None of the above
46. A certain quantity of electricity is passed through aqueous solution of AgNO₃ and CuSO₄ connected in series. If Ag (at.wt.108) deposited at the cathode is 1.08 g then Cu deposited at the cathode is (at. wt. of Cu is 63.53):
 a) 6.354 g b) 0.317 g c) 0.6354 g d) 3.177 g
47. $\text{I}_2(\text{s}) | \text{I}^- (0.1\text{ M})$ half-cell is connected to a $\text{H}^+(\text{aq}) | \text{H}_2(1\text{ bar}) | \text{Pt}$ half-cell and emf is found to be 0.7714 V. If $E_{\text{I}_2/\text{I}^-}^\circ = 0.535\text{ V}$, find the pH of H^+/H_2 half-cell
 a) 1 b) 2 c) 3 d) 5
48. The $E_{\text{M}^{3+}/\text{M}^{2+}}^\circ$ values for Cr, Mn, Fe and Co are -0.41 V, +1.57 V, +0.77 V and +1.97 V respectively. For which one of these metals the change in oxidation state from +2 to +3 is easiest?
 a) Cr b) Mn c) Fe d) Co
49. In which cell, liquid function potential need to be eliminated?

- a) $\text{Pt}/\text{H}_2(\text{P}_1)|\text{HCl}|\text{Pt}/\text{H}_2(\text{P}_2)$
 b) $\text{Pt}/\text{H}_2|_{c_1}\text{HCl}|_{c_2}\text{Pt}/\text{H}_2$
 c) Nicad cell
 d) Lead storage battery
50. Which one of the following nitrates will leave behind Ametal on strong heating?
 a) Ferric nitrate b) Copper nitrate c) Manganese nitrate d) Silver nitrate
51. $E_{\text{Cu}}^\circ = 0.34 \text{ V}$, $E_{\text{Zn}}^\circ = 0.76 \text{ V}$. A Daniel cell contains 0.1 M ZnSO_4 solution and 0.01 M CuSO_4 solution at its electrodes. EMF of the cell is
 a) 1.10 V b) 1.04 V c) 1.16 V d) 1.07 V
52. The E° of Fe^{2+}/Fe and Sn^{2+}/Sn are -0.44 V and -0.14 V respectively. If cell reaction is $\text{Fe} + \text{Sn}^{2+} \rightarrow \text{Fe}^{2+} + \text{Sn}$ then emf of the cell is
 a) +0.30 V b) -0.58 V c) +0.58 V d) -0.30 V
53. Electrolysis rules of Faraday's states that mass deposited on electrode is proportional to
 a) Q b) Q^2 c) I^2 d) None of these
54. A silver cup is plated with silver by passing 965 C of electricity. The amount of Ag deposited is
 a) 107.89 g b) 9.89 g c) 1.0002 g d) 1.08 g
55. The molecular conductivity and equivalent conductivity are same for the solution of :
 a) 1 M NaCl b) 1 M $\text{Ba}(\text{NO}_3)_2$ c) 1 M $\text{La}(\text{NO}_3)_3$ d) 1 M $\text{Th}(\text{NO}_3)_4$
56. Dipping iron article into a strongly alkaline solution of sodium phosphate
 a) Does not affect the article b) Forms $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$ on the surface
 c) Forms iron phosphate film d) Forms ferric hydroxide
57. When an electric current is passed through an aqueous solution of sodium chloride :
 a) H_2 is evolved at the anode
 b) Oxygen is evolved at the cathode
 c) Its pH progressively decreases
 d) Its pH progressively increases
58. The cell reaction of the galvanic cell $\text{Cu}(s) | \text{Cu}^{2+}(aq) || \text{Hg}^{2+}(aq) | \text{Hg}(l)$ is
 a) $\text{Hg} + \text{Cu}^{2+} \rightarrow \text{Hg}^{2+} + \text{Cu}$ b) $\text{Hg} + \text{Cu}^{2+} \rightarrow \text{Cu}^+ + \text{Hg}^+$
 c) $\text{Cu} + \text{Hg} \rightarrow \text{CuHg}$ d) $\text{Cu} + \text{Hg}^{2+} \rightarrow \text{Cu}^{2+} + \text{Hg}$
59. Calculate the volume of hydrogen at NTP obtained by passing a current of 0.4 ampere through acidified water for 30 minute :
 a) 0.0836 litre b) 0.1672 litre c) 0.0432 litre d) 0.836 litre
60. The standard emf of a cell involving one electron change is found to be 0.591 V and 25 °C. The equilibrium constant of the reaction is ($F = 96500 \text{ C mol}^{-1}$)
 a) 1.0×10^1 b) 1.0×10^5 c) 1.0×10^{10} d) 1.0×10^{30}
61. The relationship between Gibbs' free energy change (ΔG) and emf (E) of a reversible electrochemical cell is given by
 a) $\Delta G = nFE$ b) $\Delta G = nF/E$ c) $\Delta G = -nFE$ d) $\Delta G = E/nF$
62. The reduction electrode potential, E of 0.1 M solution of M^+ ions ($E_{\text{RP}} = -2.36 \text{ V}$) is
 a) -4.82 V b) -2.41 V c) +2.41 V d) None of these
63. Passage of 1 faraday of electricity through a solution of CuSO_4 , deposits :
 a) 1 mole of Cu b) 1 g-atom of Cu c) 1 molecule of Cu d) 1 g equivalent of Cu
64. The conductivity of $N/50$ solution of KCl in a cell at 25°C is $0.002765 \text{ mho cm}^{-1}$. If the resistance of a cell containing this solution is 400 ohm, the cell constant is :
 a) 1.106 cm b) 1.106 cm^{-1} c) 1 cm d) 1 cm^{-1}
65. The equilibrium constant for the reaction given below at 298 K is :



$$E_{\text{cell}}^{\circ} = 2905 \text{ V at } 298 \text{ K}$$

$$\text{a) } e^{0.32/0.0295}$$

$$\text{b) } 10^{0.595/0.76}$$

$$\text{c) } 10^{0.0250/0.32}$$

$$\text{d) } 10^{0.32/0.295}$$

66. When the sample of copper with zinc impurity is to be purified by electrolysis, the appropriate electrodes are
- | Cathode | Anode |
|------------------|---------------|
| a) Pure zinc | pure copper |
| b) Impure sample | pure copper |
| c) Impure zinc | impure sample |
| d) Pure copper | impure sample |
67. A current of 12 A is passed through an electrolytic cell containing aqueous NiSO_4 solution. Both Ni and H_2 gas are formed at the cathode. The current efficiency is 60%. What is the mass of nickel deposited on the cathode per hour?
- a) 7.883 g b) 3.941 g c) 5.91 g d) 2.645 g
68. 10^{-2} g atom of Ag can be oxidised to Ag^+ during the electrolysis of AgNO_3 solution using silver electrode by :
- a) 965 coulomb b) 96500 coulomb c) 9650 coulomb d) 96.500 coulomb
69. A gas X at 1 atm is bubbled through a solution containing a mixture of 1 M y^- and 1 M z^- at 25°C . If the order of reduction potential is $z > y > x$ then
- a) y will oxidize x and not z b) y will oxidize x and z
- c) y will oxidize z and not x d) y will reduce both x and z
70. Which one of the following solutions will have highest conductivity?
- a) 0.1 M CH_3COOH b) 0.1 M NaCl c) 0.1 M KNO_3 d) 0.1 M HCl
71. A current of strength 2.5 A was passed through CuSO_4 solution for 6 min 26 s. The amount of copper deposited is (At. Wt. of Cu = 63.5, $1F = 96500 \text{ C}$)
- a) 0.3175 g b) 3.175 g c) 0.635 g d) 6.35 g
72. A student made the following observations in the laboratory,
- i) Clean copper metal did not react with 1 molar $\text{Pb}(\text{NO}_3)_2$ solution.
- ii) Clean lead metal dissolved in a 1 molar AgNO_3 solution and crystals of Ag metal appeared.
- iii) Clean silver metal did not react with 1 molar $\text{Cu}(\text{NO}_3)_2$ solution.
- The order of decreasing reducing character of the three metals is :
- a) Cu, Pb, Ag b) Cu, Ag, Pb c) Pb, Cu, Ag d) Pb, Ag, Cu
73. The e. m. f. of the cell $\text{Zn} | \text{Zn}^{2+} (1 \text{ M}) || \text{Cu}^{2+} | \text{Cu} (1 \text{ M})$ is 1.1 volt. If the standard reduction potential of $\text{Zn}^{2+} | \text{Zn}$ is -0.78 volt, what is the oxidation potential of $\text{Cu} | \text{Cu}^{2+}$?
- a) + 1.86 V b) 0.32 V c) -0.32 V d) -1.86 V
74. Standard reduction electrode potentials of three metals A, B and C are respectively + 0.5 V, -3.0 V and -1.2 V . The reducing powers of these metals are
- a) $A > B > C$ b) $C > B > A$ c) $A > C > B$ d) $B > C > A$
75. Quantity of charge is measured in :
- a) ampere-sec. b) ampere c) ampere sec^{-1} . d) ampere $^{-1}$ sec.
76. Which of the following will form a cell with the highest voltage?
- a) 0.1 M Ag^+ , 2 M Co^{2+} b) 2 M Ag^+ , 2 M Co^{2+} c) 1 M Ag^+ , 1 M Co^{2+} d) 2 M Ag^+ , 0.1 M Co^{2+}
77. When electric current is passed through acidified water for 1930 s, 1120 mL of H_2 gas is collected (at STP) at the cathode. What is the current passed in amperes?
- a) 0.05 b) 0.50 c) 5.0 d) 50
78. In which of the following pairs, the constants/ quantities are not mathematically related to each other?
- a) Gibbs free energy and standard cell potential
- b) Equilibrium constant and standard cell potential
- c) Rate constant and activation energy
- d) Rate constant and standard cell potential

79. The charge required for reduction of 1 mole of $\text{Cr}_2\text{O}_7^{2-}$ ions to Cr^{3+} is
 a) 96500 C b) 2×96500 C c) 3×96500 C d) 6×96500 C
80. Cell constant has the unit:
 a) cm b) cm^{-1} c) cm^2 d) cm sec^{-1}
81. The resistance of 0.01 N solution of an electrolyte was found to be 210 ohm at 298 K, using a conductivity cell of cell constant 0.66 cm^{-1} . The equivalent conductivity of solution is :
 a) $314.28 \text{ mho cm}^2 \text{ eq.}^{-1}$ b) $3.14 \text{ mho cm}^2 \text{ eq.}^{-1}$ c) $314.28 \text{ mho}^{-1} \text{ cm}^2 \text{ eq.}^{-1}$ d) $3.14 \text{ mho}^{-1} \text{ cm}^2 \text{ eq.}^{-1}$
82. IV. $\text{Cu} + 2\text{HCl} \rightarrow \text{CuCl}_2 + \text{H}_2(\text{g})$
 $[E_{\text{Cu}^{2+}/\text{Cu}}^\circ = +0.34 \text{ V}]$
 V. $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2(\text{g})$
 $[E_{\text{Zn}^{2+}/\text{Zn}}^\circ = -0.76 \text{ V}]$
 VI. $\text{Ag} + 2\text{HCl} \rightarrow \text{AgCl} + \frac{1}{2}\text{H}_2(\text{g})$
 $[E_{\text{Ag}^+/\text{Ag}}^\circ = +0.80 \text{ V}]$
- Which of the following reaction is feasible ?
 a) (ii) b) (i) c) (iii) d) All of these
83. The standard potential at 25°C for the following half-reactions are given against them $\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}$, $E^\circ = -0.762 \text{ V}$
 $\text{Mg}^{2+} + 2e^- \rightarrow \text{Mg}$, $E^\circ = -2.37 \text{ V}$
 When zinc dust is added to the solution of MgCl_2 ,
 a) ZnCl_2 is formed b) Mg is precipitated
 c) Zn dissolves in the solution d) No reaction takes place
84. 1 coulomb of charge passes through solution of AgNO_3 and CuSO_4 connected in series and the concentration of two solution being in the ratio 1 : 2. The ratio of amount of Ag and Cu deposited on Pt electrode is :
 a) 107.9 : 63.54 b) 54 : 31.77 c) 107.9 : 31.77 d) 54 : 63.54
85. When lead accumulator is charged, it is :
 a) An electrolytic cell b) A galvanic cell c) A Daniell cell d) None of these
86. If the ΔG of a cell reaction $\text{AgCl} + e^- \rightarrow \text{Ag} + \text{Cl}^-$ is -21.20 kJ , the standard emf of cell is
 a) 0.239 V b) 0.220 V c) -0.320 V d) -0.110 V
87. $\Lambda_{\text{ClCH}_2\text{COONa}}^\infty = 224 \Omega^{-1} \text{ cm}^2 \text{ g equiv}^{-1}$
 $\Lambda_{\text{NaCl}}^\infty = 38.5 \Omega^{-1} \text{ cm}^2 \text{ g equiv}^{-1}$
 $\Lambda_{\text{HCl}}^\infty = 203 \Omega^{-1} \text{ cm}^2 \text{ g equiv}^{-1}$
 What is the value of $\lambda_{\text{ClCH}_2\text{COOH}} = ?$
 a) $288.5 \Omega^{-1} \text{ cm}^2 \text{ g equiv}^{-1}$ b) $289.5 \Omega^{-1} \text{ cm}^2 \text{ g equiv}^{-1}$
 c) $388.5 \Omega^{-1} \text{ cm}^2 \text{ g equiv}^{-1}$ d) $59.5 \Omega^{-1} \text{ cm}^2 \text{ g equiv}^{-1}$
88. When a copper wire is immersed in a solution of AgNO_3 , the colour of the solution becomes blue because copper :
 a) Forms a soluble complex with AgNO_3
 b) Is oxidised to Cu^{2+}
 c) Is reduced to Cu^{2-}
 d) Splits up into atomic form and dissolves
89. The electrode potential of a glass electrode depends upon :
 a) Concentration of chloride ions
 b) Concentration of hydrogen ions
 c) Concentration of KCl solution
 d) None of the above
90. 0.04 N solution of a weak acid has conductivity $4.23 \times 10^{-4} \text{ mho cm}^{-1}$. If the degree of dissociation of acid at this dilution is 0.0612, then equivalent conductivity at infinite dilution is $\text{mho cm}^2 \text{ eq.}^{-1}$:
 a) 172.8 b) 180 c) 190 d) 160

104. A galvanic cell with electrode potential of 'A' = +2.23 V and 'B' = -1.43 V. The value of E_{cell}° is
 a) 3.66 V b) 0.80 V c) -0.80 V d) -3.66 V
105. A galvanic cell is composed of two hydrogen electrodes, one of which is a standard one. In which of the following solutions should the other electrode be immersed to get maximum e. m. f.?
 a) 0.1 M HCl b) 0.1 M CH₃COOH c) 0.1 M H₃PO₄ d) 0.1 M H₂SO₄
106. Which metal does not give the following reaction?
 $M + \text{water or steam} \rightarrow \text{oxide} + \text{H}_2 \uparrow$
 a) Iron b) Sodium c) Mercury d) Magnesium
107. 4.5 g of Al (at. mass 27 amu) is deposited at cathode from Al³⁺ solution by a certain quantity of charge. The volume of H₂ produced at STP from H⁺ ions in solution by the same quantity of charge will be :
 a) 11.2 L b) 44.8 L c) 5.6 L d) 22.4 L
108. In the electrolysis of acidulated water, it is desired to obtain 1.12 cc of hydrogen per second under STP condition. The current to be passed is
 a) 1.93 A b) 9.65 A c) 19.3 A d) 0.965 A
109. The speed of ions during passage of current depends upon :
 a) Nature of ion b) Potential gradient c) Dilution of solution d) All of these
110. The best way to prevent rusting of iron is
 a) Making it cathode b) Putting in saline water
 c) Both (a) and (b) d) None of these
111. The hydrogen electrode is dipped in a solution of pH = 3 at 25°C. The reduction potential of the cell would be :
 a) 0.177 V b) -0.177 V c) 0.087 V d) 0.059 V
112. Conductivity (unit Siemen) is directly proportional to area of the vessel and the concentration of the solution in it and is inversely proportional to the length of the vessel. Then, the units of the constant of proportionality is
 a) S² m² mol b) S² m² mol⁻² c) S m² mol⁻¹ d) S m mol⁻¹
113. The metal that cannot be produced on reduction of its oxide by aluminium is :
 a) K b) Mn c) Cr d) Fe
114. In the concentration cells, the electrical energy is produced due to :
 a) Oxidation of fuel
 b) Heat energy
 c) Chemical reaction
 d) Transfer of a substance from one concentration to other
115. How many faraday are needed to reduce a mole of MnO₄⁻ of Mn²⁺?
 a) 4 b) 5 c) 3 d) 2
116. For the cell,
 T1 | T1⁺ (0.001 M) || Cu²⁺ (0.1 M) | Cu
 E_{cell} at 25°C is 0.83 V. E_{cell} can be increased
 a) By decreasing [Cu²⁺] b) By increasing [Cu²⁺]
 c) By increasing [T1⁺] d) None of these
117. In an aqueous solution, hydrogen (H₂) will not reduce :
 a) Fe³⁺ b) Cu²⁺ c) Zn²⁺ d) Ag⁺
118. How many faradays of electricity are required to electrolyse 1 mole CuCl₂ to copper metal and chlorine gas?
 a) 1 F b) 2 F c) 3 F d) 4 F
119. Which statement is not correct?
 a) Conductance of an electrolytic solution increases with dilution
 b) Conductance of an electrolytic solution decreases with dilution
 c) Specific conductance of an electrolytic solution decreases with dilution
 d) Equivalent conductance of an electrolytic solution increase with dilution

120. The correct value of e.m.f. of cell is given by :

i) $E_{\text{cell}} = E_{\text{OP anode}} - E_{\text{RP cathode}}$

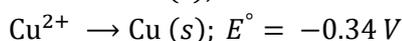
ii) $E_{\text{cell}} = E_{\text{OP anode}} + E_{\text{RP cathode}}$

iii) $E_{\text{cell}} = E_{\text{RP anode}} + E_{\text{RP cathode}}$

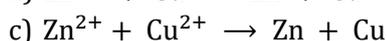
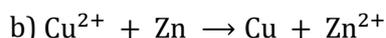
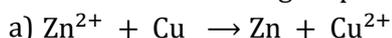
iv) $E_{\text{cell}} = E_{\text{OP anode}} - E_{\text{OP cathode}}$

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

121. $\text{Zn}^{2+} \rightarrow \text{Zn (s)}; E^\circ = -0.76 \text{ V}$



Which of the following is spontaneous?



d) None of the above

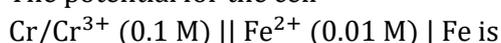
122. Reduction potentials of *A, B, C*, and *D* are 0.8 V, 0.79 V, 0.34 V and -2.37 V respectively. Which element displaces all the other three elements?

- a) *B* b) *A* c) *D* d) *C*

123. Given,

$E^\circ_{\text{Cr}^{3+}/\text{Cr}} = 0.72 \text{ V}, E^\circ_{\text{Fe}^{2+}/\text{Fe}} = 0.42 \text{ V}.$

The potential for the cell



- a) 0.26 V b) 0.399 V c) -0.339 V d) -0.26 V

124. The electroplating with chromium is undertaken because :

- a) Electrolysis of chromium is easier
b) Chromium can form alloys with other metals
c) Chromium gives a protective and decorative coating to the base metal
d) Of high reactivity of chromium metal

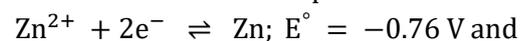
125. Which of the following is not correct?

- a) Aqueous solution of NaCl is an electrolyte.
b) The units of electrochemical equivalent are g-coulomb.
c) In the Nernst equation, *n* represents the number of electrons transferred in the electrode reaction.
d) Standard reduction potential of hydrogen electrode is zero volt.

126. H_2 cannot be displaced by

- a) Li^+ b) Sr^{2+} c) Al^{3+} d) Ag^+

127. The standard reduction potential of Zn and Ag in water at 298 K are,



128. The amount of an ion discharged during electrolysis is not dependent of :

- a) Resistance of solution
b) Time
c) Current strength
d) Electrochemical equivalent of the element

129. The conductivity of a 0.1 N KCl solution at 23°C is 0.012 ohm⁻¹ cm⁻¹. The resistance of the cell containing the solution at the same temperature was found to be 55 ohm. The cell constant will be :

- a) 0.918 cm⁻¹ b) 0.66 cm⁻¹ c) 1.142 cm⁻¹ d) 1.12 cm⁻¹

130. Reduction potential of four elements *P, Q, R, S* is -2.90, +0.34, +1.20 and -0.76. Reactivity decreases in the order

- a) $P > Q > R > S$ b) $S > R > Q > P$ c) $P > S > Q > R$ d) $Q > S > R > P$

131. Which of the following statements are correct concerning redox properties?

I A metal *M* for which E° for the half reaction

$M^{n+} + ne^{-} = M$, is very negative will be A good reducing agent.

II The oxidizing power of the halogens decreases from chlorine to iodine.

III The reducing power of hydrogen halides increases from hydrogen chloride to hydrogen iodide.

a) I, II and III b) I and II c) I only d) II and III only

132. A cell with two electrodes, one of grey tin and the other white tin, both dipping in solution of $(\text{NH}_4)_2\text{SnCl}_6$ showed zero e.m.f. at 18°C . What conclusion may be drawn from this?

- a) The e.m.f. developed at the electrode-solution phase boundary cancels the normal e.m.f.
b) Grey tin being non-metallic ceases to provide a reversible electrode reaction
c) Electrode surface develops a protective layer and the cell develops a very large internal resistance
d) The standard Gibbs energy change of the cell becomes zero

133. Aluminium displaces hydrogen from dilute HCl whereas silver does not. The emf of a cell prepared by combining Al/Al^{3+} and Ag/Ag^+ is 2.46 V. The reduction potential of silver electrode is +0.80 V. The reduction potential of aluminium electrode is

a) +1.66 V b) -3.26 V c) 3.26 V d) -1.66 V

134. For $\text{I}_2 + 2e^{-} \rightarrow 2\text{I}^{-}$, standard reduction potential = + 0.54 volt. For $2\text{Br}^{-} \rightarrow \text{Br}_2 + 2e^{-}$, standard oxidation potential = - 1.09 volt. For $\text{Fe} \rightarrow \text{Fe}^{2+} + 2e^{-}$, standard oxidation potential = + 0.44 volt. Which of the following reactions is non-spontaneous?

- a) $\text{Br}_2 + 2\text{I}^{-} \rightarrow 2\text{Br}^{-} + \text{I}_2$
b) $\text{Fe} + \text{Br}_2 \rightarrow \text{Fe}^{2+} + 2\text{Br}^{-}$
c) $\text{Fe} + \text{I}_2 \rightarrow \text{Fe}^{2+} + 2\text{I}^{-}$
d) $\text{I}_2 + 2\text{Br}^{-} \rightarrow 2\text{I}^{-} + \text{Br}_2$

135. When KMnO_4 acts as an oxidizing agent and ultimately forms MnO_4^{2-} , MnO_2 , Mn_2O_3 and Mn^{2+} then the number of electrons transferred in each case respectively, are

a) 4, 3, 1, 5 b) 1, 5, 3, 7 c) 1, 3, 4, 5 d) 3, 5, 7, 1

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

a) 10 b) 1×10^{10} c) 1×10^{-10} d) 10×10^{-2}

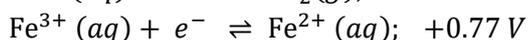
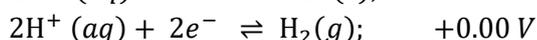
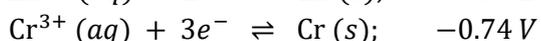
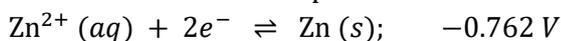
137. Which one of the following has the highest molar conductivity?

- a) Diaminedichloroplatinum (III) b) Tetraaminedichlorocobalt (III) chloride
c) Potassium hexacyanoferrate (II) d) Hexaquo chromium (III) bromide

138. Electrode potential of Zn^{2+}/Zn is - 0.76 V and that of Cu^{2+}/Cu is + 0.34 V. The emf of the cell constructed between these two electrodes is

a) 1.10 V b) - 1.10 V c) 2.20 V d) - 2.20 V

139. The standard reduction potentials at 298 K for the following half-cell reactions are given



Which one of the following is the strongest reducing agent?

a) $\text{Zn}(\text{s})$ b) $\text{Cr}(\text{s})$ c) $\text{H}_2(\text{g})$ d) $\text{Fe}^{2+}(\text{aq})$

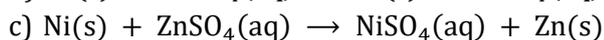
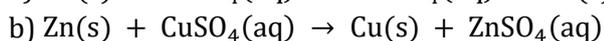
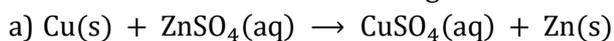
140. How long (in hours) must a current of 5.0 A be maintained to electroplate 60 g of calcium from molten CaCl_2 ?

a) 27 h b) 8.3 h c) 11 h d) 16 h

141. Use of electrolysis is

a) Electrorefining b) Electroplating c) Both (a) and (b) d) None of these

142. What is the cell reaction occurring in Daniel cell (Galvanic cell)?



156. If the current is passed into the solution of an electrolyte:
- Anions move towards anode, cations towards cathode
 - Anions and cations both move towards anode
 - Anions move towards cathode, cations towards anode
 - No movement of ions takes place
157. The element that is easiest to be reduced is :
- Fe
 - Cu
 - Ag
 - Sn
158. Standard reduction potential for, $\text{Li}^+|\text{Li}$, $\text{Zn}^{2+}|\text{Zn}$, $\text{H}^+|\text{H}_2$ and $\text{Ag}^+|\text{Ag}$ is $-3.05, -0.762, 0.00$ and $+0.80$ V. Which has highest reducing capacity?
- Ag
 - H_2
 - Zn
 - Li
159. What is the quantity of electricity (in Coulombs) required to deposit all the silver from 250mL of 1 M AgNO_3 solution?
- 2412.5
 - 24125
 - 4825.0
 - 48250
160. When 1 faraday of electricity is passed through CuSO_4 solution, number of atoms formed is :
- 6.02×10^{23}
 - 3.01×10^{23}
 - 2
 - 6.02×10^{23}
161. Hydrogen gas is not liberated when the following metal is added to dil. HCl
- Ag
 - Zn
 - Mg
 - Sn
162. In a hydrogen-oxygen fuel cell, combustion of hydrogen occurs to
- Generate heat
 - Create potential difference between the two electrodes
 - Produce high purity water
 - Remove adsorbed oxygen from electrode surfaces.
163. The sum of the two transport number of ions for an electrolyte is always equal to :
- 1
 - 2
 - 1/2
 - None of these
164. On passing 0.5 F electricity through molten sodium chloride, sodium deposited at cathode will be
- 29.25 g
 - 11.50 g
 - 58.50 g
 - 0.00 g
165. A solution of CuSO_4 is electrolysed for 10 min with a current of 1.5 A. What is the mass of copper deposited at the cathode?
- 2.096 g
 - 0.296 g
 - 3.029 g
 - 2.906 g
166. The correct order of molar conductivity at infinite dilution of LiCl, NaCl and KCl is
- $\text{LiCl} > \text{KCl} > \text{NaCl}$
 - $\text{KCl} > \text{NaCl} > \text{LiCl}$
 - $\text{LiCl} > \text{NaCl} > \text{KCl}$
 - $\text{NaCl} > \text{KCl} > \text{LiCl}$
167. Salts of A (atomic weight 7), B (atomic weight 27) and C (atomic weight 48) were electrolyzed under identical conditions using the same quantity of electricity. It was found that when 2.1 g of A was deposited, the weights of B and C deposited were 2.7 g and 7.2 g. The valencies of A, B and C are respectively :
- 3, 1 and 2
 - 1, 3 and 2
 - 3, 1 and 3
 - 2, 3 and 2
168. Indicator electrode is :
- SHE
 - Calomel electrode
 - Ag/AgCl electrode
 - Quinhydrone electrode
169. Molar conductance of electrolytic solution Λ_m is
- $\propto l$
 - $\propto (1/A)$
 - $\propto (1/C)$
 - $\propto (\sqrt{C})$
170. Which metal is most readily corroded in moist air?
- Copper
 - Iron
 - Silver
 - Nickel
171. Electrolysis of dilute aqueous NaCl solution was carried out by passing 10mA current. The time required to liberate 0.01 mole of H_2 gas at the cathode is ($1F = 96500 \text{ C mol}^{-1}$)
- $9.65 \times 10^4 \text{ s}$
 - $19.3 \times 10^4 \text{ s}$
 - $28.95 \times 10^4 \text{ s}$
 - $38.6 \times 10^4 \text{ s}$
172. The oxidation potential of Mg and Al are + 2.37 and + 1.66 volt respectively. The Mg in chemical reactions :
- Will be replaced by Al
 - Will replace Al

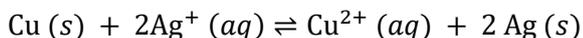
- c) Will not be able to replace Al
d) None of the above
173. The weight of silver (eq. wt. = 108) displaced by that quantity of current which displaced 5600 mL of hydrogen at STP is :
a) 54 g b) 108 g c) 5.4 g d) None of these
174. When during electrolysis of a solution of a AgNO_3 , 9650 C of charge pass through the electroplating bath, the mass of silver deposited on the cathode will be
a) 1.08 g b) 10.8 g c) 21.6 g d) 108 g
175. The standard oxidation potentials of the electrodes $\text{Ag} | \text{Ag}^+$, $\text{Sn} | \text{Sn}^{2+}$, $\text{Ca} | \text{Ca}^{2+}$, $\text{Pb} | \text{Pb}^{2+}$ are -0.8 , 0.136 , 2.866 and 0.126 V respectively. The most powerful oxidising agent among these metal ions is :
a) Pb^{2+} b) Ca^{2+} c) Sn^{2+} d) Ag^+
176. Pure water does not conduct electricity because it
a) Is neutral b) Is readily decomposed
c) Is almost totally unionized d) Has a low boiling point
177. The minimum equivalent conductance in fused state is shown by :
a) MgCl_2 b) BeCl_2 c) CaCl_2 d) SrCl_2
178. A cell necessarily does not contain :
a) An anode
b) A cathode
c) An electrolyte or a fuel
d) A porous diaphragm
179. The standard redox potentials for the reactions
 $\text{Mn}^{2+} + 2e^- \rightarrow$ and $\text{Mn}^{3+} + e^- \rightarrow \text{M}^{2+}$ are -1.18 V and 1.51 V respectively. What is the redox potential for the reaction
 $\text{Mn}^{3+} + 3e^- \rightarrow \text{Mn}$?
a) 0.33 V b) 1.69 V c) -0.28 V d) -0.85
180. During electrolysis of fused CaH_2 , H_2 is liberated at :
a) Anode b) Cathode c) Either electrode d) Not at all
181. Total charge on 1 mole of a monovalent metal ion is equal to :
a) 6.28×10^{18} coulomb b) 1.6×10^{-19} coulomb c) 9.65×10^4 coulomb d) None of these
182. For which case Λ values *vs* \sqrt{c} show a straight line?
a) KCl b) HCOOH c) CH_3NH_2 d) CH_3COOH
183. Which is not true for a standard hydrogen electrode?
a) The hydrogen ion concentration is $1M$
b) Temperature is 25°C
c) Pressure of hydrogen is 1 atmosphere
d) It contains a metallic conductor which does not adsorb hydrogen
184. The laws of electrolysis were proposed by
a) Kohlrausch b) Faraday c) Haber d) Bergius
185. The metal that cannot be obtained by electrolysis of the aqueous solution of its salts is :
a) Ag b) Cr c) Cu d) Al
186. A certain current liberated 0.504 g of hydrogen in 2 h. How many grams of copper can be liberated by the same current flowing for the same time in a copper sulphate solution?
a) 12.9 g b) 15.9 g c) 31.7 g d) 36.9 g
187. If mercury is used as cathode in the electrolysis of aqueous NaCl solution, the ions discharged at cathode are :
a) H^+ b) Na^+ c) OH^- d) Cl^-
188. Specific conductivity of a solution
a) Increases with dilution b) Decreases with dilution
c) Remains unchanged with dilution d) Depends on mass of electrolyte

189. When an electrolytic solution conducts electricity, current is carried out by :
 a) Electrons b) Cations and anions c) Neutral atoms d) None of these
190. e.m.f. of a cell in terms of reduction potential of its left and right electrode is :
 a) $E = E_L + E_R$ b) $E = E_L - E_R$ c) $E = E_R - E_L$ d) $E = -[E_R + E_L]$
191. Which defines the standard reduction electrode potential of Zn^{2+} ions?
 a) $Zn^{2+}(aq) + 2e \rightarrow Zn(s)$; $[Zn^{2+}] = 1M$
 b) $Zn(g) \rightarrow Zn^{2+} + 2e$; $[Zn^{2+}] = 1M$
 c) $Zn^{2+}(aq) \rightarrow Zn(s) + 2e$; $[Zn^{2+}] = 1M$
 d) $Zn^{2+}(g) \rightarrow Zn(s) - 2e$; $[Zn^{2+}] = 1M$
192. Given, the data at 25 °C ,
 $Ag + I^- \rightarrow AgI + e^-$; $E^\circ = 0.152 V$
 $Ag \rightarrow Ag^+ + e^-$; $E^\circ = -0.800V$
 What is the value of $\log K_{sp}$ for AgI?
 $\left(2.303 \frac{RT}{F} = 0.059 V \right)$
 a) - 8.12 b) +8.612 c) -37.83 d) -16.13
193. The molar conductivity of HCl, NaCl and CH_3COONa are 425, 188, 96 $S \text{ cm}^2 \text{ mol}^{-1}$ at 298 K. The molar conductivity of CH_3COOH at the same temperature is $S \text{ cm}^2 \text{ mol}^{-1}$.
 a) 333 b) 451 c) 325 d) 550
194. In the electrolysis of $CuCl_2$ solution using Cu electrodes the mass of cathode increases by 3.18 g. What happened at the other electrode?
 a) 0.05 mole of Cu^{2+} ions passed into solution
 b) 0.112 litre of Cl_2 was liberated
 c) 0.56 litre O_2 was liberated
 d) 0.1 mole of Cu^{2+} ions passed into the solution
195. When a quantity of electricity is passed through $CuSO_4$ solution, 0.16 g of copper gets deposited. If the same quantity of electricity is passed through acidulated water, then the volume of H_2 liberated at STP will be [Given , atomic weight of Cu = 64]
 a) 4.0 cm^3 b) 56 cm^3 c) 604 cm^3 d) 8.0 cm^3
196. Faraday's laws hold good at :
 a) All pressures b) Only at 298 K c) In different solvents d) All of these
197. The standard reduction potentials at 25°C of $Li + |Li, Ba^{2+}|Ba, Na^+|Na$ and $Mg^{2+}|Mg$ are - 3.05, -2.73, -2.71 and - 2.37 V respectively. Which is strongest reducing agent?
 a) Li b) Ba c) Na d) Mg
198. In which cell, electrical energy is converted into chemical energy?
 a) Water voltameter b) Silver voltameter c) Coulometer d) Either of these
199. Passage of 96500 coulomb of electricity liberateslitre of O_2 at NTP during electrolysis.
 a) 5.6 b) 6.5 c) 22.2 d) 11.2
200. The number of coulombs required for the deposition of 107.870 g silver is
 a) 96500 b) 48250 c) 1 d) 10000
201. The units of equivalent conductance, are
 a) $\Omega \text{ cm}^2 \text{ equiv}^{-1}$ b) $\Omega \text{ cm}^2 \text{ equiv}$ c) $\Omega^{-1} \text{ cm}^2 \text{ equiv}^{-1}$ d) $\Omega \text{ cm}^2 \text{ equiv}$
202. For strong electrolytes the plot of molar conductance *vs* \sqrt{C} is
 a) Parabolic b) Linear c) Sinusoidal d) Circular
203. The value of Λ_{eq}^∞ for NH_4Cl , NaOH and NaCl are respectively, 149.74, 248.1 and $126.4 \Omega^{-1} \text{ cm}^2 \text{ equiv}^{-1}$. The value of Λ_{eq}^∞ of NH_4OH is
 a) 371.44 b) 271.44
 c) 71.44 d) Cannot be predicted from given data

204. The standard electrode potentials of Ag^+ / Ag is +0.80 V and Cu^+ / Cu is +0.34 V. These electrodes are connected through Asalt bridge and if
- Copper electrode acts as Acathode then E_{cell}° is +0.46 V
 - Silver electrode acts as anode then E_{cell}° is -0.34 V
 - Copper electrode acts as anode then E_{cell}° is +0.46 V
 - Silver electrode acts as Acathode then E_{cell}° is -0.34 V
205. e.m.f. of cell $\text{Ni} | \text{Ni}^{2+} (0.1M) || \text{Au}^{3+} (1.0 M) | \text{Au}$ is, if E° for $\text{Ni}^{2+} | \text{Ni}$ is - 0.25 V, E° for $\text{Au}^{3+} | \text{Au}$ is 1.50 V.
- + 1.25 V
 - 1.75 V
 - + 1.75 V
 - + 4.0 V
206. The position of some metals in the electrochemical series in decreasing electropositive character is given as $\text{Mg} > \text{Al} > \text{Zn} > \text{Cu} > \text{Ag}$. What will happen, if a copper spoon is used to stir a solution of aluminium nitrate?
- The spoon will get coated with aluminium
 - An alloy of copper and aluminium is formed
 - The solution becomes blue
 - There is no reaction
207. Which of the following statements is correct? Galvanic cell converts
- Chemical energy into electrical energy
 - Electrical energy into chemical energy
 - Metal from its elemental state to the combined state
 - Electrolyte into individual ions
208. For cell reaction,
 $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$
 Cell representation is
- $\text{Zn} | \text{Zn}^{2+} || \text{Cu}^{2+} | \text{Cu}$
 - $\text{Cu} | \text{Cu}^{2+} || \text{Zn}^{2+} | \text{Zn}$
 - $\text{Cu} | \text{Zn}^{2+} || \text{Zn} | \text{Cu}^{2+}$
 - $\text{Cu}^{2+} | \text{Zn} || \text{Zn}^{2+} | \text{Cu}$
209. By passing 9.65 A current for 16 min 40 s , the volume of O_2 liberated at STP will be
- 280 mL
 - 560 mL
 - 1120 mL
 - 2240 mL
210. Consider the following disproportionation
 $2\text{ClO}_3^- \rightleftharpoons \text{ClO}_2^- + \text{ClO}_4^-$
 If the initial concentration of perchlorate ion is 0.1 M what it would be at equilibrium at 298 K?
 ($E_{\text{ClO}_4^- / \text{ClO}_3^-}^\circ = 0.36 \text{ V}$ and $E_{\text{ClO}_3^- / \text{ClO}_2^-}^\circ = 0.33 \text{ V}$)
- 0.1 M
 - 0.05 M
 - 0.07 M
 - 0.19 M
211. When Cu reacts with AgNO_3 solution, the reaction takes place is
- Oxidation of Cu
 - Reduction of Cu
 - Oxidation of Ag
 - Reduction of NO_3^-
212. E° for $\text{F}_2 + 2e = 2\text{F}^-$ is 2.8 V, E° for $1/2 \text{F}_2 + e = \text{F}^-$ is:
- 2.8 V
 - 1.4 V
 - 2.8 V
 - 1.4 V
213. Which one of the following solutions has highest conductance power?
- 0.1 M CH_3COOH
 - 0.1 M NaCl
 - 0.1 M KNO_3
 - 0.1 M HCl
214. Standard electrode potentials of $\text{Fe}^{2+} + 2e \rightarrow \text{Fe}$ and $\text{Fe}^+ + 3e \rightarrow \text{Fe}$ are - 440 V and -0.036 V respectively. The standard electrode potential (E°) for $\text{Fe}^{3+} + e \rightarrow \text{Fe}^{2+}$ is:
- 0.476 V
 - 0.404 V
 - + 0.404 V
 - + 0.772 V
215. Stainless steel does not rust because
- Chromium and nickel combine with iron
 - Chromium forms an oxide layer and protects iron from rusting
 - Nickel present in it, does not rust
 - Iron forms Ahard chemical compound with chromium present in it
216. Cu(II) sulphate solution is treated separately with KCl and KI . In which case, Cu^{2+} be reduced to Cu^+ ?
- With KCl
 - With KI
 - With both (a) and (b)
 - None of these
217. The main function of the salt bridge is :
- To allow ions to go from one cell to another

- b) To provide link between two half cells
 c) To keep the e.m.f. of the cell positive
 d) To maintain electrical neutrality of the solution in two half cells
218. When 9.65 C of electricity is passed through a solution of silver nitrate (atomic weight of Ag = 107.87 taking as 108), the amount of silver deposited is
 a) 5.8 mg b) 10.8 mg c) 15.8 mg d) 20.8 mg
219. The oxidation number of S in $\text{Na}_2\text{S}_4\text{O}_6$ is
 a) 2.5 for each S atom
 b) +2 and +3 (two S have +2 and other two have +3)
 c) +2 and +3 (three S have +2 and one S has +3)
 d) +5 and 0 (two S have +5 and other two have zero)
220. E° values of $\text{Mg}^{2+} / \text{Mg}$ is -2.37 V, of $\text{Zn}^{2+} / \text{Zn}$ is -0.76 V and $\text{Fe}^{2+} / \text{Fe}$ is -0.44 V.
 Which of the statements is correct?
 a) Zn will reduce Fe^{2+} b) Zn will reduce Mg^{2+}
 c) Mg oxidises Fe d) Zn oxidises Fe
221. Kohlrausch's law states that at :
 a) Infinite dilution each ion makes definite contribution to equivalent conductance of an electrolyte depending on the nature of the other ion of the electrolyte
 b) Infinite dilution, each ion makes definite contribution to conductance of an electrolyte whatever be the nature of the other ion of the electrolyte
 c) Infinite dilution, each ion makes definite contribution to equivalent conductance of an electrolyte, whatever be the nature of the other ion of the electrolyte
 d) Finite dilution, each ion makes definite contribution to equivalent conductance of an electrolyte, whatever be the nature of the other ion of the electrolyte
222. What is the potential of platinum wire dipped into a solution of 0.1 M in Sn^{2+} and 0.01 M in Sn^{4+} ?
 a) E° b) $E^\circ + 0.059$ c) $E^\circ + \frac{0.059}{2}$ d) $E^\circ - 0.059$
223. Electrolysis involves oxidation and reduction respectively at
 a) Anode and cathode b) Cathode and anode
 c) At both the electrodes d) None of these
224. The equivalent conductivity of two strong electrolytes at infinite dilution are :
 $\Lambda_{\text{CH}_3\text{COONa(aq)}} = 91.0 \text{ S cm}^2 \text{ eq.}^{-1}$
 $\Lambda_{\text{HCl(aq)}} = 426.2 \text{ S cm}^2 \text{ eq.}^{-1}$
 What additional information one needs to calculate Λ of an aqueous CH_3COOH ?
 a) Λ of ClCH_2COOH b) Λ of CH_3COOK c) Λ of NaCl d) Λ of H^+
225. In the electrolysis of CuSO_4 , the reaction $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$, takes place at :
 a) Anode b) Cathode c) In solution d) None of these
226. The standard emf of a galvanic cell involving 2 moles of electrons in its redox reaction is 0.59 V. The equilibrium constant for the redox reaction of the cell is
 a) 10^{20} b) 10^5 c) 10 d) 10^{10}
227. The amount of silver deposited by passing 241.25C of current through silver nitrate solution is
 a) 2.7 g b) 2.7 mg c) 0.27 g d) 0.54 g
228. The conductivity of saturated solution of CaF_2 is $3.86 \times 10^{-5} \text{ mho cm}^{-1}$ and that of water used for solution is 0.15×10^{-5} . The conductivity of CaF_2 alone is :
 a) 3.71×10^{-5} b) 4.01×10^{-5} c) 3.7×10^{-4} d) 3.86×10^{-4}
229. Four alkali metals A, B, C and D are having respectively standard electrode potential as -3.05, -1.66, -0.40 and 0.80. Which one will be the most reactive?
 a) A b) B c) C d) D
230. An increase in equivalent conductivity of strong electrolyte with dilution is mainly due to:
 a) Increase in ionic molality ions

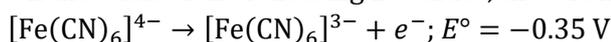
- b) 100% ionisation of electrolyte at normal dilution
 c) Increase in both i.e., no. of ions and ionic mobility
 d) Increase in no. of ions
231. The number of coulombs required to reduce 12.3 g of nitrobenzene to aniline is
 a) 115800 C b) 5790 C c) 28950 C d) 57900 C
232. The cell constant is
 a) $\frac{l}{a}$ b) $\frac{a}{l}$ c) $a \times l$ d) $\frac{\kappa}{R}$
233. The factors which influence the conductance of solution.
 a) Solute-solute interaction
 b) Solute-solvent interaction
 c) Temperature
 d) All of the above
234. In a cell containing zinc electrode and standard hydrogen electrode(SHE), the zinc electrode acts as :
 a) Anode
 b) Cathode
 c) Neither cathode nor anode
 d) Both anode and cathode
235. The best conductor of electricity is a 0.1 M solution of:
 a) Boric acid b) Sulphuric acid c) Acetic acid d) Propionic acid
236. Electrode potential of hydrogen electrode is volt.
 a) 0 b) +1 c) -1 d) None of these
237. Which aqueous solution will conduct an electric current quite well?
 a) Glycerol b) Sugar c) Hydrochloric acid d) Pure water
238. Use of electrolysis is not done in
 a) Production of Na b) Production of water
 c) Purification of metals d) Production of KOH
239. Beryllium is placed above magnesium in the II group. Beryllium dust, therefore, when added to $MgCl_2$ solution will :
 a) Have no effect
 b) Precipitate Mg metal
 c) Precipitate MgO
 d) Lead to dissolution of Be metal
240. When electric current is passed through an ionic hydride in molten state
 a) Hydrogen is liberated at anode b) Hydrogen is liberated at cathode
 c) No change takes place d) Hydride ions migrates towards cathode
241. Which of the following electrolytic solutions has the least specific conductance?
 a) 0.02 N b) 0.2 N c) 2 N d) 0.002 N
242. During the electrolysis of an electrolyte, the number of ions produced, is directly proportional to the
 a) Time consumed b) Mass of electrons
 c) Quantity of electricity passed d) Electrochemical equivalent of electrolytes
243. 1.8 g of metal were deposited by a current of 3 amperes for 50 minute. The equivalent wt. of metal is :
 a) 20.5 b) 25.8 c) 19.3 d) 30.7
244. Which substance is obtained in the solution on electrolysis of aqueous $CuSO_4$ solution using graphite electrodes?
 a) H_2O b) H_2SO_4 c) Na_2SO_4 d) $Cu(OH)_2$
245. During the electrolysis of fused NaCl, which reaction occurs at anode?
 a) Chloride ions are oxidized b) Sodium ions are oxidized
 c) Chloride ions are reduced d) Sodium ions are reduced
246. Which one of the following condition will increase the voltage of the cell represented by the equation?



- a) Increase in the dimension of Cu electrode
 b) Increase in the dimension of Ag electrode
 c) Increase in the concentration of Cu^{2+} ions
 d) Increase in the concentration of Ag^+ ions
247. Which will reduce zinc oxide to zinc?
 a) Mg
 b) Pb
 c) Cu
 d) Fe
248. The unit of electrochemical equivalent is :
 a) gram
 b) Gram/ampere
 c) Kg/coulomb
 d) Coulomb/gram
249. $\text{Sn}^{4+} + 3e^- \rightarrow \text{Sn}^{2+}, E^\circ = 0.13\text{ V}$
 $\text{Br}_2 + 2e^- \rightarrow 2\text{Br}^-, E^\circ = 1.08\text{ V}$
 Calculate K_{eq} for the cell reaction for the cell formed by two electrodes.
 a) 10^{41}
 b) 10^{32}
 c) 10^{-32}
 d) 10^{-42}
250. SI unit of conductivity is :
 a) $\text{ohm}^{-1}\text{ cm}^{-1}$
 b) $\text{ohm}^{-1}\text{ cm}^{-1}$ or Sm^{-1}
 c) ohm m^{-1}
 d) ohm cm^{-1}
251. Ionic mobility of Ag^+ is
 ($\lambda_{\text{Ag}^+} = 5 \times 10^{-1}\ \Omega^{-1}\text{ cm}^2\ \text{equiv}^{-1}$)
 a) 5.2×10^{-9}
 b) 2.4×10^{-9}
 c) 1.52×10^{-9}
 d) 8.25×10^{-9}
252. $E_{\text{Fe}^{3+}/\text{Fe}}^\circ = -0.036\text{ V}, E_{\text{Fe}^{2+}/\text{Fe}}^\circ = -0.439\text{ V}$. The value of standard electrode potential for the charge,
 $\text{Fe}^{3+} (\text{aq}) + e^- \rightarrow \text{Fe}^{2+} (\text{aq})$ will be
 a) -0.072 V
 b) 0.385 V
 c) 0.770 V
 d) -0.270 V
253. Whether tin can displace lead from aqueous lead bromide solution?
 a) No
 b) Yes, because standard reduction potential of Sn < Pb
 c) Yes, because standard reduction potential of Sn > Pb
 d) None of the above
254. Given the standard reduction potentials
 $\text{Zn}^{2+}/\text{Zn} = -0.74\text{ V}, \text{Cl}_2/\text{Cl}^- = 1.36\text{ V}$
 $\text{H}^+/\frac{1}{2}\text{H}_2 = 0\text{ V}$ and $\text{Fe}^{2+}/\text{Fe}^{3+} = 0.77\text{ V}$
 The order of increasing strength as reducing agent is
 a) $\text{Cl}^-, \text{Zn}, \text{H}_2, \text{Fe}^{2+}$
 b) $\text{H}_2, \text{Zn}, \text{Fe}^{2+}, \text{Cl}^-$
 c) $\text{Cl}^-, \text{Fe}^{2+}, \text{Zn}, \text{H}_2$
 d) $\text{Cl}^-, \text{Fe}^{2+}, \text{H}_2, \text{Zn}$
255. Molar conductivities (Λ_m) at infinite dilution of NaCl, HCl and CH_3COONa are 126.4, 425.9 and 91.0 $\text{S cm}^2\ \text{mol}^{-1}$ respectively. Λ_m for CH_3COOH will be :
 a) $425.5\text{ S cm}^2\ \text{mol}^{-1}$
 b) $180.5\text{ S cm}^2\ \text{mol}^{-1}$
 c) $290.8\text{ S cm}^2\ \text{mol}^{-1}$
 d) $390.5\text{ S cm}^2\ \text{mol}^{-1}$
256. $\text{KCl}(\text{aq})$ cannot be used as a salt bridge for the cell $\text{Cu}(\text{s})|\text{CuSO}_4(\text{aq})||\text{AgNO}_3(\text{aq})|\text{Ag}(\text{s})$ because :
 a) CuCl_2 is precipitated
 b) Cl_2 gas is given out
 c) AgCl is precipitated
 d) All of these
257. The ionic conductance of Ba^{2+} and Cl^- are respectively 127 and $76\ \Omega^{-1}\text{ cm}^2$ at infinite dilution. The equivalent conductance (in $\Omega^{-1}\text{ cm}^2$) of BaCl_2 at infinite dilution will be
 a) 139.5
 b) 203
 c) 279
 d) 101.5
258. The Gibbs energy for the decomposition of Al_2O_3 at 500°C is as follows :
 $\frac{2}{3}\text{Al}_2\text{O}_3 \rightarrow \frac{4}{3}\text{Al} + \text{O}_2, \Delta_r G = +966\text{ kJ mol}^{-1}$.
 The potential difference needed for electrolytic reduction of Al_2O_3 at 500°C is atleast :
 a) 5.0 V
 b) 4.5 V
 c) 3.0 V
 d) 2.5 V
259. Which of the following statements (or equation) is correct?
 a) The units of cell emf are V. cm^{-1}
 b) $\Delta G = -\frac{nF}{E_{\text{cell}}}$
 c) In galvanic cell, chemical energy is transformed into electrical energy.
 d) Oxidation state of Mn in potassium permanganate is +6

260. Faraday's law of electrolysis fails when :
- Temperature is increased
 - Inert electrodes are used
 - A mixture of electrolytes is used
 - In none of the above cases
261. Conductance 'C' (in S) is directly proportional to the area of the electrode and concentration and inversely proportional to length of separation of electrode, the unit of constant of proportionality is:
- $S\ m\ mol^{-1}$
 - $S\ m^2\ mol^{-1}$
 - $S^{-2}\ m^2\ mol$
 - $S^2\ m^2\ mol^2$
262. A certain metal fails to liberate H_2 gas from a moderately conc. HCl solution. However, it displaces Ag from $AgNO_3$ solution. Which among the followings may it be?
- Mg
 - Fe
 - Cu
 - Cd
263. For the cell reaction $Fe + 2Fe^{3+} = 3Fe^{2+}$, which is not possible?
- One cell can be constructed
 - Three different cells with different E_{cell}° are possible
 - Three different cells with different number of electrons used in redox reaction are possible
 - Three different cells with same ΔG° value are possible
264. Copper sulphate solution does not react with
- Zinc
 - Iron
 - Silver
 - All of these
265. Rust is
- $FeO + Fe(OH)_2$
 - Fe_2O_3
 - $Fe_2O_3 + Fe(OH)_2$
 - Fe_2O_3 and $Fe(OH)_3$
266. The conductivity of strong electrolyte
- Increases on dilution slightly
 - Decreases on dilution
 - Does not change with dilution
 - Depends upon density of electrolyte itself
267. An electric current of c ampere was passed through a solution of an electrolyte for ' t ' second depositing P g of the metal M on the cathode. The equivalent weight E of the metal will be :
- $E = \frac{c \times t}{P \times 96500}$
 - $E = \frac{c \times P}{t \times 96500}$
 - $E = \frac{96500 \times P}{c \times t}$
 - $E = \frac{c \times t \times 9650}{P}$
268. Cu^+ ion is not stable in aqueous solution because of disproportionation reaction. E° value for disproportionation of Cu^+ is (given, $E_{Cu^{2+}/Cu^+}^\circ = 0.15$, $E_{Cu^{2+}/Cu}^\circ = 0.34$ V)
- + 0.38 V
 - 0.38 V
 - + 0.49 V
 - 0.49 V
269. The molar conductivities Λ_{NaOAc}° and Λ_{HCl}° at infinite dilution in water at $25^\circ C$ are 91.0 and 426.2 S cm^2/mol respectively. To calculate Λ_{HOAc}° , the additional value required is
- $\Lambda_{H_2O}^\circ$
 - Λ_{KCl}°
 - Λ_{NaOH}°
 - Λ_{NaCl}°
270. The molar conductivity 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 conductivity of CH_3COOH at infinite dilution is :
- 201.28 $ohm^{-1}\ cm^2\ mol^{-1}$
 - 698.28 $ohm^{-1}\ cm^2\ mol^{-1}$
 - 390.71 $ohm^{-1}\ cm^2\ mol^{-1}$
 - 540.48 $ohm^{-1}\ cm^2\ mol^{-1}$
271. If the electrolyte used in problem 4 is $Ba(NO_3)_2$, then molecular conductivity of solution is :
- 628.56 $mho^{-1}\ cm^2\ mol$
 - 628.56 $mho\ cm^2\ mol^{-1}$
 - 6.28 $mho\ cm^2\ mol^{-1}$
 - 6.28 $mho^{-1}\ cm^2\ mol^{-1}$
272. The equivalent conductivity of KCl at infinite dilution is 130 $mho\ cm^2\ eq^{-1}$. The transport number of Cl^- ion in KCl at the same temperature is 0.505. The transport number of K^+ ion is :
- 0.495
 - 0.505
 - 0.0495
 - Cannot be predicted
273. A cell in which electric current is produced by net oxidation and reduction process is called :
- Voltaic cell
 - Electrolytic cell
 - Concentration cell
 - None of these
274. Electrolysis of aq. Solution of LiCl shows :
- $pH < 7$
 - $pH = 7$
 - $pH > 7$
 - No change

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



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

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

- a) 0.66 cm^{-1} b) 1.12 cm^{-1} c) 0.918 cm^{-1} d) 1.66 cm^{-1}

277. 20 g of chlorine are evolved in 6 hour from sodium chloride solution by the current of :

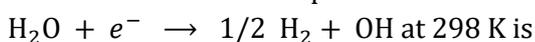
- a) 5 ampere b) 10 ampere c) 2.5 ampere d) 50 ampere

278. For the electrochemical cell, $M | M^+ || X^- | X$,

$E^\circ(M^+ | M) = 0.44 \text{ V}$, $E^\circ(X | X^-) = 0.33 \text{ V}$. From this data one can deduce that

- a) $E^\circ_{\text{cell}} = -0.77 \text{ V}$
b) $M^+ + X^- \rightarrow M + X$ is the spontaneous reaction
c) $M + X \rightarrow M^+ + X^-$ is the spontaneous reaction
d) $E^\circ_{\text{cell}} = 0.77 \text{ V}$

279. The standard reduction potential of the reaction,



a) $E^\circ = \frac{RT}{F} \ln K_w$

b) $E^\circ = -\frac{RT}{F} \ln[\text{P}_{\text{H}_2}]^{1/2}[\text{OH}^-]$

c) $E^\circ = -\frac{RT}{F} \ln \frac{[\text{P}_{\text{H}_2}]^{1/2}}{[\text{H}^+]}$

d) $E^\circ = -\frac{RT}{F} \ln K_w$

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

- a) $\text{Cr} > \text{Mn} > \text{Fe} > \text{Co}$ b) $\text{Mn} > \text{Cr} > \text{Fe} > \text{Co}$ c) $\text{Cr} > \text{Fe} > \text{Mn} > \text{Co}$ d) $\text{Fe} > \text{Mn} > \text{Cr} > \text{Co}$

281. The increase in equivalent conductivity of a weak electrolyte solution with dilution is attributed to :

- a) Increase in degree of dissociation
b) Increase in ionic mobility
c) Both (a) and (b)
d) None of the above

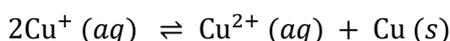
282. Resistance of 0.2 M solution of an electrolyte is 50Ω . The specific conductance of the solution is 1.3 S m^{-1} . If resistance of the 0.4M solution of the same electrolyte is 260Ω , its molar conductivity is

- a) $6250 \text{ Sm}^2 \text{ mol}^{-1}$ b) $6.25 \times 10^{-4} \text{ Sm}^2 \text{ mol}^{-1}$
c) $625 \times 10^{-4} \text{ Sm}^2 \text{ mol}^{-1}$ d) $62.5 \text{ Sm}^2 \text{ mol}^{-1}$

283. When electric current is passed through a cell having an electrolyte, the positive ions move towards the cathode and the negative ions towards the anode. If the cathode is pulled out of the solution :

- a) The positive and the negative ions both will move towards the anode
b) The positive ions will start moving towards the anode; the negative ions will stop moving
c) The negative ions will continue to move towards the anode; the positive ions will stop moving
d) The positive ions and the negative ions will start moving randomly

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



choose correct E° for the above reaction if

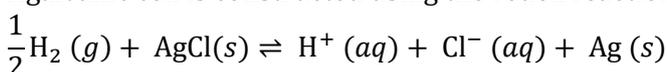
$$E^\circ_{\text{Cu}^{2+}/\text{Cu}} = 0.34 \text{ V}, E^\circ_{\text{Cu}^{2+}/\text{Cu}^+} = 0.15 \text{ V}$$

- a) -0.38 V b) $+0.49 \text{ V}$ c) $+0.38 \text{ V}$ d) -0.19 V

285. Standard electrode potential of NHE at 298 K is

- a) 0.05 V b) 0.10 V c) 0.50 V d) 0.00 V

286. Galvanic cell is constructed using the redox reaction,



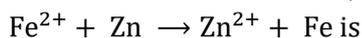
It is represented as

- a) $\text{Pt} | \text{H}_2(\text{g}) | \text{HCl solution} || \text{AgNO}_3 \text{ solution} |$
 Ag
- b) $\text{Ag} | \text{AgCl(s)} | \text{KCl solution} | | \text{HCl solution} |$
 $| \text{H}_2(\text{g}) | \text{Pt}$
- c) $\text{Pt} | \text{H}_2(\text{g}) | \text{KCl solution} || \text{AgCl(s)} | \text{Ag}$
- d) $\text{Pt} | \text{H}_2(\text{g}), \text{HCl solution} || \text{AgCl(s)} | \text{Ag}$
287. $\text{Zn} | \text{Zn}^{2+} (\text{A} = 0.1 \text{ M}) || \text{Fe}^{2+} (\text{A} = 0.01 \text{ M}) | \text{Fe}$.
 The emf of the above cell is 0.2905 V. Equilibrium constant for the cell reaction is
- a) $10^{0.32/0.0591}$ b) $10^{0.32/0.0295}$ c) $10^{0.26/0.0295}$ d) $e^{0.32/0.0295}$
288. The conductance of a solution of an electrolyte is same as that of its conductivity. The cell used can be said to have cell constant equal to :
- a) 1 b) Zero c) 100 d) 10
289. A conductivity cell has two platinum electrodes of 1.2 cm^2 area, separated by a distance of 0.8 cm. The cell constant is :
- a) 0.66 cm^{-1} b) 1.5 cm^{-1} c) 0.96 cm^{-1} d) 0.66 cm
290. A current of i ampere was passed for t second through three cells P, Q and R connected in series. These contain respectively silver nitrate, mercuric nitrate and mercurous nitrate. At the cathode of the cell P , 0.216 g of Ag was deposited. The weights of mercury deposited in the cathode of Q and R respectively are :
- a) 0.4012 and 0.8024 g b) 0.4012 and 0.2006 g c) 0.2006 and 0.4012 g d) 0.1003 and 0.2006 g
291. Out of Cu, Al, Fe and Zn, metal which can displace all others from their salt solution is
- a) Al b) Cu c) Zn d) Fe
292. 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
 $\Lambda^\circ_{\text{CH}_3\text{COONa}} = 91.0 \text{ S cm}^2/\text{equiv}$
 $\Lambda^\circ_{\text{HCl}} = 426.2 \text{ S cm}^2/\text{equiv}$
 What additional information/quantity one needs to calculate Λ° of an aqueous solution of acetic acid?
- a) Λ° of NaCl
 b) Λ° of CH_3COOK
 c) The limiting equivalent conductance of H^+ ($\lambda^\circ_{\text{H}^+}$)
 d) Λ° of chloroacetic acid (ClCH_2COOH)
293. The emf of the cell
 $\text{Ni} | \text{Ni}^{2+} (1.0 \text{ M}) || \text{Au}^{3+} (1.0 \text{ M}) | \text{Au}$
 is [$E^\circ (\text{Ni}^{2+} / \text{Ni}) = -0.25 \text{ V}$ and
 $E^\circ (\text{Au}^{3+} / \text{Au}) = +1.5 \text{ V}$]
- a) 2.00 V b) 1.25 V c) -1.25 V d) 1.75 V
294. The standard reduction potential for Fe^{2+}/Fe and Sn^{2+}/Sn electrodes are -0.44 and -0.14 V respectively. For the given cell reaction, $\text{Fe}^{2+} + \text{Sn} \rightarrow \text{Fe} + \text{Sn}^{2+}$, the standard emf is
- a) 0.42 V b) -0.42 V c) -0.30 V d) -1.10 V
295. In a cell that utilises the reaction,
 $\text{Zn}(\text{s}) + 2\text{H}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{H}_2(\text{g})$
 addition of H_2SO_4 to cathode compartment, will
- a) Lower the E and shift the equilibrium to the left
 b) Lower the E and shift the equilibrium to the right
 c) Increase the E and shift the equilibrium to the right
 d) Increase the E and shift the equilibrium to the left
296. Which will increase the voltage of the cell $\text{Sn}(\text{s}) + 2\text{Ag}^+(\text{aq}) \rightarrow \text{Sn}^{2+}(\text{aq}) + 2\text{Ag}(\text{s})$?
- a) Increase in size of the silver rod
 b) Increase in the concentration of Sn^{2+} ions
 c) Increase in the concentration of Ag^+ ions

- d) None of the above
297. Given standard electrode potentials
 $Fe^{2+} + 2e^- \rightarrow Fe$ $E^\circ = -0.440 V$
 $Fe^{3+} + 3e^- \rightarrow Fe$ $E^\circ = -0.036 V$
 The standard electrode potential (E°) for
 $Fe^{3+} + e^- \rightarrow Fe^{2+}$ is
 a) +0.772 V b) -0.772 V c) +0.417 V d) -0.414 V
298. Normal Al – AlCl₃ coupled with normal hydrogen electrode gives an e.m.f. of 1.66 V. The standard oxidation electrode potential of aluminium is :
 a) – 1.66 V b) + 1.66 V c) – 0.83 V d) + 0.83 V
299. Which of the following statements is true for fuel cells?
 a) They are more efficient b) They are free from pollution
 c) They run till reactants are active d) All of the above
300. The Λ° of NH₄OH at infinite dilution is ... S cm² eq.⁻¹. Given $\lambda_{OH^-}^\infty = 174$, $\lambda_{Cl^-}^\infty = 66$ and $\lambda_{NH_4Cl}^\infty = 130$ S cm² eq.⁻¹ :
 a) 238 b) 218 c) 198 d) 160
301. The metal that forms a self protecting film of oxide to prevent corrosion is
 a) Na b) Al c) Cu d) Au
302. The number of Faraday's needed to reduce 4 g-equivalents of Cu²⁺ to Cu metal will be
 a) 1 b) 2 c) 4 d) 8
303. The atomic weight of Al is 27. When a current of 5 faraday is passed through a solution of Al³⁺ ions, the wt.of Al deposited is :
 a) 27 g b) 36 g c) 45 g d) 9 g
304. Which is correct representation for a cell at equilibrium?
 a) $\Delta G^\circ = -2.303 RT \log K_{eq}$.
 b) $E^\circ = \frac{2.303RT}{nF} \log K_{eq}$.
 c) $-\Delta G^\circ = RT \ln K_{eq}$
 d) All of the above.
305. Consider the following E° values
 $E^\circ_{Fe^{3+}/Fe^{2+}} = +0.77 V$
 $E^\circ_{Sn^{2+}/Sn} = - 0.14 V$
 Under standard conditions, the potential for the reaction
 $Sn(s) + 2Fe^{3+}(aq) \rightarrow 2Fe^{2+}(aq) + Sn^{2+}(aq)$ is
 a) – 0.91 V b) + 0.91 V c) – 0.41 V d) + 0.41 V
306. According to Kohlrausch's law the limiting value of equivalent conductivity of an electrolyte A₂B is given by :
 a) $\lambda_{A^+}^\infty + \lambda_{B^{-2}}^\infty$ b) $\frac{1}{2}\lambda_{A^+}^\infty + \lambda_{B^{-2}}^\infty$ c) $\lambda_{A^+}^\infty + \frac{1}{2}\lambda_{B^{-2}}^\infty$ d) $2\lambda_{A^+}^\infty + \lambda_{B^{-2}}^\infty$
307. Standard reduction potential of an element is equal to :
 a) + 1 × its reduction potential
 b) – 1 × its standard oxidation potential
 c) 0.00 V
 d) + 1 × its standard oxidation potential
308. Rusting of iron is catalysed by which of the following?
 a) Fe b) Zn c) O₂ d) H⁺
309. The equivalent conductivity of monobasic acid at infinite dilution is 348 ohm⁻¹ cm² eq.⁻¹. If the resistivity of the solution containing 15 g acid (mol. wt. 49) in 1 litre is 18.5 ohm cm, what is the degree of dissociation of acid?

- a) 45.9% b) 40.2% c) 60.4% d) 50.7%
310. The standard reduction potential E° for half reactions are
 $\text{Zn} = \text{Zn}^{2+} + 2e^- \quad E^\circ = +0.76 \text{ V}$
 $\text{Fe} = \text{Fe}^{2+} + 2e^- \quad E^\circ = +0.41 \text{ V}$
 The emf of the cell reaction
 $\text{Fe}^{2+} + \text{Zn} = \text{Zn}^{2+} + \text{Fe}$ is
 a) -0.35 V b) +0.35 c) +1.17 V d) -1.17 V
311. If a salt bridge is removed from the two half cells, the voltage
 a) Drops to zero b) Does not change c) Increase gradually d) Increase rapidly
312. The standard oxidation potentials of Zn, Cu, Ag and Ni electrodes are +0.76, -0.34, -0.80 and +0.25 V respectively. Which of the following reaction will provide maximum voltage?
 a) $\text{Cu} + 2\text{Ag}^+(aq) \rightarrow \text{Cu}^{2+}(aq) + 2\text{Ag}$ b) $\text{Zn} + 2\text{Ag}^+(aq) \rightarrow \text{Zn}^{2+}(aq) + 2\text{Ag}$
 c) $\text{H}_2 + \text{Ni}^{2+}(aq) \rightarrow 2\text{H}^+(aq) + \text{Ni}$ d) $\text{Zn} + \text{Cu}^{2+}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Cu}$
313. An apparatus used for the measurement of quantity of electricity is known as a :
 a) Calorimeter b) Cathetometer c) Coulometer d) colorimeter
314. For the cell prepared from electrode A and B; Electrode A : $\text{Cr}_2\text{O}_7^{2-} | \text{Cr}^{3+}$, $E^\circ_{\text{red}} = +1.33 \text{ V}$ and Electrode B : $\text{Fe}^{3+} | \text{Fe}^{2+}$, $E^\circ_{\text{red}} = 0.77 \text{ V}$. Which of the following statement is correct?
 a) The electrons will flow from B to A when connection is made
 b) The e.m.f. of the cell will be 0.56 V
 c) A will be positive electrode
 d) All of the above
315. Which colourless gas evolves when NH_4Cl reacts with zinc in Adry cell battery?
 a) NH_3 b) N_2 c) H_2 d) Cl_2
316. The standard E°_{red} values of A, B and C are +0.68 V, -2.54 V, -0.50 V respectively. The order of their reducing power is
 a) $A > B > C$ b) $A > C > B$ c) $C > B > A$ d) $B > C > A$
317. Based on the data given below, the correct order of reducing power is :
 $\text{Fe}^{3+}(aq) + e^- \rightarrow \text{Fe}^{2+}(aq); E^\circ = +0.77 \text{ V}$
 $\text{Al}^{3+}(aq) + 3e^- \rightarrow \text{Al}(s); E^\circ = -1.66 \text{ V}$
 $\text{Br}_2(aq) + 2e^- \rightarrow 2\text{Br}^-(aq); E^\circ = +1.08 \text{ V}$
 a) $\text{Br}^- < \text{Fe}^{2+} < \text{Al}$ b) $\text{Fe}^{2+} < \text{Al} < \text{Br}^-$ c) $\text{Al} < \text{Br}^- < \text{Fe}^{2+}$ d) $\text{Al} < \text{Fe}^{2+} < \text{Br}^-$
318. Small quantities of solutions of compounds TX, TY and TZ are put into separate test tubes containing X, Y and Z solutions. TX does not react with any of these. TY reacts with both X and Z. TZ reacts with X. The decreasing order of ease of oxidation of the anions X^- , Y^- , Z^- is
 a) Y^-, Z^-, X^- b) Z^-, X^-, Y^- c) Y^-, X^-, Z^- d) X^-, Z^-, Y^-
319. What flows in the internal circuit of a galvanic cell?
 a) Ions b) Electrons c) Electricity d) Atoms
320. The standard electrode potential of Zn^{2+}/Zn and Ag^+/Ag are -0.763 V and +0.799 V respectively. The standard potential of the cell is
 a) 1.56 V b) -1.56 V c) 0.036 V d) -0.036 V
321. Consider the following E° values
 $E^\circ_{\text{Fe}^{3+}/\text{Fe}^{2+}} = +0.77 \text{ V}$
 $E^\circ_{\text{Sn}^{2+}/\text{Sn}} = -0.14 \text{ V}$
 Under standard conditions the potential for the reaction
 $\text{Sn}(s) + 2\text{Fe}^{3+}(aq) \rightarrow 2\text{Fe}^{2+}(aq) + \text{Sn}^{2+}(aq)$ is
 a) 1.68 V b) 1.40 V c) 0.91 V d) 0.63 V
322. The standard electrode potential for the half-cell reactions are
 $\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}; E^\circ = -0.76 \text{ V}$
 $\text{Fe}^{2+} + 2e^- \rightarrow \text{Fe}; E^\circ = -0.44 \text{ V}$

The emf of the cell reaction,



- a) -0.32 V b) -1.20 V c) +1.20 V d) +0.32 V

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

- a) $p(\text{H}_2) = 1 \text{ atm}$ and $[\text{H}^+] = 2.0 \text{ M}$ b) $p(\text{H}_2) = 1 \text{ atm}$ and $[\text{H}^+] = 1.0 \text{ M}$
c) $p(\text{H}_2) = 2 \text{ atm}$ and $[\text{H}^+] = 1.0 \text{ M}$ d) $p(\text{H}_2) = 2 \text{ atm}$ and $[\text{H}^+] = 2.0 \text{ M}$

324. Give the products available on the cathode and the anode respectively during the electrolysis of an aqueous solution of MgSO_4 between inert electrodes.

- a) $\text{H}_2(\text{g})$ and $\text{O}_2(\text{g})$ b) $\text{O}_2(\text{g})$ and $\text{H}_2(\text{g})$ c) $\text{O}_2(\text{g})$ and $\text{Mg}(\text{s})$ d) $\text{O}_2(\text{g})$ and $\text{SO}_2(\text{g})$

325. Which of the following statements is not applicable to electrolytic conductors?

- a) A single stream of electrons flows from cathode to anode
b) Show a positive temperature coefficient for conductance
c) New products show up at the electrodes
d) Ions are responsible for carrying the current

326. Which of the following expressions correctly represents the equivalent conductance at infinite dilution of $\text{Al}_2(\text{SO}_4)_3$. Given that $\Lambda_{\text{Al}^{3+}}^\infty$ and $\Lambda_{\text{SO}_4^{2-}}^\infty$ are the equivalent conductances at infinite dilution of the respective ions?

- a) $2\Lambda_{\text{Al}^{3+}}^\infty + 3\Lambda_{\text{SO}_4^{2-}}^\infty$ b) $2\Lambda_{\text{Al}^{3+}}^\infty + \Lambda_{\text{SO}_4^{2-}}^\infty$ c) $(\Lambda_{\text{Al}^{3+}}^\infty + \Lambda_{\text{SO}_4^{2-}}^\infty) \times 6$ d) $\frac{1}{3}\Lambda_{\text{Al}^{3+}}^\infty + \frac{1}{2}\Lambda_{\text{SO}_4^{2-}}^\infty$

327. Conductivity of a strong electrolyte

- a) Decreases on dilution b) Increases on dilution
c) Does not change considerably on dilution d) Depends on density

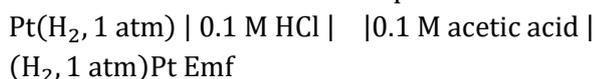
328. Which of the following compounds will not undergo decomposition on passing electricity through aqueous solution?

- a) Sugar b) Sodium acetate c) Sodium chloride d) Sodium bromide

329. Which loses charge at cathode?

- a) Ions
b) Cations
c) Anions
d) Both anions and cations

330. An electrochemical cell is set up as follows



Emf of this cell will not be zero because

- a) The pH of 0.1 M HCl and 0.1 M acetic acid is not the same b) Acids used in two compartments are different
c) Emf of a cell depends on the molarities of acids used d) The temperature is constant

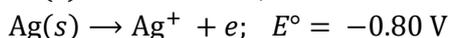
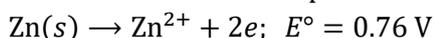
331. Which of the following reactions cannot be a base for electrochemical cell?

- a) $\text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O}$
b) $\text{AgNO}_3 + \text{Zn} \rightarrow \text{Zn}(\text{NO}_3)_2 + \text{Ag}$
c) $\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} \downarrow + \text{NaNO}_3$
d) $\text{KMnO}_4 + \text{FeSO}_4 + \text{H}_2\text{SO}_4 \rightarrow \text{K}_2\text{SO}_4 + \text{Fe}_2(\text{SO}_4)_3 + \text{MnSO}_4 + \text{H}_2\text{O}$

332. The emf of a galvanic cell, with electrode potentials of silver = + 0.80 V and that of copper = + 0.34 V, is

- a) + 0.46 V b) + 0.66 V c) + 0.86 V d) - 0.66 V

333. The standard oxidation potentials of Zn and Ag in water at 25°C are,



Which reaction actually takes place?

- a) $\text{Zn}(s) + 2\text{Ag}^+(aq) \rightarrow \text{Zn}^{2+} + \text{Ag}(s)$
 b) $\text{Zn}^{2+} + 2\text{Ag}^+(s) \rightarrow 2\text{Ag}^+(aq) + \text{Zn}(s)$
 c) $\text{Zn}(s) + 2\text{Ag}(s) \rightarrow \text{Zn}^{2+}(aq) + \text{Ag}^+(aq)$
 d) $\text{Zn}^{2+}(aq) + \text{Ag}^+(aq) \rightarrow \text{Zn}(s) + \text{Ag}(s)$
334. Amount of electricity that can deposit 108 g of silver from AgNO_3 solution is
 a) 1 F b) 1 A c) 1 C d) None of these
335. Also the $[\text{H}^+]$ for problem 9 using the same data is :
 a) 0.00133 M b) 0.133 M c) 0.0133 M d) None of these
336. A hydrogen electrode placed in a buffer solution of CH_3COONa and acetic acid in the ratio's $x : y$ and $y : x$ has electrode potential values E_1 volt and E_2 volt respectively at 25°C . The pK_a values of acetic acid is (E_1 and E_2 are oxidation potential) :
 a) $\frac{E_1 + E_2}{0.118}$
 b) $\frac{E_2 - E_1}{0.118}$
 c) $\frac{E_1 - E_2}{0.118}$
 d) $\frac{E_1 + E_2}{0.118}$
337. When an electric current is passed through acidulated water, 112 mL of hydrogen gas at STP collects at the cathode in 965 second. The current passed, in ampere is :
 a) 1.0 b) 0.5 c) 0.1 d) 2.0
338. The hydrogen electrode is dipped in a solution of pH 3 at 25°C . The potential would be (the value of $2.303 RT / F$ is 0.059 V)
 a) 0.177 V b) 0.087 V c) 0.059 V d) -0.177 V
339. The potential of the following cell is 0.34 V at 25°C . Calculate the standard reduction potential of the copper half-cell.
 $\text{Pt} | \text{H}_2(1 \text{ atm}) | \text{H}^+(1 \text{ M}) || \text{Cu}^{2+}(1 \text{ M}) | \text{Cu}$
 a) -3.4 V b) +3.4 V c) -0.34 V d) +0.34 V
340. The hydrogen electrode is dipped in a solution of pH = 3 at 25°C . The potential of the cell would be (the value of $\frac{2.303RT}{T}$ is 0.059 V)
 a) 0.059 V b) 0.088 V c) 0.178 V d) -0.177 V
341. The apparatus in which electrical energy is converted into chemical energy is known as :
 a) Voltmeter b) Coulometer c) Both (a) and (b) d) None of these
342. When 1 F of electricity is passed through acidulated water, O_2 evolved is
 a) 1.0 dm^3 b) 5.6 dm^3 c) 11.2 dm^3 d) 22.4 dm^3
343. A current of strength 2.5 A was passed through CuSO_4 solution for 6 min 26 s. The amount of copper deposited is:
 (Atomic weight of Cu = 63.5) (1 F = 96500 C)
 a) 0.3175 g b) 3.0175 g c) 7.0135 g d) 6.0275 g
344. The potential of the cell for the reaction,
 $\text{M}(s) + 2\text{H}^+(1\text{M}) \rightarrow \text{H}_2(g)(1 \text{ atm}) + \text{M}^{2+}(0.1 \text{ M})'$
 is 1.500 V. The standard reduction potential for $\text{M}^+ / \text{M}(s)$ couple is
 a) 0.1470 V b) 1.470 V c) 14.70 V d) None of these
345. For the cell $\text{Zn} | \text{Zn}^{2+} || \text{Cu}^{2+} | \text{Cu}$, if the concentration of Zn^{2+} and Cu^{2+} ions is doubled, the e.m.f. of the cell :
 a) Doubles b) Reduces to half c) Remains same d) Becomes zero
346. The ionic conductivity of H^+ and OH^- at 298 K are 349.8 and 198.5 $\text{mho cm}^2 \text{eq}^{-1}$ respectively. The equivalent conductivity of H_2O at infinite dilution is :
 a) 548.3 b) 151.3 c) 699.6 d) 54.83

347. The equivalent conductivity of a solution containing 2.54 g of CuSO_4 per L is $91.0 \Omega^{-1} \text{cm}^2 \text{eq}^{-1}$. Its conductivity would be
 a) $2.9 \times 10^{-3} \Omega^{-1} \text{cm}^2$ b) $1.8 \times 10^{-2} \Omega^{-1} \text{cm}^2$ c) $2.4 \times 10^{-4} \Omega^{-1} \text{cm}^2$ d) $3.6 \times 10^{-3} \Omega^{-1} \text{cm}^2$
348. Standard electrode potential data are useful for understanding the suitability of an oxidant in a redox titration. Some half cell reactions and their standard potentials are given below ;
 $\text{MnO}_4^- (aq) + 8\text{H}^+ (aq) + 5e^- \rightarrow \text{Mn}^{2+} (aq) + 4\text{H}_2\text{O}; E^\circ = 1.51 \text{ V}$
 $\text{Cr}_2\text{O}_7^{2-} (aq) + 14\text{H}^+ (aq) + 6e^- \rightarrow 2\text{Cr}^{3+} (aq) + 7\text{H}_2\text{O}; E^\circ = 1.38 \text{ V}$
 $\text{Fe}^{3+} (aq) + e^- \rightarrow \text{Fe}^{2+} (aq); E^\circ = 0.77 \text{ V}$
 $\text{Cl}_2 (g) + 2e^- \rightarrow 2\text{Cl}^- (aq); E^\circ = 1.40 \text{ V}$
 Identify the only incorrect statement regarding the quantitative estimation of aqueous $\text{Fe}(\text{NO}_3)_2$:
 a) MnO_4^- can be used in aqueous HCl
 b) $\text{Cr}_2\text{O}_7^{2-}$ can be used in aqueous HCl
 c) MnO_4^- can be used in aqueous H_2SO_4
 d) $\text{Cr}_2\text{O}_7^{2-}$ can be used in aqueous H_2SO_4
349. The standard H electrode is written as :
 a) Pt, H_2 , $\text{H}^+ (a = 1)$
 b) $\text{PtH}_2 / \text{H}^+ (a = 1)$
 c) $\text{PtH}_2 (g) (1 \text{ atm}) / \text{H}^+ (a = 1)$
 d) None of the above
350. Standard electrode potential of cell $\text{H}_2 | \text{H}^+ || \text{Ag}^+ | \text{Ag}$ is
 a) 0.8 V b) -0.8 V c) -1.2 V d) 1.2 V
351. A dilute solution of Li_2SO_4 is electrolyzed. The products formed at the anode and cathode, respectively are :
 a) S and Li b) O_2 and Li c) SO_2 and O_2 d) O_2 and H_2
352. 3 F electricity was passed through an aqueous solution of iron (II) bromide. The weight of iron metal (at. Wt. = 56) deposited at the cathode (in g) is
 a) 65 b) 84 c) 112 d) 168
353. 5 A is passed through a solution of zinc sulphate for 40 min. Find the amount of zinc deposited at the cathode
 a) 4.065 g b) 8.065 g c) 16.065 g d) 32.065 g
354. Saturated solution of KNO_3 is used to make 'salt-bridge' because
 a) Velocities of both K^+ and NO_3^- are nearly the same
 b) Velocity of K^+ is greater than that of NO_3^-
 c) Velocity of NO_3^- is greater than that of K^+
 d) KNO_3 is highly soluble in water
355. The calomel electrode is a :
 a) Standard hydrogen electrode
 b) Reference electrode
 c) Platinum electrode
 d) Mercury electrode
356. Calculate the emf of the cell
 $\text{Cu} (s) | \text{Cu}^{2+} (aq) || \text{Ag}^+ (aq) | \text{Ag} (s)$
 Given,
 $E^\circ_{\text{Cu}^{2+}/\text{Cu}} = +0.34 \text{ V}, E^\circ_{\text{Ag}^+/\text{Ag}} = 0.80 \text{ V},$
 a) +0.46 V b) +1.14 V c) +0.57 V d) -0.46 V
357. The electrolytic conductance is a direct measure of
 a) Resistance b) Potential c) Dissociation d) Concentration

358. On the basis of electrochemical theory of aqueous corrosion, the reaction occurring at the cathode is
- a) $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$ b) $H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$
 c) $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$ d) $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$
359. The resistance of $\frac{N}{10}$ solution is found to be $2.5 \times 10^3 \Omega$. The equivalent conductance of the solution is (cell constant = 1.25 cm^{-1})
- a) $2.5 \Omega^{-1} \text{ cm}^2 \text{ equiv}^{-1}$ b) $5.0 \Omega^{-1} \text{ cm}^2 \text{ equiv}^{-1}$
 c) $2.5 \Omega^{-1} \text{ cm}^{-2} \text{ equiv}^{-1}$ d) $5.0 \Omega^{-1} \text{ cm}^{-2} \text{ equiv}^{-1}$
360. In a concentration cell :
- a) Two electrodes are of different elements
 b) Two electrolytic solutions of the same electrolyte but having different concentrations are used
 c) Electrolyte of one strength but electrodes of two different concentrations are used
 d) Both (b) and (c)
361. Using the following data, for the electrode potentials calculate ΔG° , in kJ, for the indicated reaction
- $5Ce^{4+}(aq) + Mn^{2+}(aq) + 4H_2O(l) \rightarrow 5Ce^{3+}(aq) + MnO_4^-(aq) + 8H^+(aq)$
 $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l) E^\circ = +1.51 \text{ V}$
 $Ce^{4+}(aq) + e^- \rightarrow Ce^{3+}(aq) E^\circ = +1.61 \text{ V}$
- a) -36.24 b) -48.25 c) -31.54 d) -19.65
362. During electrolysis of an aqueous solution of Cu^{2+} sulphate, 0.635 g of copper was deposited at cathode. The amount of electricity consumed in coulomb is :
- a) 1930 b) 3860 c) 9650 d) 4825
363. Each of the three metals X, Y and Z were put in turn into aqueous solution of the other two. $X + \text{Salt of } Y \text{ (or } Z) = Y \text{ (or } Z) + \text{Salt of } X$. Which observation is probably incorrect?
- a) $Y + \text{Salt of } X = \text{No action observed}$
 b) $Y + \text{Salt of } Z = \text{Nothing can be decided}$
 c) $Z + \text{Salt of } X = X + \text{Salt of } Z$
 d) $Z + \text{Salt of } X = \text{No action observed}$
364. During the charging of lead storage battery, the reaction at anode is represented by :
- a) $Pb^{2+} + SO_4^{2-} \rightarrow PbSO_4$
 b) $PbSO_4 + H_2O \rightarrow PbO_2 + SO_4^{2-} + 2H^+$
 c) $Pb \rightarrow Pb^{2+} + 2e^-$
 d) $Pb^{2+} + 2e^- \rightarrow Pb$
365. Which of the formula given below is correct?
- a) $\kappa = \frac{1}{R} \times \frac{1}{a}$ b) $\kappa = C \times \frac{1}{a}$ c) $\Lambda_{eq} = \kappa \times V_{in \text{ mL}}$ d) All of these
366. The number of faraday required to generate 1 g-atom of Mg from $MgCl_2$ is :
- a) 1 b) 2 c) 3 d) 4
367. During the electrolysis of molten NaCl solution, 230 g of sodium metal is deposited on the cathode, then how many moles of chlorine will be obtained at anode?
- a) 10.0 b) 5.0 c) 35.5 d) 17.0
368. 1.05 g of lead ore containing impurity of Ag was dissolved in HNO_3 and the volume was made 350 mL. An Ag electrode was dipped in the solution. $Pt(H_2) | H^+(1 \text{ M}) || Ag^+ | Ag$
 The E_{cell} is 0.503 V at 298 K. The percent of lead in the ore is ($E^\circ_{Ag^+|Ag} = 0.80 \text{ V}$)
- a) 0.033% b) 0.050% c) 0.066% d) 0.13%
369. The equivalent conductivity of 0.05 N solution of a monobasic acid is $15.8 \text{ mho cm}^2 \text{ eq}^{-1}$. If equivalent conductivity of the acid at infinite dilution is $350 \text{ mho cm}^2 \text{ eq}^{-1}$ at the same temperature. What is its degree of dissociation?
- a) 0.04514 b) 0.4514 c) 4.514 d) 0.004514
370. What is the potential of the cell containing two hydrogen electrodes as represented ahead,



- a) - 0.295 V b) - 0.0591 V c) 0.295 V d) 0.0591 V

371. The standard emf for the given cell reaction, $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Cu} + \text{Zn}^{2+}$ is 1.10 V at 25°C. The emf for the cell reaction, when 0.1 M Cu^{2+} and 0.1 M Zn^{2+} solutions are used, at 25°C, is

- a) 1.10 V b) -1.10 V c) 2.20 V d) -2.20 V

372. Four colourless salt solutions are placed in separate test tubes and a strip of copper is placed in each. Which solution finally turns blue?

- a) $\text{Pb}(\text{NO}_3)_2$ b) $\text{Zn}(\text{NO}_3)_2$ c) AgNO_3 d) $\text{Cd}(\text{NO}_3)_2$

373. The same amount of electricity was passed through two separate electrolytic cells containing solutions of nickel nitrate and chromium nitrate respectively. If 0.3 g of nickel was deposited in the first cell, the amount of chromium deposited is :

(at. wt. Ni = 59, Cr = 52)

- a) 0.1 g b) 0.17 g c) 0.3 g d) 0.6 g

374. The molar conductivities of KCl, NaCl and KNO_3 are 152, 128 and $111 \text{ S cm}^2 \text{ mol}^{-1}$ respectively. What is the molar conductivity of NaNO_3 ?

- a) $101 \text{ S cm}^2 \text{ mol}^{-1}$ b) $87 \text{ S cm}^2 \text{ mol}^{-1}$ c) $-101 \text{ S cm}^2 \text{ mol}^{-1}$ d) $-391 \text{ S cm}^2 \text{ mol}^{-1}$

375. The degree of ionisation of weak electrolytes is influenced by :

- a) Temperature
b) Concentration of electrolyte
c) Nature of solvent
d) All of the above

376. At 25°C temperature, the cell potential of a given electrochemical cell is 1.92 V. Find the value of x.



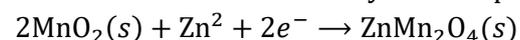
$$E^\circ \text{ Mg} / \text{Mg}^{2+}(\text{aq}) = 2.37 \text{ V} \quad E^\circ \text{ Fe} / \text{Fe}^{2+}(\text{aq}) = 0.45 \text{ V}$$

- a) $x = 0.01 \text{ M}$ b) $x < 0.01 \text{ M}$
c) $x > 0.01 \text{ M}$ d) x cannot be predicted

377. The corrosion of iron object is favoured by :

- a) Presence of H^+ ion
b) Presence of moisture in air
c) Presence of impurities in iron object
d) All of the above

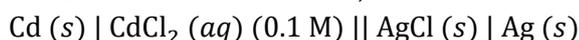
378. The cathodic reaction of a dry cell is represented by



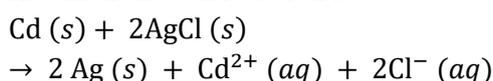
If, there are 8 g of MnO_2 in the cathodic compartment then the time for which the dry cell will continue to give a current of 2 mA is

- a) 25.675 day b) 51.35 day c) 12.8 day d) 6.423 day

379. The standard emf of the cell,



In which the cell reaction is



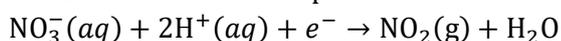
is 0.6915 V at 0°C and 0.6753 V at 25°C. The enthalpy change of the reaction at 25°C is

- a) -176 kJ b) -234.7 kJ c) +123.5 kJ d) -167.26 kJ

380. The factor which is not affecting the conductivity of any solution is

- a) Temperature b) Dilution c) Nature of electrolyte d) None of these

381. The standard reduction potential for the half-cell having reaction



is 0.78 V. What will be the reduction potential of the half-cell in a neutral solution?

- a) 0.78 V b) 0.89 V c) 0.36 V d) 0.59 V

382. Two different electrolytic cells filled with molten $\text{Cu}(\text{NO}_3)_2$ and molten $\text{Al}(\text{NO}_3)_3$ respectively are connected in series. When electricity is passed 2.7 g Al is deposited on electrode. Calculate the weight of Cu deposited on cathode.

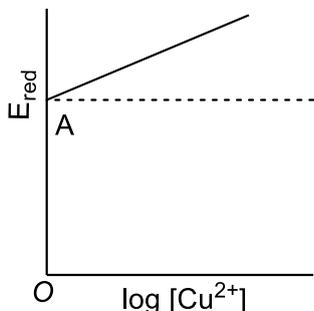
[Cu = 63.5; Al = 27.0 g mol⁻¹]

- a) 190.5 g b) 9.525 g c) 63.5 g d) 31.75 g

383. 1 volt coulomb is :

- a) Equal to 1 joule b) Equal to 10⁷ erg c) An unit of energy d) All of these

384. $\text{Cu}^{2+} + 2e \rightarrow \text{Cu}$; $\log[\text{Cu}^{2+}]$ vs. E_{red} graph is of the type as shown in figure where $OA = 0.34$ V, then electrode potential of the half cell of Cu | $\text{Cu}^{2+}(0.1 \text{ M})$ will be :



- a) $-0.34 + \frac{0.0591}{2} \text{ V}$ b) $0.34 + 0.0591 \text{ V}$ c) 0.34 V d) None of these

385. If ϕ denotes standard reduction potential, which is true:

- a) $E_{\text{cell}}^\circ = \phi_R - \phi_L$ b) $E_{\text{cell}}^\circ = \phi_L + \phi_R$ c) $E_{\text{cell}}^\circ = \phi_L - \phi_R$ d) $E_{\text{cell}}^\circ = (\phi_L + \phi_R)$

386. A substance that will reduce Ag^+ to Ag but will not reduce Ni^{2+} to Ni is :

- a) Zn b) Pb c) Mg d) Al

387. The correct order of the mobility of the alkali metal ions in aqueous solution is :

- a) $\text{K}^+ > \text{Rb}^+ > \text{Na}^+ > \text{Li}^+$
 b) $\text{Rb}^+ > \text{K}^+ > \text{Na}^+ > \text{Li}^+$
 c) $\text{Li}^+ > \text{Na}^+ > \text{K}^+ > \text{Rb}^+$
 d) $\text{Na}^+ > \text{K}^+ > \text{Rb}^+ > \text{Li}^+$

388. Calculate the volume of H_2 gas at NTP obtained by passing 4 A through acidified H_2O for 30 min is

- a) 0.0836 L b) 0.0432 L c) 0.1672 L d) 0.836 L

389. Resistance of a conductivity cell filled with a solution of an electrolyte of concentration 0.1 M is 100 Ω . The conductivity of this solution is 1.29 S m⁻¹. Resistance of the same cell when filled with 0.2 M of the same solution is 520 Ω . The molar conductivity of 0.02 M solution of the electrolyte will be

- a) $124 \times 10^{-4} \text{ S m}^2 \text{ mol}^{-1}$ b) $1240 \times 10^{-4} \text{ S m}^2 \text{ mol}^{-1}$
 c) $1.24 \times 10^{-4} \text{ S m}^2 \text{ mol}^{-1}$ d) $12.4 \times 10^{-4} \text{ S m}^2 \text{ mol}^{-1}$

390. Ionic mobility of Ag^+ at infinite dilution is :

($\lambda_{\text{Ag}^+}^0 = 5 \times 10^{-3} \text{ S m}^2 \text{ eq}^{-1}$)

- a) 5.2×10^{-8} b) 2.4×10^{-8} c) 1.52×10^{-8} d) 8.25×10^{-8}

391. The number of electrons passing per second through a cross-section of copper wire carrying 10⁻⁶ ampere :

- a) 6.2×10^{23} b) 6.2×10^{12} c) 6.2×10^{10} d) None of these

392. The amount of substance deposited by the passage of 1 A of current for 1 s is equal to

- a) Equivalent mass b) Molecular mass
 c) Electrochemical equivalent d) Specific equivalent

393. 9.65 C electric current is passed through fused anhydrous MgCl_2 . The magnesium metal thus obtained is completely converted into a Grignard reagent. The number of moles of Grignard reagent obtained is

- a) 5×10^{-4} b) 1×10^{-4} c) 5×10^{-5} d) 1×10^{-5}

394. Which one is correct relation :

- a) $\Delta S = \left(\frac{\partial E}{\partial T}\right)_P \times nF$

$$b) \left(\frac{\partial E}{\partial T}\right)_p = \frac{\Delta G - \Delta H}{T}$$

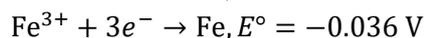
$$c) \left(\frac{\partial E}{\partial T}\right)_p = \frac{\partial(\Delta S)}{\partial T}$$

$$d) -\Delta S = \left(\frac{\partial E}{\partial T}\right)_p \times nF$$

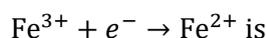
395. A current is passed through two voltmeters connected in series. The first voltmeter contains $XSO_4(aq)$ while the second voltmeter contains $Y_2SO_4(aq)$. The relative atomic masses of X and Y are in the ratio of 2 : 1. The ratio of the mass of X liberated to the mass of Y liberated is :

- a) 1 : 1 b) 1 : 2 c) 2 : 1 d) None of these

396. Given, standard electrode potentials

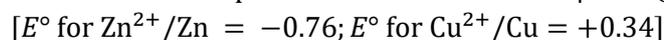


The standard electrode potential (E°) for



- a) + 0.771 V b) - 0.771 V c) + 0.417 V d) - 0.417 V

397. The standard cell potential for the cell is : $Zn | Zn^{2+}(1M) || Cu^{2+}(1M) | Cu$



- a) $-0.76 + 0.34 = -0.42 \text{ V}$
 b) $-0.34 - (-0.76) = +0.42 \text{ V}$
 c) $0.34 - (-0.76) = +1.10 \text{ V}$
 d) $-0.76 - (+0.34) = -1.10 \text{ V}$

398. The speed of migration of Ag^+ ion and NO_3^- ion are $0.00057 \text{ cm sec}^{-1}$ and $0.00063 \text{ cm sec}^{-1}$ at infinite dilution. The equivalent conductivity of $AgNO_3$ at infinite dilution is:

- a) 140.2 b) 130.1 c) 120.8 d) 115.8

399. In electrochemical corrosion of metals, the metal undergoing corrosion :

- a) Acts as anode b) Acts as cathode c) Undergoes reduction d) None of these

400. Which does not get oxidised by bromine water?

- a) Fe^{2+} to Fe^{3+} b) Cu^+ to Cu^{2+} c) Mn^{2+} to MnO_4^- d) Sn^{2+} to Sn^{4+}

401. 3 faraday of electricity is passed through molten Al_2O_3 , aqueous solution of $CuSO_4$ and molten $NaCl$ taken in three different electrolytic cells. The amount of Al , Cu and Na deposited at the cathodes will be in the ratio of :

- a) 1 mole : 2 mole : 3 mole
 b) 3 mole : 2 mole : 1 mole
 c) 1 mole : 1.5 mole : 3 mole
 d) 1.5 mole : 2 mole : 3 mole

402. In Galvanic cell, the electrons flow from

- a) Anode to cathode through the solution b) Cathode to anode through the solution
 c) Anode to cathode through the external circuit d) Cathode to anode through the external circuit

403. Which of the following reactions is correct for a given electrochemical cell at 25 °C?



- a) $2Br^-(aq) + Cl_2(g) \rightarrow 2Cl^-(aq) + Br_2(g)$ b) $Br_2(g) + 2Cl^-(aq) \rightarrow 2Br^-(aq) + Cl_2(g)$
 c) $Br_2(g) + Cl_2(g) \rightarrow 2Br^-(aq) + 2Cl^-(aq)$ d) $2Br^-(aq) + 2Cl^-(aq) \rightarrow Br_2(g) + Cl_2(g)$

404. Which gains electrons more easily?

- a) H^+ b) Na^+ c) K^+ d) Mg^{2+}

405. Two electrolytic cells, one containing acidified ferrous chloride and another acidified ferric chloride are connected in series. The ratio of iron deposited at cathodes in the two cells when electricity is passed through the cells will be :

- a) 3 : 1 b) 2 : 1 c) 1 : 1 d) 3 : 2

406. Limiting molar conductivity of NH_4OH , i.e., $\Lambda_m(NH_4OH)$ is equal to :

- a) $\Delta_m(\text{NH}_4\text{OH}) + \Delta_m(\text{NH}_4\text{Cl}) - \Delta_m(\text{HCl})$
 b) $\Delta_m(\text{NH}_4\text{Cl}) + \Delta_m(\text{NaOH}) - \Delta_m(\text{NaCl})$
 c) $\Delta_m(\text{NH}_4\text{Cl}) + \Delta_m(\text{NaCl}) - \Delta_m(\text{NaOH})$
 d) $\Delta_m(\text{NaOH}) + \Delta_m(\text{NaCl}) - \Delta_m(\text{NH}_4\text{Cl})$

407. Given :

- i) $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}, E^\circ = 0.337 \text{ V}$
 ii) $\text{Cu}^{2+} + e^- \rightarrow \text{Cu}^+, E^\circ = 0.153 \text{ V}$
 Electrode potential, E° for the reaction,
 $\text{Cu}^+ + e^- \rightarrow \text{Cu}$, will be :

- a) 0.38 V b) 0.52 V c) 0.90 V d) 0.30 V

408. The reaction taking place at anode when an aqueous solution of CuSO_4 is electrolysed using inert Pt electrode :

- a) $2\text{SO}_4^{2-} \rightarrow \text{S}_2\text{O}_8^{2-} + 2e^-$
 b) $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$
 c) $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4e^-$
 d) $2\text{H}^+ + 2e^- \rightarrow \text{H}_2$

409. Deduce from the following E° values of half cells, what combination of two half cells would result in a cell with the largest potential?

- i) $\text{A}^{3-} \rightarrow \text{A}^{2-} + e^-; E^\circ = 1.5 \text{ V}$
 ii) $\text{B}^{2+} + e^- \rightarrow \text{B}^+; E^\circ = -2.1 \text{ V}$
 iii) $\text{C}^{2+} + e^- \rightarrow \text{C}^+; E^\circ = +0.5 \text{ V}$
 iv) $\text{D} \rightarrow \text{D}^{2+} + 2e^-; E^\circ = -1.5 \text{ V}$

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

410. An ion is reduced to the element when it absorbs 6×10^{20} electrons. The number of equivalents of the ion is :

- a) 0.10 b) 0.01 c) 0.001 d) 0.0001

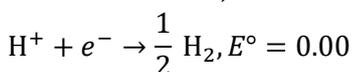
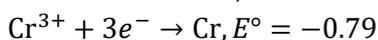
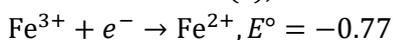
411. The standard e.m.f. of a galvanic cell can be calculated from :

- a) The size of the electrode
 b) The pH of the solution
 c) The amount of metal in the anode
 d) The E° values of the two half cells

412. The charge in coulomb on Cu^{2+} ion is :

- a) 3.2×10^{-19} b) 2.3×10^{-12} c) 0.23×10^{-19} d) 0.32×10^{-19}

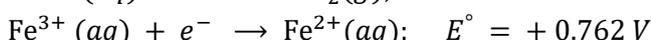
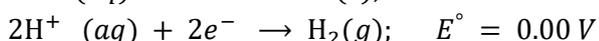
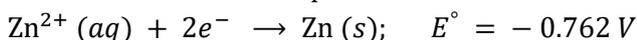
413. $\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}(s), E^\circ = -0.76$



Strongest reducing agent is

- a) H_2 b) Zn c) Fe^{2+} d) Cr

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



The strongest reducing agent is

- a) Zn (s) b) Cr (s) c) $\text{H}_2(\text{g})$ d) $\text{Fe}^{2+}(\text{aq})$

415. Strong electrolytes are those which

- a) Conduct electricity b) Dissolve readily in water

427. the amount of sodium deposited by 5 ampere current for 10 minute from fused NaCl is :
 a) 0.715 g b) 71.5 g c) 5.17 g d) 0.517 g
428. Which of the following reactions is used to make a fuel cell?
 a) $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$
 b) $2\text{Fe}(\text{s}) + \text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) \rightarrow 2\text{Fe}^{2+}(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$
 c) $\text{Pb}(\text{s}) + \text{PbO}_2(\text{s}) + 2\text{H}_2\text{SO}_4(\text{aq}) \rightarrow 2\text{PbSO}_4(\text{s}) + \text{H}_2\text{O}(\text{l})$
 d) $\text{Cd}(\text{s}) + 2\text{Ni}(\text{OH})_3(\text{s}) \rightarrow \text{CdO}(\text{s}) + 2\text{Ni}(\text{OH})_2 + 2\text{H}_2\text{O}(\text{l})$
429. When 1 F of electricity is passed through acidulated water, O_2 evolved is
 a) 11.2 dm^3 b) 5.6 dm^3 c) 22.4 dm^3 d) 1.0 dm^3
430. The Gibbs energy for the decomposition of Al_2O_3 at 500°C is as follows :
 $\frac{2}{3}\text{Al}_2\text{O}_3 \rightarrow \frac{4}{3}\text{Al} + \text{O}_2;$
 $\Delta_r G = +960 \text{ kJ mol}^{-1}.$
 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
431. Which one will liberate Br_2 from KBr?
 a) HI b) I_2 c) Cl_2 d) SO_2
432. In a galvanic cell, the electrons flow from
 a) Anode to cathode through the external circuit b) Anode to cathode through the solution
 c) Cathode to anode through the external circuit d) Cathode to anode through the solution
433. The value of molar conductance of HCl is greater than that of NaCl at a particular temperature and dilution because :
 a) mol. wt. of HCl < mol. wt. of NaCl
 b) $u_{\text{H}^+} > u_{\text{Na}^+}$ (u is speed of ion)
 c) HCl is acid
 d) Ionisation of HCl is more than NaCl
434. Maximum number of mole of oxygen gas that can be obtained by the electrolytic decomposition of 90 g of water will be:
 a) 1 b) 2.5 c) 5 d) 9
435. More electropositive elements have :
 a) Positive reduction potential
 b) Tendency to gain electrons
 c) Negative reduction potential
 d) Negative oxidation potential
436. 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) $1.3 \times 10^4 \text{ g}$ b) $9.0 \times 10^3 \text{ g}$ c) $8.05 \times 10^4 \text{ g}$ d) $2.4 \times 10^5 \text{ g}$
437. The reaction,
 $\frac{1}{2}\text{H}_2(\text{g}) + \text{AgCl}(\text{s}) \rightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{Ag}(\text{s})$
 Occurs in the galvanic cell
 a) Pt/ $\text{H}_2(\text{g})$ | KCl (sol) || AgCl(s) | Ag b) Pt/ $\text{H}_2(\text{g})$ | HCl(sol) || AgNO₃ (sol) | Ag
 c) Pt/ $\text{H}_2(\text{g})$ | HCl(sol) || AgCl(s) | Ag d) Ag/AgCl(s) | KCl(sol) || AgNO₃ | Ag
438. Which of the following ions can be replaced by H^+ ions when H_2 gas is bubbled through the solutions containing these ions?
 a) Li^+ b) Ba^{2+} c) Cu^{2+} d) Be^{2+}
439. The cell reaction, $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$ is best represented by :
 a) Cu/ Cu^{2+} || Zn^{2+} / Zn b) Zn/ Zn^{2+} || Cu^{2+} / Cu c) Cu^{2+} / Cu || Zn/ Zn^{2+} d) Pt/ Zn^{2+} || Pt/ Cu^{2+}
440. Ionic mobility (u^∞) of an ion at infinite dilution is related to its ionic conductance (λ_∞) by :

- a) $\lambda_{\infty} = u_{\infty} \times \text{Faraday}$ b) $u_{\infty} = \lambda_{\infty} \times \text{Faraday}$ c) $\text{Faraday} = u_{\infty} \times \lambda_{\infty}$ d) None of these
441. Coulomb is the quantity of current defined as :
- One ampere of current passing for 1 sec
 - One which deposits 0.001118 g of Ag on cathode
 - One which deposits electrochemical equivalence of metal
 - All of the above
442. The standard electrode potential is measured by
- Electrometer
 - Voltmeter
 - Pyrometer
 - Galvanometer
443. Chlorine gas is passed into a solution containing KF, KI, and KBr and CHCl_3 is added. The initial colour in CHCl_3 layer is :
- Violet due to formation of I_2
 - Orange due to formation of Br_2
 - Colourless due to formation of F_2
 - No colour change due to no reaction
444. On passing 3 A of electricity for 50 min, 1.8 g metal deposits. The equivalent mass of metal is
- 9.3
 - 19.3
 - 38.3
 - 39.9
445. How many atoms of calcium will be deposited from a solution of CaCl_2 by a current 0.25 mA flowing for 60 s?
- 4.68×10^{18}
 - 2.34×10^{18}
 - 1.24×10^{18}
 - 0.46×10^{18}
446. If 'F' is faraday and 'N' is Avogadro number, then charge of electron can be expressed as
- $F \times N$
 - $\frac{F}{N}$
 - $\frac{N}{F}$
 - $F^2 N$
447. By how much is the oxidizing power of $\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}$ couple decreased if the H^+ concentration is decreased from 1 M to 10^{-3} M at 25°C ?
- 0.207 V
 - 0.414 V
 - 0.001 V
 - 0.287 V
448. Which process involves corrosion?
- Brown deposits on iron articles
 - Green deposits on battery terminals
 - Black deposits on silver coin
 - All of the above
449. The electric conduction of a salt solution in water depends on the
- Size of its molecules
 - Shape of its molecules
 - Size of solvent molecules
 - Extent of its ionization
450. The electrode potentials for
- $$\text{Cu}^{2+}(\text{aq}) + e^- \rightarrow \text{Cu}^+(\text{aq})$$
- and
- $$\text{Cu}^+(\text{aq}) + e^- \rightarrow \text{Cu}(\text{s})$$
- are + 0.15 V and + 0.50 V respectively. The value of $E_{\text{Cu}^{2+}/\text{Cu}}^\circ$ will be :
- 0.150 V
 - 0.500 V
 - 0.325 V
 - 0.650 V
451. By diluting a weak electrolyte, specific conductivity (K_c) and equivalent conductivity (λ_c) change as
- Both increase
 - K_c increases, λ_c decreases
 - K_c decreases, λ_c increases
 - Both decrease
452. The cell reaction for the given cell is :
- $$\text{Pt}(\text{H}_2) \mid \text{pH} = 2 \mid \mid \text{pH} = 3 \mid \text{Pt}(\text{H}_2)$$
- $P_1 = 1 \text{ atm}$ $P_2 = 1 \text{ atm}$
- Spontaneous
 - Non-spontaneous
 - In equilibrium
 - Either of these
453. If the molar conductance value of Ca^{2+} and Cl^- at infinite dilution are respectively $118.88 \times 10^{-4} \text{ m}^2 \text{ mho mol}^{-1}$ and $77.33 \times 10^{-4} \text{ m}^2 \text{ mho mol}^{-1}$ then that of CaCl_2 is (in $\text{m}^2 \text{ mho mol}^{-1}$)
- 118.88×10^{-4}
 - 154.66×10^{-4}
 - 273.54×10^{-4}
 - 196.21×10^{-4}
454. During electrolysis, the species discharged at cathode are
- Anion
 - Cation
 - Ions
 - All of these

455. In the electrolysis of which solution, OH^- ions are discharged in preference to Cl^- ions?
 a) Dilute NaCl b) Very dilute NaCl c) Fused NaCl d) Solid NaCl
456. A cell constructed by coupling a standard copper electrode and a standard magnesium electrode has emf of 2.7 V. If the standard reduction potential of copper electrode is +0.34 V then that of the magnesium electrode is
 a) + 2.36 V b) - 2.36 V c) + 3.26 V d) - 3.26 V
457. Variation of equivalent conductivity with concentration of strong electrolyte is given by Hückel-Onsager equation expressed as :
 a) $\Lambda_M = \Lambda^\infty - b\sqrt{c}$ b) $\Lambda_\infty = \Lambda M - b\sqrt{c}$ c) $\Lambda_M = b\sqrt{c} - \Lambda^\infty$ d) None of these
458. An electric current is passed through silver nitrate solution using silver electrodes. 10.79 g of silver was found to be deposited on the cathode. If the same amount of electricity is passed through copper sulphate solution using copper electrodes, the weight of copper deposited on the cathode is
 a) 1.6 g b) 2.3 g c) 3.2 g d) 6.4 g
459. The amount of energy expended during the passage of one ampere current for 100 second under a potential of 115 V is :
 a) 20 kJ b) 11.5 kJ c) 115 kJ d) 0.115 kJ
460. If a strip of copper metal is placed in a solution of ferrous sulphate :
 a) Copper will precipitate out
 b) Iron will precipitate out
 c) Both copper and iron will be dissolved
 d) No reaction will take place
461. The process of zinc-plating on iron sheet is known as
 a) Annealing b) Roasting c) Galvanisation d) Smelting
462. For the following cell with hydrogen electrodes at two different pressures p_1 and p_2
 $\text{Pt}(\text{H}_2) | \text{H}^+ (aq) | \text{Pt}(\text{H}_2)$
 $p_1 \quad 1 \text{ M} \quad p_2$
 emf is given by
 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) $\frac{RT}{2F} \log_e \frac{p_2}{p_1}$
463. During the electrolysis of a solution of AgNO_3 , 9650 C of charge is passed through the electroplating bath. The mass of silver deposited at the cathode will be
 a) 108 g b) 10.8 g c) 1.08 g d) 18.10 g
464. What is the time (in sec) required for depositing all the silver present in 125mL of 1 M AgNO_3 solution by passing a current of 241.25 A ? ($1F = 96500 \text{ C}$)
 a) 10 b) 50 c) 1000 d) 100
465. For the redox reaction,
 $\text{Zn}(s) + \text{Cu}^{2+}(0.1 \text{ M}) \rightarrow \text{Zn}^{2+}(1 \text{ M}) + \text{Cu}(s)$
 taking place in a cell, E_{cell}° is 1.10 V. E_{cell}° for the cell will be
 $\left(2.303 \frac{RT}{F} = 0.0591\right)$
 a) 2.14 V b) 1.80 V c) 1.07 V d) 0.82 V
466. The limiting molar conductivities Λ° for NaCl, KBr and KCl are 126, 152 and 150 $\text{S cm}^2 \text{ mol}^{-1}$ respectively. The Λ° for NaBr is
 a) 128 $\text{S cm}^2 \text{ mol}^{-1}$ b) 248 $\text{S cm}^2 \text{ mol}^{-1}$ c) 328 $\text{S cm}^2 \text{ mol}^{-1}$ d) 348 $\text{S cm}^2 \text{ mol}^{-1}$
467. The emf of the cell,
 $\text{Ag} | \text{Ag}^+ (0.1 \text{ M}) || \text{Ag}^+ (1 \text{ M}) | \text{Ag}$ at 298 K is
 a) 0.0059 V b) 0.059 V c) 5.9 V d) 0.59 V
468. A solution of sodium sulphate in water is electrolysed using inert electrodes. The products at the cathode and anode are respectively

- a) H_2, O_2 b) O_2, H_2 c) O_2, Na d) O_2, SO_2
469. The standard electrode potential for the change ;
 $\text{Sn}(s) + 2\text{Fe}^{3+}(aq) \rightarrow 2\text{Fe}^{2+}(aq) + \text{Sn}^{2+}(aq)$ is :
 (Given $E^\circ_{\text{Fe}^{3+}/\text{Fe}^{2+}} = 0.77 \text{ V}$ and $E^\circ_{\text{Sn}^{2+}/\text{Sn}} = -0.14 \text{ V}$)
 a) 0.63 V b) 1.40 V c) 0.91 V d) 1.68 V
470. Hydrogen-oxygen fuel cells are used in spacecraft to supply :
 a) Power for heat and light
 b) Power for pressure
 c) Oxygen
 d) None of the above
471. The resistance of 0.01 N solution of an electrolyte was found to be 210 ohm at 298 K. Its conductance is :
 a) 4.76×10^{-3} mho b) 4.76 mho c) 210 mho d) None of these
472. The amount of silver deposited on passing 2 F of electricity through aqueous solution of AgNO_3 is
 a) 54 g b) 108 g c) 216 g d) 324 g
473. Cell constant of a conductivity cell is usually derived by using a solution of :
 a) KCl b) NaCl c) NH_4Cl d) LiCl
474. Four successive members of the first series of the transition metals are listed below. For which one of them the standard potential ($E^\circ_{M^{2+}/M}$) value has a positive sign?
 a) Co (Z = 27) b) Ni (Z = 28) c) Cu (Z = 29) d) Fe (Z = 26)
475. When same quantity of electricity is passed through aqueous AgNO_3 and H_2SO_4 solutions connected in series, 5.04×10^{-2} g of H_2 is liberated. What is the mass of silver (in grams) deposited? (Eq. wts. of hydrogen = 1.008, silver = 108)
 a) 54 b) 0.54 c) 5.4 d) 10.8
476. The term infinite dilution refers when :
 a) $\alpha \rightarrow 1$, for weak electrolytes
 b) An electrolyte is 100% dissociated
 c) All interionic effects disappears
 d) All of the above
477. In the problem 15, the limiting mobility of K^+ ion is :
 a) 6.1×10^{-4} b) 6.67×10^{-4} c) 7.1×10^{-4} d) 7.67×10^{-4}
478. The standard reduction potential at 290 K for the following half reactions are,
 (i) $\text{Zn}^{2+} + 2e \rightarrow \text{Zn}(s); E^\circ = -0.762 \text{ V}$
 (ii) $\text{Cr}^{3+} + 3e \rightarrow \text{Cr}(s); E^\circ = -0.740 \text{ V}$
 (iii) $2\text{H}^+ + 2e \rightarrow \text{H}_2(g); E^\circ = -0.000 \text{ V}$
 (iv) $\text{Fe}^{3+} + e \rightarrow \text{Fe}^{2+}; E^\circ = +0.77 \text{ V}$
 Which is the strongest reducing agent?
 a) Zn b) Cr c) Fe^{2+} d) H_2
479. Daniel cell, anode and cathode are respectively
 a) $\text{Zn} | \text{Zn}^{2+}$ and $\text{Cu}^{2+} | \text{Cu}$ b) $\text{Cu} | \text{Cu}^{2+}$ and $\text{Zn}^{2+} | \text{Zn}$ c) $\text{Fe} | \text{Fe}^{2+}$ and $\text{Cu}^{2+} | \text{Cu}$ d) $\text{Cu} | \text{Cu}^{2+}$ and $\text{Fe}^{2+} | \text{Fe}$
480. Iron sheets are galvanized to :
 a) Prevent action of O_2 and H^+ on Fe
 b) Prevent oxidation of Fe
 c) Prevent rusting
 d) All of the above
481. The conductance of all the ions present in a solution containing 1 g equivalent in it is known as :
 a) Conductivity
 b) Equivalent conductivity
 c) Molecular conductivity
 d) None of the above

482. For the reduction of silver ions with copper metal, the standard cell potential is 0.46 V at 25°C. The value of standard Gibbs energy ΔG° will be :

- a) - 89.0 kJ b) - 89.0 J c) - 44.5 kJ d) - 98.0 kJ

483. E° of an electrode is :

- a) Extensive property b) Constitutive property c) Colligative property d) Intensive property

484. The one which decreases with dilution is

- a) Molar conductance b) Conductance
c) Specific Conductance d) Equivalent conductance

485. Which of the following metal can replace zinc from $ZnSO_4$ solution?

- a) Cu b) Hg c) Fe d) Al

486. Several blocks of magnesium are fixed to the bottom of Aship to

- a) Keep away the sharks b) Make the ship lighter
c) Prevent action of water and salt d) Prevent puncturing by under- sea Rocks

487. $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$; $E^\circ = 1.51 V$

$MnO_4 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$; $E^\circ = 1.23 V$

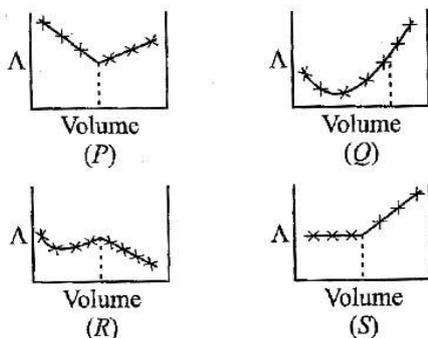
$E^\circ_{MnO_4^-|MnO_2}$ is

- a) 1.70 V b) 0.91 V c) 1.37 V d) 0.548 V

488. In a salt bridge, KCl is used because :

- a) It is an electrolyte
b) It is good conductor of electricity
c) The transport number of K^+ and Cl^- ions are nearly same or both have same ionic mobility.
d) It is ionic compound.

489. $AgNO_3 (aq)$ was added to an aqueous KCl solution gradually and the conductivity of the solution was measured. The plot of conductance (Λ) versus the volume of $AgNO_3$ is



- a) (P) b) (Q) c) (R) d) (S)

490. For the electrochemical cell, $M | M^+ || X^- | X$, $E^\circ(M^+ | M) = 0.44 V$ and $E^\circ(X | X^-) = 0.33 V$. From this data one can deduce that

- a) $E_{cell} = 0.77 V$
b) $-0.77 V$
c) $M^+ + X^- \rightarrow M + X$ is the spontaneous reaction
d) $M + X \rightarrow M^+ + X^-$ is the spontaneous reaction

491. The specific conductance (κ) of an electrolyte of 0.1 N concentration is related to equivalent conductance (Λ) by the following formula

- a) $\Lambda = \kappa$ b) $\Lambda = 10\kappa$ c) $\Lambda = 100\kappa$ d) $\Lambda = 10000\kappa$

492. Which is the correct representation for Nernst equation?

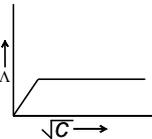
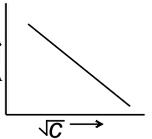
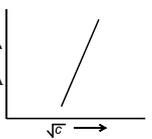
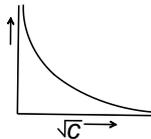
- a) $E_{RP} = E^\circ_{RP} + \frac{0.059}{n} \log \frac{[\text{oxidant}]}{[\text{reductant}]}$
b) $E_{OP} = E^\circ_{OP} - \frac{0.059}{n} \log \frac{[\text{oxidant}]}{[\text{reductant}]}$
c) $E_{OP} = E^\circ_{OP} + \frac{0.059}{n} \log \frac{[\text{reductant}]}{[\text{oxidant}]}$

- d) All of the above
493. The number of electrons required to deposit 1 g atom of Al (at. wt. = 27) from a solution of AlCl_3 are :
 a) 1 N b) 2 N c) 3 N d) 4 N
494. The standard reduction potential of some electrodes are,
 $E^\circ(\text{K}^+/\text{K}) = -2.9 \text{ V}$,
 $E^\circ(\text{Zn}^{2+}/\text{Zn}) = -0.76 \text{ V}$, $E^\circ(\text{H}^+/\text{H}_2) = -0.00 \text{ V}$,
 $E^\circ(\text{Cu}^{2+}/\text{Cu}) = +0.34 \text{ V}$
 The Strongest oxidant is :
 a) Copper b) Zinc c) Hydrogen d) Cu^{2+}
495. 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
496. The weight ratio of Mg and Al deposited during the passage of same current through their molten salts :
 a) 12 : 9 b) 9 : 12 c) 6 : 2 d) 2 : 3
497. $2\text{Fe}^{3+} + 3\text{I}^- \rightleftharpoons 2\text{Fe}^{2+} + \text{I}_3^-$
 The standard reduction potentials in acidic conditions are 0.77 and 0.54 V respectively for $\text{Fe}^{3+}/\text{Fe}^{2+}$ and I_3^-/I^- couples. The equilibrium constant for the reaction is
 a) 6.26×10^{-7} b) 5.33×10^{-4} c) 6.26×10^7 d) 5.33×10^4
498. In a cell that utilizes the reaction
 $\text{Zn}(\text{s}) + 2\text{H}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{H}_2(\text{g})$
 addition of H_2SO_4 to cathode compartment will
 a) Lower the E and shift equilibrium to the right
 b) Lower the E and shift equilibrium to the left
 c) Increase the E and shift equilibrium to the right
 d) Increase the E and shift equilibrium to the left
499. On passing electricity through dilute H_2SO_4 solution the amount of substance liberated at the cathode and anode are in the ratio :
 a) 1 : 8 b) 8 : 1 c) 16 : 1 d) 1 : 16
500. The increase in equivalent conductivity of a strong electrolytic solution with dilution is attributed to :
 a) Increase in number of ions per unit volume
 b) Increase in molecular attraction
 c) Increase in degree of dissociation
 d) Increase in ionic mobility
501. The cell, $\text{Zn} | \text{Zn}^{2+} (1 \text{ M}) || \text{Cu}^{2+} (1 \text{ M}) | \text{Cu}$
 $(E^\circ_{\text{cell}} = 1.10 \text{ V})$, was allowed to be completely discharged at 298 K. The relative concentration of Zn^{2+} to Cu^{2+} $\left(\frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}\right)$ is
 a) Antilog (24.08) b) Antilog (37.3) c) $10^{37.3}$ d) 9.65×10^4
502. The algebraic sum of potentials of two electrodes of a galvanic cell is called :
 a) Potential difference b) Ionic difference c) e.m.f. d) Electrode difference
503. The standard oxidation potentials, E° for the half reactions are ;
 $\text{Zn} \rightarrow \text{Zn}^{2+} + 2e$; $E^\circ = +0.76 \text{ V}$
 $\text{Ag} \rightarrow \text{Ag}^+ + e$; $E^\circ = -0.77 \text{ V}$
 The standard e. m. f. of the cell,
 $\text{Ag}^+ + \text{Zn} \rightarrow \text{Zn}^{2+} + \text{Ag}$ is :
 a) + 1.53 V b) - 1.53 V c) -0.01 V d) + 0.01 V
504. Rust is a mixture of :
 a) FeO and $\text{Fe}(\text{OH})_2$ b) FeO and $\text{Fe}(\text{OH})_3$ c) Fe_2O_3 and $\text{Fe}(\text{OH})_3$ d) Fe_3O_4 and $\text{Fe}(\text{OH})_3$
505. A metal having negative reduction potential when dipped in the solution of its own ions, has a tendency :
 a) To pass into the solution

- b) To be deposited from the solution
 c) To become electrically positive
 d) To remain neutral
506. The resistance of 0.5 N solution of an electrolyte in a conductivity cell was found to be 45 ohms. The equivalent conductivity of the same solution isif the electrodes in the cell are 2.2 cm apart and have an area of 3.8 cm².
 a) 25.73 b) 30.75 c) 35.75 d) 15.75
507. The SI unit for ionic mobility is:
 a) m²volt⁻¹sec⁻¹ b) cm²volt⁻¹sec⁻¹ c) cm volt⁻¹sec⁻¹ d) cm⁻²volt⁻¹sec⁻¹
508. Which modifications are necessary to determine resistance of solution by usual method of measurements on Wheatstone bridge principle?
 a) A.C. should be used
 b) A conductivity cell is used
 c) Galvanometer is replaced by magic eye or head phone arrangement
 d) All of above
509. The number of electrons passing per second through a cross-section of Cu wire carrying 10 ampere is :
 a) 6×10^{19} b) 8×10^{19} c) 1×10^{19} d) 1.6×10^{19}
510. Electrolytic reduction of alumina to aluminium by Hall-Heroult process is carried out :
 a) In the presence of NaCl
 b) In the presence of fluoride
 c) In the presence of cryolite, which forms a melt with lower melting temperature
 d) In the presence of cryolite, which forms a melt with higher melting temperature
511. Electrolytes, when dissolved in water, dissociate into their constituent ions. The degree of dissociation of a weak electrolyte increases with
 a) The presence of a substance yielding common ion
 b) Decreasing temperature
 c) Decreasing concentration of the electrolyte
 d) Increasing concentration of the electrolyte
512. The electrolytic bath used in gold plating of copper articles contains :
 a) Molten gold b) CuSO₄ c) AuCl₃ d) AuCl₃ + NaCN
513. Pure water does not conduct electricity because it is
 a) Basic b) Almost not ionised
 c) Decomposed easily d) Acidic
514. Galvanic cell is a device in which
 a) Chemical energy is converted into electrical energy.
 b) Electrical energy is converted into chemical energy.
 c) Chemical energy is seen in the form of heat.
 d) Thermal energy from an outside source is used to drive the cell reaction.
515. A standard hydrogen electrode has zero electrode potential because
 a) Hydrogen is easier to oxidise b) This electrode potential is assumed to be zero
 c) Hydrogen atom has only one electron d) Hydrogen is the lightest element
516. The molar conductivity at infinite dilution of AgNO₃, NaCl and NaNO₃ are 116.5, 110.3 and 105.2 mho cm²mol⁻¹ respectively. The molar conductivity of AgCl is :
 a) 121.6 b) 111.4 c) 130.6 d) 150.2
517. Which is correct about fuel cells?
 a) Cells continuously run as long as fuels are supplied
 b) These are more efficient and free from pollution
 c) These are used to provide power and drinking water to astronauts in space programme
 d) All of the above
518. The value of electronic charge is equal to :

- a) $\frac{\text{Faraday}}{\text{Av. number}}$
 b) Faraday \times Av. number
 c) $\frac{\text{Av. number}}{\text{Faraday}}$
 d) None of these
519. The formula $\alpha = \frac{\Lambda_v}{\Lambda_\infty}$ is valid for :
 a) Weak electrolytes b) Strong electrolytes c) Salts d) None of these
520. A hypothetical electrochemical cell is shown below;
 $A|A^+(xM)||B^+(yM)|B$
 The e.m.f. measured is + 0.20 V. The cell reaction is :
 a) The cell reaction cannot be predicted
 b) $A + B^+ \rightarrow A^+ + B$
 c) $A^+ + B \rightarrow A + B^+$
 d) $A^+ + e^- \rightarrow A; B^+ + e^- \rightarrow B$
521. The laws of electrolysis were proposed by
 a) Kohlraush b) Faraday c) Nernst d) Berthelot
522. When X amperes of current is passed through molten $AlCl_3$ for 96.5 s. 0.09 g of aluminium is deposited. What is the value of X ?
 a) 10 A b) 20 A c) 30 A d) 40 A
523. It is impossible to measure the actual voltage of any half cell by itself because:
 a) Both half cell reactions takes place simultaneously
 b) Of resistance of wire
 c) A reaction does not take place on its own
 d) None of the above
524. The art of electroplating was given by :
 a) Faraday b) Edison c) Graham d) Brugan
525. If 1 faraday of charge is passed through a solution of $CuSO_4$, the amount of copper deposited will be equal to its :
 a) Gram equivalent weight
 b) Gram molecular weight
 c) Atomic weight
 d) Electrochemical equivalent
526. The oxidation potential values of A, B, C and D are -0.03, +0.108 V, -0.07 V and +0.1 V respectively. The non-spontaneous cell reaction takes place between
 a) A and B b) B and D c) D and A d) B and C
527. The conductivity of $N/10$ KCl solution at $20^\circ C$ is $0.0212 \text{ ohm}^{-1} \text{ cm}^{-1}$ and the resistance of cell containing this solution at $20^\circ C$ is 55 ohm. The cell constant is:
 a) 2.173 cm^{-1} b) 1.166 cm^{-1} c) 4.616 cm^{-1} d) 3.324 cm^{-1}
528. What is the value of E_{cell} ?
 $Cr | Cr^{3+} (0.1 M) || Fe^{2+} (0.01 M) | Fe$
 Given, $E^\circ_{Cr^{3+}/Cr} = -0.74 V$
 and $E^\circ_{Fe^{2+}/Fe} = -0.44 V$
 a) +0.2941 V b) +0.5212 V c) +0.1308 V d) -0.2606 V
529. The E° for OCl^-/Cl^- and $Cl^-/\frac{1}{2}Cl_2$ are 0.94 V and $-1.36 V$; E° for $OCl^-/\frac{1}{2}Cl_2$ is :
 a) $-0.42 V$ b) $-2.20 V$ c) $0.52 V$ d) $1.04 V$
530. The cell reaction for the given cell is spontaneous if :
 $Pt(H_2)|H^+(1M)||H^+(1M)|Pt(H_2)$
 P_1 P_2

- a) 1 : 2 : 3 : 1 b) 6 : 3 : 2 : 6 c) 6 : 3 : 0 : 0 d) 3 : 2 : 1 : 0
554. In the problem 15, ionic conductance of K^+ ion is :
a) 64.35 b) 60.20 c) 262.26 d) 26.22
555. In the electrochemical reaction,
 $2Fe^{3+} + Zn \rightarrow Zn^{2+} + 2Fe^{2+}$
increasing the concentration of Fe^{2+}
a) Increases cell emf b) Increases the current flow
c) Decreases the cell emf d) Alter the pH of the solution
556. How many electrons are there in one coulomb?
a) 6.02×10^{21} b) 6.24×10^{18} c) 6.24×10^{15} d) 6.02×10^{16}
557. The element which can displace three other halogens from their compound is
a) F b) Cl c) Br d) I
558. The units of equivalent conductivity is
a) $S\ cm^2$ b) $ohm\ cm^2(g - equivalent)$
c) $ohm\ cm$ d) $ohm^{-1}\ cm^2(g - equivalent)^{-1}$
559. Calculate the equilibrium constant for the reaction, at $25^\circ C$
 $Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$
at $25^\circ C$, $E^\circ_{cell} = 0.47\ V$, $R = 8.314\ JK^{-1}$
 $F = 96500\ C$ is
a) 1.8×10^{15} b) 8.5×10^{15} c) 1.8×10^{10} d) 85×10^{15}
560. The ratio of weights of hydrogen and magnesium deposited by the same amount of electricity from H_2SO_4 and $MgSO_4$ in aqueous solution are :
a) 1 : 8 b) 1 : 12 c) 1 : 16 d) None of these
561. The Λ^∞ of NH_4Cl , $NaOH$ and $NaCl$ are 129.8, 217.4 and 108.9 $ohm^{-1}\ cm^2\ eq^{-1}$ respectively. The λ_∞ of NH_4OH is $ohm^{-1}\ cm^2\ eq^{-1}$.
a) 238.3 b) 218 c) 240 d) 260
562. The reaction at cathode during the electrolysis of aqueous solution of $NaCl$ in Nelson cell is :
a) $2Cl^- \rightarrow Cl_2 + 2e$
b) $2H^+ + 2e \rightarrow H_2$
c) $2OH^- \rightarrow H_2 + O_2 + 2e$
d) $Na^+ + e \rightarrow Na$
563. Which of the following is an additive property?
a) Conductance b) Viscosity c) Surface tension d) None of these
564. The limiting molar conductivities of $NaCl$, KBr and KCl are 126, 152 and 150 $S\ cm^2\ mol^{-1}$ respectively. The Λ for $NaBr$ is :
a) 302 $S\ cm^2\ mol^{-1}$ b) 176 $S\ cm^2\ mol^{-1}$ c) 278 $S\ cm^2\ mol^{-1}$ d) 128 $S\ cm^2\ mol^{-1}$
565. The calomel electrode is reversible with respect to :
a) Hg_2^{2+} b) H^+ c) Hg^{2+} d) Cl^-
566. Reaction taking place at anode in dry cell is :
a) $Zn^{2+} + 2e \rightarrow Zn(s)$ b) $Zn(s) \rightarrow Zn^{2+} + 2e$ c) $Mn^{2+} + 2e \rightarrow Mn(s)$ d) $Mn(s) \rightarrow Mn^{2+} + 2e$
567. Number of faraday required to liberate 8 g of H_2 is :
a) 8 b) 16 c) 4 d) 2
568. The number of coulombs required to reduce 12.3 g of nitrobenzene to aniline, is
a) 96500 C b) 5790 C c) 95700 C d) 57900 C
569. On passing 0.1 F of electricity through aluminium metal deposited at cathode is ($Al = 27$)
a) 0.3 g b) 0.6 g c) 0.9 g d) 1.2 g
570. During electrolysis of H_2O , the molar ratio of H_2 and O_2 formed is :
a) 2 : 1 b) 1 : 2 c) 1 : 3 d) 1 : 1
571. At infinite dilution stage, the solution of CH_3COOH in water does not contain :

- a) H^+ ion b) CH_3COO^- ion c) CH_3COOH molecule d) All of these
572. 1 faraday of electricity will liberate 1 g-atom of the metal from the solution of :
 a) NaCl b) BaCl_2 c) CuSO_4 d) AlCl_3
573. The standard electrode potential of hydrogen electrode at 1 M concentration and hydrogen gas at 1 atm pressure is
 a) 1 V b) 6 V c) 8 V d) 0 V
574. The emf of a Daniell cell at 298 K is E_1 , $\text{Zn} | \text{ZnSO}_4 || \text{CuSO}_4 | \text{Cu}$. When the concentration of (0.01 M) (1.0 M) ZnSO_4 is 1.0 M and that of CuSO_4 is 0.01 M, the emf 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) $E_2 = 0 \neq E_1$
575. The acid used in lead storage battery is
 a) H_2SO_4 b) H_3PO_4 c) HCl d) HNO_3
576. The conductance of 1 cm^3 of a solution is known as its :
 a) Resistance
 b) Conductivity
 c) Equivalent conductivity
 d) Molecular conductivity
577. The limiting molar conductivities Λ° for NaCl, KBr and KCl are 126, 152 and 150 $\text{S cm}^2 \text{mol}^{-1}$ respectively. The Λ° for NaBr is
 a) 128 $\text{S cm}^2 \text{mol}^{-1}$ b) 176 $\text{S cm}^2 \text{mol}^{-1}$ c) 278 $\text{S cm}^2 \text{mol}^{-1}$ d) 302 $\text{S cm}^2 \text{mol}^{-1}$
578. The variation of equivalent conductivity of weak electrolyte with $\sqrt{\text{concentration}}$ is correctly shown in figure :
 a)  b)  c)  d) 
579. The electrode potential measures the :
 a) Tendency of the electrode to gain or lose electrons
 b) Tendency of the cell reaction to occur
 c) Difference in the ionisation potential of electrode and metal ion
 d) Current carried by an electrode
580. Metals can be prevented from rusting by :
 a) Connecting iron to more electropositive metal, *i. e.*, cathodic protection
 b) Connecting iron to more electropositive metal, *i. e.*, anodic protection
 c) Connecting iron to less electropositive metal, *i. e.*, anodic protection
 d) Connecting iron to less electropositive metal, *i. e.*, cathodic protection
581. The number of faraday required to liberate 1 mole of any element indicates :
 a) Weight of element
 b) Conductance of electrolyte
 c) Charge on the ion of that element
 d) None of the above
582. Lithium is generally used as an electrode in high energy density batteries. This is because:
 a) Lithium is the lightest element
 b) Lithium has quite high negative reduction potential
 c) Lithium is quite reactive
 d) Lithium does not corrode easily
583. Corrosion of iron is essentially an electrochemical phenomenon where the cell reactions are
 a) Fe is oxidised to Fe^{2+} and dissolved oxygen in water is reduced to OH^- b) Fe is oxidised to Fe^{3+} and H_2O is reduced to O_2^{2-}

- c) Fe is oxidised to Fe^{2+} and H_2O is reduced to O_2^- d) Fe is oxidised to Fe^{2+} and H_2O is reduced to O_2
584. In the electrodeposition of Ag, the silver ions are:
 a) Reduced at anode b) Reduced at cathode c) Oxidised at anode d) Oxidised at cathode
585. Standard E° of the half cell $\text{Fe}|\text{Fe}^{2+}$ is + 0.44 V and standard E° of half cell $\text{Cu}|\text{Cu}^{2+}$ is -0.32 V then :
 a) Cu oxidises Fe^{2+} ion b) Cu^{2+} oxidises Fe c) Cu reduces Fe^{2+} ion d) Cu^{2+} reduces Fe
586. Which of the following is displaced by Fe?
 a) Ag b) Zn c) Na d) All of these
587. The electrochemical cell stops working after sometimes because
 a) Electrode potential of both the electrodes becomes zero
 b) Electrode potential of both the electrodes becomes equal
 c) One of the electrodes is eaten away
 d) The cell reaction gets reversed
588. The resistance of 0.01 N solution of an electrolyte was found to be 210 ohm at 298 K, using a conductivity cell of cell constant 0.66 cm^{-1} . The conductivity of solution is :
 a) $3.14 \times 10^{-3} \text{ mho cm}^{-1}$
 b) $3.14 \times 10^{-3} \text{ mho}^{-1} \text{ cm}$
 c) 3.14 mho cm^{-1}
 d) $3.14 \text{ mho}^{-1} \text{ cm}^{-1}$
589. The molar conductivity of acetic acid at infinite dilution is 390.7 and for 0.1 M acetic acid solution is 5.2 $\text{mho cm}^2 \text{ mol}^{-1}$. The degree of dissociation of 0.1 M CH_3COOH solution is :
 a) 13.3% b) 0.0133% c) 1.33% d) 133%
590. When a lead storage battery is charged, it acts as
 a) A primary cell b) A galvanic cell c) A concentration cell d) An electrolytic cell
591. For gold plating, the electrolyte used is
 a) AuCl_3 b) HAuCl_4 c) $\text{K}[\text{Au}(\text{CN})_2]$ d) None of these
592. How many coulomb of electricity are consumed when 100 mA current is passed through a solution of AgNO_3 for 30 minute during an electrolysis experiment?
 a) 108 b) 18000 c) 180 d) 3000
593. How many kJ of energy is evolved, when a current of 2.00 A passes for 200 s under the potential of 230 V?
 a) 56 kJ b) 86 kJ c) 36 kJ d) 92 kJ
594. What will be the emf for the given cell $\text{Pt} | \text{H}_2(p_1) | \text{H}^+(\text{aq}) || \text{H}_2(p_2) | \text{Pt}$?
 a) $\frac{RT}{2F} \log \frac{p_1}{p_2}$ b) $\frac{RT}{F} \log \frac{p_1}{p_2}$ c) $\frac{RT}{F} \log \frac{p_2}{p_1}$ d) None of these
595. The time required to coat a metal surface of 80 cm^2 with $5 \times 10^{-3} \text{ cm}$ thick layer of silver (density 1.05 g cm^{-3} with the passage of 3A current through a silver nitrate solution is :
 a) 115 sec b) 125 sec c) 135 sec d) 145 sec
596. On electrolysing a solution of dilute H_2SO_4 between platinum electrodes, the gas evolved at the anode and cathode are respectively :
 a) SO_2 and O_2 b) SO_3 and H_2 c) O_2 and H_2 d) H_2 and O_2
597. The electrochemical equivalent of silver is 0.0011180 g. When an electric current of 0.5 ampere is passed through an aqueous silver nitrate solution of 200 sec, the amount of silver deposited is:
 a) 1.1180 g b) 0.11180 g c) 5.590 g d) 0.5590 g
598. Galvanised iron sheets have coating of :
 a) Cu b) Sn c) Zn d) Carbon
599. Ionisation depends upon
 a) Pressure b) Volume c) Dilution d) None of these
600. Standard free energies of formation (in kJ/mol) at 298 K are - 237.2, -394.4 and - 8.2 for $\text{H}_2\text{O}(l)$, $\text{CO}_2(g)$ and pentane (g), respectively. The value of E°_{cell} for the pentane-oxygen fuel cell is :
 a) 2.0968 V b) 1.0968 V c) 0.0968 V d) 1.968 V

601. In the electrolysis of water, 1 F of electrical energy would evolve
 a) 1 mole of oxygen b) 1 g atom of oxygen c) 8 g of oxygen d) 22.4 L of oxygen
602. Given $l/a = 0.5\text{cm}^{-1}$, $R = 50\text{ ohm}$, $N = 1.0$. The equivalent conductance of the electrolytic cell is
 a) $10\Omega^{-1}\text{ cm}^2\text{ g equiv}^{-1}$ b) $20\Omega^{-1}\text{ cm}^2\text{ g equiv}^{-1}$
 c) $300\Omega^{-1}\text{ cm}^2\text{ g equiv}^{-1}$ d) $100\Omega^{-1}\text{ cm}^2\text{ g equiv}^{-1}$
603. If 3 F of electricity is passed through the solutions of AgNO_3 , CuSO_4 and AuCl_3 , the molar ratio of the cations deposited at the cathodes will be
 a) 1:1:1 b) 1:2:3 c) 3:2:1 d) 6:3:2
604. If $\text{Mg}^{2+} + 2e \rightarrow \text{Mg}(s)$; $E = -2.37\text{ V}$,
 $\text{Cu}^{2+} + 2e \rightarrow \text{Cu}(s)$; $E = +0.34\text{ V}$?
 The e.m.f. of the cell $\text{Mg} | \text{Mg}^{2+} || \text{Cu}^{2+} | \text{Cu}$ is :
 a) 2.71 V b) 2.30 V c) 2.80 V d) 1.46 V
605. The standard reduction potentials of $\text{Zn}^{2+} | \text{Zn}$, $\text{Cu}^{2+} | \text{Cu}$ and $\text{Ag}^+ | \text{Ag}$ are respectively - 0.76, 0.34 and 0.8 V. The following cells were constructed
 I $\text{Zn} | \text{Zn}^{2+} || \text{Cu}^{2+} | \text{Cu}$
 II $\text{Zn} | \text{Zn}^{2+} || \text{Ag}^+ | \text{Ag}$
 III $\text{Cu} | \text{Cu}^{2+} || \text{Ag}^+ | \text{Ag}$
 What is the correct order of E_{cell}° of these cells?
 a) II > III > I b) II > I > III c) I > II > III d) III > I > II
606. What is the effect of dilution on the equivalent conductance of strong electrolyte?
 a) Decreases on dilution b) Remains unchanged
 c) Increases on dilution d) None of these
607. For which electrolyte the evaluation of A^∞ is not possible by extrapolation of Λ vs \sqrt{c} curves to zero concentration?
 a) KCl b) NH_4OH c) NaCl d) K_2SO_4
608. The standard reduction potential, E° for the half-reactions are as
 $\text{Zn} \rightleftharpoons \text{Zn}^{2+} + 2e^-$, $E^\circ = \ominus 0.76\text{ V}$
 $\text{Fe} \rightleftharpoons \text{Fe}^{2+} + 2e^-$, $E^\circ = +0.41\text{ V}$
 The E°_{cell} for the cell formed by these two electrodes is
 a) -0.35 V b) -1.17 V c) +0.35 V d) +1.17 V
609. In the electrochemical cell, $\text{H}_2(\text{g}) 1\text{ atm} | \text{H}^+ (1\text{ M}) || \text{Cu}^{2+} (1\text{ M}) | \text{Cu}(s)$
 Which one of the following statements is true?
 a) H_2 is anode, Cu is cathode b) Cu is anode, H_2 is cathode
 c) Oxidation occurs at Cu electrode d) Reduction occurs at H_2 electrode
610. Which of the following does not conduct electricity?
 a) Fused NaCl b) Solid NaCl c) Brine solution d) Copper
611. The ionic mobility of alkali metal ions in aqueous solution is maximum for :
 a) K^+ b) Rb^+ c) Li^+ d) Na^+
612. The e.m.f. of the cell involving following changes,
 $\text{Zn}(s) + \text{Ni}^{2+}(1\text{M}) \rightarrow \text{Zn}^{2+}(1\text{M}) + \text{Ni}(s)$ is 0.5105 V. The standard e.m.f. of the cell is :
 a) 0.540 V b) 0.4810 V c) 0.5696 V d) 0.5105 V
613. The factor temperature coefficient of e.m.f. is :
 a) $(\partial E/\partial T)_P$ b) $(\partial E/\partial T)_T$ c) $(\partial E/\partial V)_T$ d) None of these
614. On passing 1 F of electricity through the electrolytic cells containing Ag^+ , Ni^{2+} and Cr^{3+} ions solution, the deposited Ag (at. wt. = 108), Ni (at. wt. = 59) and Cr (at. wt. = 52) is

Ag	Ni	Cr	Ag	Ni	Cr
a) 108 g	29.5 g	17.3 g	b) 108 g	59.5 g	52.0 g
c) 108 g	108 g	108 g	d) 108 g	117.5 g	166 g

615. Which of the following expression is correct?

- a) $\Delta G^\circ = -nFE_{\text{cell}}^\circ$ b) $\Delta G^\circ = +nFE_{\text{cell}}^\circ$
 c) $\Delta G^\circ = -2.303 RT nFE_{\text{cell}}^\circ$ d) $\Delta G^\circ = -nF \log K_c$

616. For which cell e.m.f. is independent of the concentration of electrolytes used?

- a) Fe | FeO(s)|KOH(aq)| Nb) Pt(H₂)| HCl |Pt(Cl₂) c) Zn|Zn(NO₃)₂|| CuSO₄ |Cd) Hg, HgCl₂| KCl ||AgNO₃|.

617. In the problem 13, the dissociation constant of acid is :

- a) 2.067×10^{-4} b) 1.02×10^{-4} c) 1.02×10^{-3} d) 1.02×10^{-5}

618. Which are used as secondary reference electrodes?

- a) Calomel electrode
 b) Ag/AgCl electrode
 c) Hg/Hg₂Cl₂ – KCl electrode
 d) All of the above

619. The amount of electricity required to produce one mole of copper from copper sulphate solution will be

- a) 1 F b) 2.33 F c) 2 F d) 1.33 F

620. The weight ratio of Al and Ag deposited using the same quantity of current is :

- a) 9 : 108 b) 2 : 12 c) 108 : 9 d) 3 : 8

621. When same electric current is passed through the solution of different electrolytes in series the amounts of the element deposited on the electrode are in the ratio of their:

- a) At.no. b) At. wt. c) Sp. gravity d) Eq. wt.

622. The metal used to recover copper from Asolution of CuSO₄ is

- a) Fe b) He c) Na d) Ag

623. If the half-cell reaction $A + e \rightarrow A^-$ has a large negative reduction potential, it follows that :

- a) A is readily reduced b) A is readily oxidized c) A⁻ is readily reduced d) A⁻ is readily oxidized

624. Same amount of electric current is passed through solutions of AgNO₃ and HCl. If 1.08 g of silver is obtained in the first case, the amount of hydrogen liberated at STP in the second case is

- a) 224 cm³ b) 1.008 g c) 112 cm³ d) 22400 cm³

625. The standard emf of Agalvanic 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, $F = 96500 \text{ C mol}^{-1}$, $R = 8.314 \text{ JK}^{-1} \text{ mol}^{-1}$)

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

626. The correct order of chemical reactivity with water according to electrochemical series is:

- a) $K > \text{Mg} > \text{Zn} > \text{Cu}$ b) $\text{Mg} > \text{Zn} > \text{Cu} > \text{K}$ c) $K > \text{Zn} > \text{Mg} > \text{Cu}$ d) $\text{Cu} > \text{Zn} > \text{Mg} > \text{K}$

627. Calculate using appropriate molar conductance of the CH₃COOH from the molar conductances of electrolytes listed below at infinite dilution in H₂O at 25°C :

Electrode	KCl	NaCl	HCl	NaOAc	KNO ₃
S Cm ² mol ⁻¹	149.9	126.5	426.2	91.0	145.0

- a) 51.2 b) 552.7 c) 390.7 d) 217.5

628. The E° for half-cell Fe/Fe²⁺ and Cu/Cu²⁺ are -0.44 V and +0.32 V respectively, then

- a) Cu²⁺ oxidises Fe b) Cu oxidises FeFe²⁺ c) Cu reduces Fe²⁺ d) Cu²⁺ oxidises Fe²⁺

629. The same amount of electricity was passed through two cells containing molten Al₂O₃ and molten NaCl. If 1.8 g of Al were liberated in one cell, the amount of Na liberated in the other cell is :

- a) 4.6 g b) 2.3 g c) 6.4 g d) 3.2 g

630. 1 mole of Al is deposited by X coulomb of electricity passing through aluminium nitrate solution. The number of mole of silver deposited by X coulomb of electricity from silver nitrate solution is :

- a) 3 b) 4 c) 2 d) 1

631. The platinum electrodes were immersed in a solution of cupric sulphate and electric current was passed through the solution. After some time, it was found that colour of copper sulphate disappeared with evolution of gas at the electrode. The colourless solution contain

- a) Copper sulphate b) Copper hydroxide c) Platinum sulphate d) Sulphuric acid

CHEMISTRY (QUESTION BANK)**3.ELECTROCHEMISTRY****: ANSWER KEY :**

1)	c	2)	b	3)	a	4)	d	145)	d	146)	c	147)	a	148)	a
5)	a	6)	d	7)	d	8)	a	149)	d	150)	d	151)	a	152)	b
9)	a	10)	c	11)	a	12)	a	153)	b	154)	c	155)	b	156)	a
13)	a	14)	d	15)	b	16)	c	157)	c	158)	d	159)	b	160)	b
17)	a	18)	c	19)	a	20)	a	161)	a	162)	b	163)	a	164)	b
21)	a	22)	b	23)	b	24)	a	165)	b	166)	b	167)	b	168)	d
25)	a	26)	b	27)	d	28)	d	169)	c	170)	b	171)	b	172)	b
29)	d	30)	c	31)	b	32)	d	173)	a	174)	b	175)	d	176)	c
33)	b	34)	d	35)	d	36)	a	177)	b	178)	b	179)	c	180)	a
37)	c	38)	a	39)	b	40)	a	181)	c	182)	a	183)	d	184)	b
41)	d	42)	c	43)	d	44)	c	185)	d	186)	b	187)	b	188)	b
45)	a	46)	b	47)	c	48)	a	189)	b	190)	c	191)	a	192)	d
49)	b	50)	d	51)	a	52)	a	193)	a	194)	a	195)	d	196)	a
53)	a	54)	d	55)	a	56)	a	197)	a	198)	d	199)	a	200)	a
57)	d	58)	d	59)	a	60)	c	201)	c	202)	b	203)	b	204)	c
61)	c	62)	b	63)	d	64)	b	205)	c	206)	d	207)	a	208)	a
65)	d	66)	d	67)	a	68)	a	209)	c	210)	d	211)	a	212)	a
69)	a	70)	d	71)	a	72)	c	213)	d	214)	d	215)	b	216)	b
73)	c	74)	d	75)	a	76)	d	217)	d	218)	b	219)	d	220)	a
77)	c	78)	d	79)	b	80)	b	221)	c	222)	c	223)	a	224)	c
81)	a	82)	a	83)	d	84)	c	225)	b	226)	a	227)	c	228)	a
85)	a	86)	b	87)	c	88)	b	229)	a	230)	a	231)	d	232)	a
89)	b	90)	a	91)	a	92)	b	233)	d	234)	a	235)	b	236)	a
93)	c	94)	b	95)	a	96)	c	237)	c	238)	b	239)	a	240)	a
97)	c	98)	c	99)	c	100)	c	241)	d	242)	c	243)	c	244)	a
101)	d	102)	b	103)	a	104)	a	245)	a	246)	d	247)	a	248)	c
105)	d	106)	c	107)	c	108)	b	249)	b	250)	b	251)	a	252)	c
109)	d	110)	a	111)	b	112)	c	253)	b	254)	d	255)	d	256)	c
113)	a	114)	d	115)	b	116)	b	257)	a	258)	d	259)	c	260)	d
117)	c	118)	b	119)	b	120)	d	261)	b	262)	c	263)	a	264)	c
121)	b	122)	c	123)	a	124)	c	265)	c	266)	a	267)	c	268)	a
125)	b	126)	d	127)	b	128)	a	269)	d	270)	c	271)	b	272)	a
129)	b	130)	c	131)	a	132)	d	273)	a	274)	c	275)	b	276)	a
133)	d	134)	d	135)	c	136)	b	277)	c	278)	b	279)	a	280)	b
137)	c	138)	a	139)	a	140)	d	281)	c	282)	b	283)	d	284)	c
141)	c	142)	b	143)	b	144)	b	285)	d	286)	d	287)	b	288)	a

289) a	290) c	291) a	292) a	461) c	462) b	463) b	464) b
293) d	294) c	295) c	296) c	465) c	466) a	467) b	468) a
297) a	298) b	299) d	300) a	469) c	470) a	471) a	472) c
301) b	302) c	303) c	304) d	473) a	474) c	475) c	476) d
305) b	306) c	307) b	308) d	477) b	478) a	479) a	480) d
309) d	310) b	311) a	312) b	481) b	482) a	483) d	484) c
313) c	314) d	315) c	316) d	485) d	486) c	487) a	488) c
317) a	318) a	319) a	320) a	489) d	490) c	491) d	492) d
321) c	322) d	323) c	324) a	493) c	494) d	495) c	496) a
325) a	326) d	327) c	328) a	497) c	498) c	499) a	500) d
329) b	330) a	331) d	332) a	501) b	502) c	503) a	504) c
333) a	334) a	335) a	336) a	505) a	506) a	507) a	508) d
337) a	338) d	339) d	340) d	509) a	510) c	511) c	512) d
341) c	342) b	343) a	344) b	513) b	514) a	515) b	516) a
345) c	346) a	347) a	348) a	517) d	518) a	519) a	520) b
349) c	350) a	351) d	352) b	521) b	522) a	523) a	524) a
353) a	354) a	355) b	356) a	525) a	526) a	527) b	528) a
357) c	358) a	359) b	360) d	529) c	530) a	531) a	532) c
361) b	362) a	363) c	364) b	533) c	534) b	535) a	536) d
365) d	366) b	367) b	368) a	537) b	538) c	539) b	540) d
369) a	370) c	371) a	372) c	541) a	542) c	543) d	544) c
373) b	374) b	375) d	376) a	545) b	546) b	547) a	548) b
377) d	378) b	379) d	380) d	549) d	550) d	551) c	552) b
381) c	382) b	383) d	384) a	553) c	554) a	555) c	556) b
385) a	386) b	387) b	388) d	557) a	558) d	559) b	560) d
389) d	390) a	391) b	392) c	561) a	562) b	563) a	564) d
393) c	394) a	395) a	396) a	565) d	566) b	567) a	568) d
397) c	398) d	399) a	400) c	569) c	570) a	571) c	572) a
401) c	402) c	403) a	404) a	573) d	574) b	575) a	576) b
405) d	406) b	407) b	408) c	577) a	578) d	579) a	580) a
409) c	410) c	411) d	412) a	581) c	582) b	583) a	584) b
413) d	414) a	415) d	416) c	585) b	586) a	587) b	588) a
417) a	418) d	419) a	420) d	589) d	590) d	591) c	592) c
421) a	422) c	423) c	424) d	593) d	594) a	595) b	596) c
425) d	426) c	427) a	428) a	597) b	598) c	599) c	600) b
429) b	430) c	431) c	432) a	601) c	602) a	603) d	604) a
433) b	434) b	435) c	436) c	605) b	606) c	607) b	608) c
437) c	438) c	439) b	440) a	609) a	610) b	611) b	612) d
441) d	442) b	443) a	444) b	613) a	614) a	615) a	616) a
445) a	446) b	447) b	448) d	617) b	618) d	619) c	620) a
449) d	450) c	451) c	452) b	621) d	622) a	623) d	624) c
453) c	454) b	455) b	456) b	625) d	626) a	627) c	628) a
457) a	458) c	459) b	460) d	629) a	630) a	631) d	

CHEMISTRY (QUESTION BANK)

3.ELECTROCHEMISTRY

: HINTS AND SOLUTIONS :

- 1 (c)
 $\text{Al} \rightarrow \text{Al}^{3+} + 3e^{-}$
 The charge required = $3 \times 96500 \text{ C}$
- 2 (b)
 Eq. of $\text{H}_2 = \text{Eq. of Cu}$
 $\therefore \frac{0.504}{1} = \frac{W}{63.5/2}$
 $\therefore W_{\text{Cu}} = 16 \text{ g}$
- 3 (a)
 $E^{\circ} = \frac{0.059}{n} \log K_{eq}$ and $\Delta G^{\circ} = -nE^{\circ}F$
 $\therefore \Delta G^{\circ} = +ve, E^{\circ}$ will be $-ve$ and $K_{eq} < 1$; one should not write $\Delta G^{\circ} > 0$.
- 4 (d)
 $E_{\text{cell}}^{\circ} = 0.87 + 0.40 = 1.27 \text{ V}$
 Cell reaction : $\text{Fe} + \text{Ni}_2\text{O}_3 \rightarrow \text{FeO} + 2\text{NiO}$
- 5 (a)
 The given values are E_{RP}° . More is E_{RP}° more is the tendency to gain electron or to show reduction or to show strong oxidant nature.
- 6 (d)
 $E_{\text{cell}}^{\circ} = E_{OP_{\text{Sn}}}^{\circ} + E_{RP_{\text{Fe}}}^{\circ} = 0.14 + (-0.44)$
 $= -0.30 \text{ V}$
- 8 (a)
 Smallest ion possesses maximum mobility.
- 9 (a)
 $\Lambda_{\text{M}}^{\circ} = \Lambda_{\text{a}}^{\circ} + \Lambda_{\text{c}}^{\circ}$
- 10 (c)
 Molten NaCl possesses Na^{+} and Cl^{-} ions.
- 11 (a)
 Given, that
 $\text{Zn} / \text{Zn}^{2+} || \text{Cu}^{2+} / \text{Cu}$
 $\therefore \text{Zn}$ is anode and Cu is cathode.
 Given,
 $\text{Zn}^{2+} / \text{Zn} = -0.76 \text{ V}$
- $\text{Cu}^{2+} / \text{Cu} = +0.34 \text{ V}$
 $E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}$
 $= 0.34 - (-0.76)$
 $= 0.34 + 0.76$
 $= 1.10 \text{ V}$
- 12 (a)
 Net redox change is zero.
- 13 (a)
 Cathode $2\text{H}_2\text{O} + 2e^{-} \rightarrow \text{H}_2 + 2\text{OH}^{-}$
 Anode : $\text{H}_2\text{O} \rightarrow 2\text{H}^{+} + \frac{1}{2}\text{O}_2 + 2e^{-}$
- 14 (d)
 More is E_{OP}° , more is reducing power of metal.
- 15 (b)
 The charge required to liberate one gram equivalent of an element is always equal to 1 faraday (*i.e.*, = 96500 C).
- 16 (c)
 $\text{pH} < 7$; Aqueous solution of CuSO_4 is acidic in nature. Furthermore some drops of H_2SO_4 is also added during electrolysis.
- 17 (a)
 Anode is positive electrode and cathode is $-ve$ electrode in electrolytic cell whereas, anode is $-ve$ electrode and cathode is $+ve$ electrode in electrochemical cells.
- 18 (c)
 $\Delta G = \Delta H - T \Delta S$
 For a spontaneous cell reaction, ΔH should be negative and ΔS should be positive. Hence, ΔG should be negative.
- 19 (a)
 Cell reaction is $\text{Mg} + \text{Sn}^{2+} \rightarrow \text{Mg}^{2+} + \text{Sn}$
 $E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{[\text{Mg}^{2+}]}{[\text{Sn}^{2+}]}$
 $= (2.34 - 0.14) - \frac{0.0591}{2} \log \frac{10^{-2}}{10^{-1}} = 2.23 \text{ V}$
- 20 (a)

- pH of solution increases due to formation of LiOH or due to increase in $[\text{OH}^-]$ because H^+ ions are discharged at cathode in preference to Li^+ .
- 21 (a) In the process of electro decomposition for purification of metal, impure metal acts as anode.
- 22 (b) Specific conductivity (κ)

$$= \frac{1}{R} \times \text{cell constant}$$
Cell constant = $\kappa \times R$

$$= 0.0129 \times 100 = 1.29$$
- 23 (b) According to Nernst equation.

$$E_{\text{cell}} = E_{\text{Cell}}^{\circ} + \frac{0.0591}{2} \log \frac{[\text{Cu}^{2+}]}{[\text{Zn}^{2+}]}$$

$$E_{\text{cell}} = E_{\text{Cell}}^{\circ} - \frac{0.0591}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$
Or $y = c + (-m)x$
Thus, the slope is negative.
- 24 (a) In MnO_4^- the oxidation number of Mn is + 7.

$$\begin{array}{ccc} +7 & & +2 \\ \text{Mn} & + & 5e^- \rightarrow \text{Mn} \end{array}$$
In the reaction, 5 electrons are involved hence 5 Faraday will be needed for the reduction of 1 mole of MnO_4^- .
Therefore, for 0.5 mole of MnO_4^- , number of Faradays required = 2.5 F
- 25 (a) Anode is electrode at which oxidation occurs.
- 26 (b) MnO_2 in Lechlanche cell.
- 27 (d) As Cr has maximum oxidation potential value, therefore its oxidation should be easiest
- 28 (d) More is reduction potential, more is the power to get itself reduced or greater is oxidising power.
- 29 (d) $F = N \times e$
- 30 (c) NaCl gives Na^+ and Cl^- ions;
At anode : $\text{Cl}^- \rightarrow (1/2)\text{Cl}_2 + e$
At cathode : $\text{H}^+ \rightarrow (1/2)\text{H}_2 + e$
- 31 (b) Electrons flow from Zn to Cu in outside circuit and current from Cu to Zn.
- 32 (d)

$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$$

$$\text{Ni} / \text{Ni}^{2+} [1.0 \text{ M}] \parallel \text{Au}^{3+} [1.0 \text{ M}] | \text{Au}$$

$$E_{\text{cell}} (\text{Au}^{3+} / \text{Au}) = 0.150 \text{ V}$$

$$E_{\text{cell}} (\text{Ni}^{2+} / \text{Ni}) = - 0.25 \text{ V}$$

$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$$

$$= 0.150 - (- 0.25)$$

$$= 0.15 + 0.25$$

$$= + 0.4 \text{ V}$$
- 33 (b) 50 % H_2SO_4 aqueous solution can be electrolysed by using Pt electrodes as

$$2\text{H}_2\text{SO}_4 \rightarrow 2\text{HSO}_4^- + 2\text{H}^+$$

$$2\text{HSO}_4^- \rightarrow \text{H}_2\text{S}_2\text{O}_8 + 2e^- \quad (\text{at anode})$$
- 34 (d) It is fact.
- 35 (d) For the given cell,

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$
1.
$$E_1 = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{1}{0.1}$$

$$= E_{\text{cell}}^{\circ} - \frac{0.0591}{2}$$
2.
$$E_2 = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{1}{1}$$

$$= E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \times 0$$

$$= E_{\text{cell}}^{\circ}$$
3.
$$E_3 = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{0.1}{1}$$

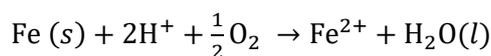
$$= E_{\text{cell}}^{\circ} + \frac{0.0591}{2}$$

$$\therefore E_3 > E_2 > E_1$$
- 36 (a) Transport number of an ion

$$= \frac{\text{current carried by that ion}}{\text{total current carried by both the ions}}$$
- 37 (c) Reduction is always carried out at cathode.
- 38 (a) Reactions
(i) $\text{Fe}(s) \rightarrow \text{Fe}^{2+} + 2e^-$, $E^{\circ} = + 0.44 \text{ V}$
and $\Delta G_1^{\circ} = - nE^{\circ}F = - 2 \times 0.44 \times F$
(ii) $2\text{H}^+ + 2e^- + \frac{1}{2}\text{O}_2 \rightarrow \text{H}_2\text{O}(l)$; $E^{\circ} = +1.23 \text{ V}$

$$\text{and } \Delta G_2^\circ = -2 \times (+1.23) \times F$$

Net reaction,



$$\begin{aligned} \Delta G_3^\circ &= \Delta G_1^\circ + \Delta G_2^\circ \\ &= -2 \times (+0.44)F + (-2 \times 1.23 \times F) \\ &= -0.88F \times -2.46F = -3.34F \\ &= -3.34 \times 96500 \text{ J} \\ &= -322.31 \text{ kJ} = -322 \text{ kJ} \end{aligned}$$

39 (b) 2 faraday will deposit 2 eq. or 1 mole of Cu.

40 (a) Cl_2 is placed above F_2 in electrochemical series, halogen placed below replaces the other from its solution.

41 (d)

$$E_{\text{cell}}^\circ = \frac{2.303RT}{nF} \log K_{\text{eq}}$$

$$0.295 = \frac{0.0591}{2} \log K_{\text{eq}}$$

$$\therefore \log K_{\text{eq}} = 10$$

$$\therefore K_{\text{eq}} = 10^{10}$$

42 (c)

$$k = \frac{1}{R} \times \frac{1}{a} = \frac{1}{32} \times \frac{1.8}{5.4} = 0.0104$$

And $\lambda = k \times V = 0.0104 \times 10,000 = 104$

43 (d)

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

44 (c)

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

$$\therefore 0.295 = \frac{0.059}{2} \log K_c$$

$$\therefore K_c = 10^{10}$$

45 (a) High value for E_{red}° . Shows more electronegativity *i.e.*, Zn is more electropositive than Fe.

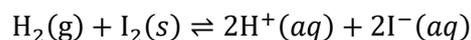
$$(E_{\text{Zn}^{2+}/\text{Zn}}^\circ < E_{\text{Fe}^{2+}/\text{Fe}}^\circ)$$

46 (b) Eq. of Cu = Eq. of Ag

$$\therefore \frac{W}{63.5/2} = \frac{1.08}{108}$$

$$\therefore W_{\text{Cu}} = 0.3175 \text{ g}$$

47 (c) The cell reaction is



$$0.7714 = 0.535 - \frac{0.0591}{2} \log \frac{[\text{H}^+]^2 [\text{I}^-]^2}{p_{\text{H}_2}}$$

$$\therefore \text{pH} = 3$$

48 (a)

$$E_{\text{Cr}^{3+}/\text{Cr}^{2+}}^\circ = -0.41 \text{ V}$$

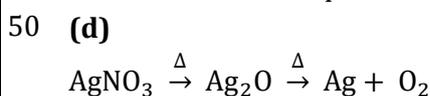
$$E_{\text{Mn}^{3+}/\text{Mn}^{2+}}^\circ = +1.57 \text{ V}$$

$$E_{\text{Fe}^{3+}/\text{Fe}^{2+}}^\circ = +0.77 \text{ V}$$

$$E_{\text{Co}^{3+}/\text{Co}^{2+}}^\circ = +1.97 \text{ V}$$

More negative value of E_{red}° indicates better reducing agent thus easily oxidized. Thus, oxidation of Cr^{2+} to Cr^{3+} is the easiest.

49 (b) In other cells, two liquid are not present.

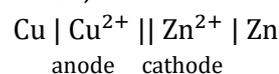


51 (a)

$$\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}, \quad E^\circ = 0.34$$

$$\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}, \quad E^\circ = 0.76$$

In the cell,



In the cell,

$$\begin{aligned} E_{\text{cell}}^\circ &= E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ \\ &= 0.76 - (-0.34) \\ &= 1.10 \text{ V} \end{aligned}$$

52 (a) Here Fe acts as anode while Sn act as cathode. We know that,

$$\begin{aligned} E_{\text{cell}}^\circ &= E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ \\ &= (-0.14) - (-0.44) \\ &= -0.14 + 0.44 \\ &= 0.30 \text{ V} \end{aligned}$$

54 (d)

$$\text{Ag}^+ + e^- \rightarrow \text{Ag}$$

$$\therefore 96500 \text{ C are required to deposit Ag} = 108 \text{ g}$$

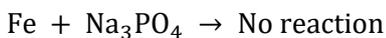
$$\therefore 965 \text{ C are required to deposit Ag} = \frac{108}{96500} \times 965 = 1.08 \text{ g}$$

55 (a) $\Lambda_m = \Lambda_{\text{eq}} \times \text{valency factor};$
For NaCl, valency factor = 1;
Molecular conductivity Λ_m is defined as the conductance of all the ions present in a solution containing 1g molecule in it; Λ_{eq} is defined as the

conductance of all the ions present in a solution containing 1g equivalent in it.

56 (a)

In electrochemical series, iron is placed below sodium, so it cannot displace sodium from its salt solution. Hence, no reaction takes place.



57 (d)

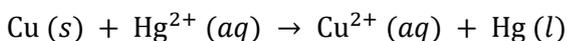
During electrolysis of $\text{NaCl}(aq)$, H^+ ions are discharged at cathode and the pH of solution increases due to decrease in $[\text{H}^+]$.

58 (d)

Galvanic cell is



In the above cell, oxidation of copper and reduction of mercury takes place. Its cell reaction is written as



59 (a)

$$W = \frac{E.i.t}{96500} = \frac{1 \times 0.4 \times 30 \times 60}{96500}$$

$$= 7.46 \times 10^{-3} \text{ g and volume} = \frac{7.46 \times 10^{-3} \times 22.4}{2}$$

$$= 0.0836 \text{ litre}$$

60 (c)

$$E_{\text{cell}}^{\circ} = \frac{2.303RT}{nF} \log K_{eq}$$

$$E_{\text{cell}}^{\circ} = \frac{0.0591}{n} \log K_{eq} \quad [\text{At } 298 \text{ K}]$$

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

$$\therefore \log K_{eq} = 10$$

$$\therefore K_{eq} = 1 \times 10^{10}$$

61 (c)

The Gibb's free energy change ΔG and emf (E°) of a reversible electrochemical cell are related by the following expression.

$$\Delta G = -nFE_{\text{cell}}^{\circ}$$

$$\text{or } = -nFE$$

62 (b)

$$E = E_{\text{RP}}^{\circ} + \frac{0.0591}{n} \log[M^+]$$

Given,

$$E_{\text{RP}}^{\circ} = -2.36 \text{ V}, [M^+] = 0.1 \text{ M}$$

$$n = 1 \text{ (for } M^+ \rightarrow M)$$

$$E = E_{\text{RP}}^{\circ} + \frac{0.0591}{n} \log[M^+]$$

$$= -2.36 + \frac{0.0591}{1} \log 0.1$$

$$= -2.36 + 0.0591 \times (-1)$$

$$= -2.36 - 0.0591$$

$$= -2.419 \text{ V}$$

(d)

1 faraday deposits 1 g equivalent of any substance.

64 (b)

$$\frac{1}{a} = k \times R = 0.002765 \times 400$$

$$= 1.106 \text{ cm}^{-1}.$$

65 (d)

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} + \frac{0.059}{2} \log \frac{[\text{Fe}^{2+}]}{[\text{Zn}^{2+}]}$$

$$0.2905 = E_{\text{cell}}^{\circ} + \frac{0.059}{2} \log \frac{0.01}{0.10}$$

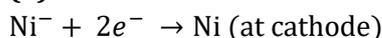
$$\therefore E_{\text{cell}}^{\circ} = 0.32$$

$$\text{No, } E_{\text{cell}}^{\circ} = \frac{0.059}{2} \log_{10} K$$

$$\therefore 0.32 = \frac{0.059}{2} \log_{10} K$$

$$K = 10^{0.32/0.0295}$$

67 (a)



$$\text{Equivalent weight of Ni} = \frac{\text{mol.wt.}}{\text{gain electron}}$$

$$= \frac{58.7}{2}$$

$$= 29.35$$

$$i = 12 \text{ A}, t = 1 \text{ h} = 60 \times 60 \text{ s.}$$

$$Z = \frac{\text{eq. wt.}}{96500}$$

$$\text{Weight of deposit Ni} = \frac{Zit \times \text{efficiency}}{100}$$

$$= \frac{29.35 \times 12 \times 60 \times 60 \times 60}{96500 \times 100}$$

$$= 7.883 \text{ g}$$

68 (a)

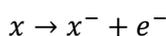
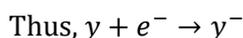
$$\frac{W}{E} = \frac{i \times t}{96500}$$

$$\therefore \frac{W}{E} = 10^{-2} \text{ (Ag is monovalent)}$$

$$\therefore Q = i \times t = 96500 \times 10^{-2} = 965 \text{ C}$$

69 (a)

The tendency to gain electron is in the order $z > y > x$



70 (d)

NaCl , KNO_3 , HCl are strong electrolytes but the size of H^+ is smallest. Smaller the size of the ions, greater is the conductance and hence greater is the conductivity ($\kappa = C \times \text{cell constant}$).

71 (a)

$$\text{Given, } i = 2.5 \text{ A}$$

$$t = 6 \text{ min } 26 \text{ s} = 6 \times 60 + 26 = 386 \text{ s}$$

$$\begin{aligned} \text{Number of coulomb passed} &= i \times t \\ &= 2.5 \times 386 \\ &= 965 \text{ C} \\ \text{Cu}^{2+} + 2e^- &\rightarrow \text{Cu} \\ \therefore 2 \times 96500 \text{ C charge} &\text{deposits Cu} = 63.5 \text{ g} \\ \therefore 965 \text{ C charge} &\text{deposits} \\ \text{Cu} &= \frac{63.5}{2 \times 96500} \times 965 \\ &= 0.3175 \text{ g} \end{aligned}$$

72 (c) Metal placed above in electrochemical series replaces the other from its salt solutions.

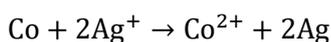
73 (c)

$$\begin{aligned} E_{\text{cell}} &= E_{\text{OP}_{\text{Zn}}}^\circ + E_{\text{RP}_{\text{Cu}}}^\circ + \frac{0.059}{2} \log \frac{[\text{Cu}^{2+}]}{[\text{Zn}^{2+}]} \\ \therefore 1.1 &= 0.78 + E_{\text{RP}_{\text{Cu}}}^\circ + \frac{0.059}{2} \log 1 \\ \therefore E_{\text{RP}_{\text{Cu}^{2+}/\text{Cu}}}^\circ &= 0.32 \\ \therefore E_{\text{RP}_{\text{Cu}^{2+}/\text{Cu}}}^\circ &= -0.32 \text{ V} \end{aligned}$$

74 (d) More the reduction potential, more is the power to get itself reduced or lesser is reducing power or greater is oxidizing power

75 (a) Quantity of current is charge, *i.e.*, coulomb or ampere sec.

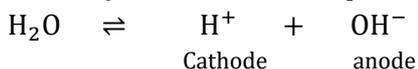
76 (d) Cobalt is anode, *ie*, oxidation takes place on cobalt electrode *ie*, cell reaction is



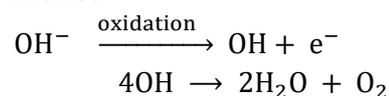
$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{RT}{nF} \ln \frac{[\text{Co}^{2+}]}{[\text{Ag}^+]^2}$$

Thus, less is the factor $\frac{[\text{Co}^{2+}]}{[\text{Ag}^+]^2}$, greater is the E_{cell}°

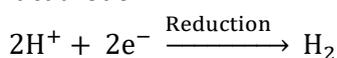
77 (c) Electrolysis of water takes place as follows



At anode



At cathode



Given, time, $t = 1930\text{s}$

Number of moles of hydrogen collected

$$\begin{aligned} &= \frac{1120 \times 10^{-3}}{22.4} \text{ moles} \\ &= 0.05 \text{ moles} \end{aligned}$$

\therefore 1 mole of hydrogen is deposited by = 2 moles of electrons
 \therefore 0.05 moles of hydrogen will be deposited by = 2×0.05
 = 0.10 mole of electrons

Charge, $Q = nF$
 = 0.1×96500

Charge, $Q = it$
 $0.1 \times 96500 = i \times 1930$

$$\begin{aligned} i &= \frac{0.1 \times 96500}{1930} \\ &= 5.0 \text{ A} \end{aligned}$$

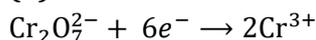
78 (d)

4. $\Delta G^\circ = -nFE_{\text{cell}}^\circ$

5. $E_{\text{cell}}^\circ = \frac{2.303 RT}{nF} \log K_c$

6. $k = Ae^{-E_a/RT}$

79 (b)



Reduction of 1 mol of $\text{Cr}_2\text{O}_7^{2-}$ to Cr^{3+} required 6 moles of electrons. Hence, charge required = $2 \times 96500 \text{ C}$

80 (b)

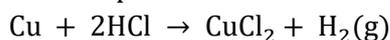
Cell constant = $\frac{l}{a} = \frac{\text{length}}{\text{area}}$
 \therefore unit is cm^{-1} .

81 (a)

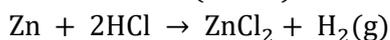
$$\begin{aligned} \lambda &= k \times V = \frac{1}{R} \times \frac{1}{a} \times V \\ &= \frac{1}{210} \times 0.66 \times 100000 \\ &= 314.28 \text{ mho cm}^2 \text{ eq}^{-1} \end{aligned}$$

82 (a)

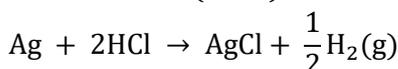
For a reaction to be feasible, the value of E_{cell}° must be positive.



$$\begin{aligned} E_{\text{cell}}^\circ &= E_{\text{H}^+/\text{H}_2}^\circ - E_{\text{Cu}^{2+}/\text{Cu}}^\circ \\ &= 0.00 - (+0.34) = -0.34 \text{ V} \end{aligned}$$



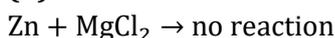
$$\begin{aligned} E_{\text{cell}}^\circ &= E_{\text{H}^+/\text{H}_2}^\circ - E_{\text{Zn}^{2+}/\text{Zn}}^\circ \\ &= 0.00 - (-0.76) = +0.76 \text{ V} \end{aligned}$$



$$\begin{aligned} E_{\text{cell}}^\circ &= E_{\text{H}^+/\text{H}_2}^\circ - E_{\text{Ag}^{2+}/\text{Ag}}^\circ \\ &= 0.00 - (0.80) = -0.80 \text{ V} \end{aligned}$$

Hence, only reaction (ii) is feasible.

83 (d)



This type of reaction does not occur because

$$\text{Mg}^{2+} E^\circ = -2.37 \text{ V, while } \text{Zn}^{2+} E^\circ = -0.76 \text{ V}$$

84 (c)

The ratio of amount deposited during the same charge takes place in the ratio of their equivalent weights and is independent of solution concentration.

85 (a)

Electrical energy is provided during charging of battery.

86 (b)

$$\Delta G^\circ = -nFE^\circ$$

$$-21.2 = -1 \times 96500 \times E$$

$$E = \frac{21.2}{96500} = 0.220 \text{ V}$$

87 (c)

According to Kohlrausch's law

$$\Lambda_{\text{ClCH}_2\text{COOH}}^\infty = \Lambda_{\text{ClCH}_2\text{COO}^-}^\infty + \Lambda_{\text{H}^+}^\infty$$

Given from Kohlrausch law,

$$\Lambda_{\text{ClCH}_2\text{COONa}}^\infty = 224 \Omega^{-1} \text{ cm}^2 \text{ g eq}^{-1}$$

$$\Lambda_{\text{ClCH}_2\text{COONa}}^\infty = \Lambda_{\text{ClCH}_2\text{COO}^-}^\infty + \Lambda_{\text{Na}^+}^\infty \dots \text{(i)}$$

$$\Lambda_{\text{HCl}}^\infty = 203 \Omega^{-1} \text{ cm}^2 \text{ g eq}^{-1}$$

$$\Lambda_{\text{HCl}}^\infty = \Lambda_{\text{H}^+}^\infty + \Lambda_{\text{Cl}^-}^\infty \dots \text{(ii)}$$

$$\Lambda_{\text{NaCl}}^\infty = 38.5 \Omega^{-1} \text{ cm}^2 \text{ g eq}^{-1}$$

$$\Lambda_{\text{NaCl}}^\infty = \Lambda_{\text{Na}^+}^\infty + \Lambda_{\text{Cl}^-}^\infty \dots \text{(iii)}$$

Adding Eqs. (i) and (ii) and subtracting Eq. (iii)

$$\Lambda_{\text{ClCH}_2\text{COO}^-}^\infty + \Lambda_{\text{H}^+}^\infty = \Lambda_{\text{ClCH}_2\text{COONa}}^\infty + \Lambda_{\text{HCl}}^\infty - \Lambda_{\text{NaCl}}^\infty$$

$$= 224 + 203 - 38.5$$

$$= 427 - 38.5$$

$$= 388.5 \Omega^{-1} \text{ cm}^2 \text{ g eq}^{-1}$$

88 (b)

Cu is above Ag in electrochemical series and thus, $\text{Cu} + 2\text{Ag}^+ \rightarrow \text{Cu}^{2+} + 2\text{Ag}$ reaction occurs.

89 (b)

The electrode potential of glass electrode depends only on $[\text{H}^+]$.

90 (a)

$$\Lambda_v = \frac{K \times 1000}{c} = \frac{4.23 \times 10^{-4}}{0.04} \times 1000$$

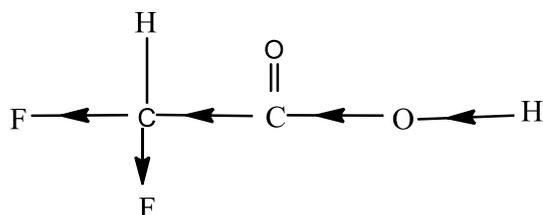
$$\text{Also, } \alpha = \frac{\Lambda_v}{\Lambda^\infty}$$

$$\therefore \Lambda^\infty = \frac{\Lambda_v}{\alpha} = \frac{4.23 \times 10^{-4} \times 1000}{0.04 \times 0.0612}$$

$$= 172.8$$

91 (a)

Fluoro group causes negative inductive effect increasing ionization, thus 0.1 M difluoroacetic acid has highest electrical conductivity.

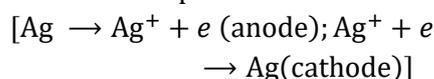


92 (b)

Zn is above iron in electrochemical series.

93 (c)

In case of attacked electrodes, metal dissolves at anode and deposits at cathode.



Thus, concentration of salt does not change.

94 (b)

$$\text{Equivalent conductivity } (\Lambda_{\text{eq}}) = \frac{\kappa \times 1000}{c}$$

$$\text{Conductivity } (\kappa) = \frac{\text{cell constant}}{\text{resistance}} = \frac{1.15}{250} \text{ S cm}^{-1}$$

$$\therefore \Lambda_{\text{eq}} = \frac{1.15 \times 1000}{250 \times 1}$$

$$\Lambda_{\text{eq}} = 4.6 \Omega^{-1} \text{ cm}^2 \text{ equiv}^{-1}$$

95 (a)

$$\text{Given, } E_{\text{Zn}^{2+}/\text{Zn}}^\circ = 0.76 \text{ V}$$

$$E_{\text{Cu}^{2+}/\text{Cu}}^\circ = 0.34 \text{ V}$$

\therefore Zn is anode (\because It has higher oxidation potential)

$$\therefore E_{\text{Zn}^{2+}/\text{Zn}}^\circ = -0.76 \text{ V}$$

$$\text{and } E_{\text{Cu}^{2+}/\text{Cu}}^\circ = -0.34 \text{ V}$$

$$E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ$$

$$= -0.34 \text{ V} - (-0.76 \text{ V})$$

$$= 0.34 \text{ V} + 0.76 \text{ V}$$

$$= 0.42 \text{ V}$$

96 (c)

$$\text{Ionic mobility} = \frac{\text{speed of ions}}{\text{pot.gradient}}$$

97 (c)

$$E = E^\circ - \frac{0.059}{2} \log[\text{H}^+]^2$$

$$= 1.30 - \frac{0.059}{2} \log(10^{-2})^2$$

$$= 1.30 + \frac{0.236}{2} = 1.418 \text{ V}$$

98 (c)

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

$$0.46 = \frac{0.059}{2} \log K_c$$

$$\log K_c = 15.59$$

$$\therefore K_c = 3.9 \times 10^{15}$$

99 (c)

Cl in OCl^- has oxidation number as +1.

Thus, $\text{Cl}^+ + 2e \rightarrow \text{Cl}^-$ (*i.e.*, reduction of OCl^-)

100 (c)

The oxidizing power of $\text{MnO}_4^- / \text{Mn}^{2+}$ couple decreases by 0.38 V.

101 (d)

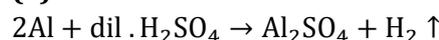
$$E_{\text{cell}}^\circ = \frac{0.059}{n} \log K_c$$

$$0.295 = \frac{0.059}{2} \log K_c; \quad \therefore K_c = 10^{10}$$

102 (b)

Fe being above Cu in electrochemical series and thus, liberates Cu from CuSO_4 .

103 (a)



104 (a)

As 'A' has more E_{red}° value than B, A will act as cathode in the galvanic cell.

Hence,

$$E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ$$

$$= (2.23) - (-1.43)$$

$$= 2.23 + 1.43 = 3.66 \text{ V}$$

105 (d)

H_2SO_4 will furnish maximum H^+ .

106 (c)

Hg is placed below H in electrochemical series.

107 (c)

$$\text{Eq. of H}_2 = \text{Eq. of Al} = \frac{4.5}{27/3} = 0.5$$

$$\therefore 1 \text{ eq. H}_2 = 11.2 \text{ L}$$

$$\therefore 0.5 \text{ eq. H}_2 = 5.6 \text{ L}$$

108 (b)

$$\text{No. of moles of H}_2 = \frac{11.2}{22400}$$

No. of equivalence of hydrogen

$$= \frac{11.2 \times 2}{22400} = 10^{-4}$$

No. of Faradays required = 10^{-4}

\therefore Current to be passed in one second

$$= 96500 \times 10^{-4}$$

$$= 9.65 \text{ A}$$

109 (d)

Nature of ion also includes size, charge on ion.

110 (a)

During rusting, oxidation of iron takes place, *i.e.*, it acts as anode. Hence, coating/connecting iron with metal of lower reduction potential (in comparison to iron) is the best way to prevent its rusting. In this process, the metal with low reduction potential undergoes oxidation (*i.e.* acts as anode) while iron acts as cathode. **(Cathodic protection)**

Use of saline water accelerates the process of rusting thereby increasing the electric conduction of electrolyte solution formed on the metal surface.

111 (b)

$$E_{\text{RP}} = E_{\text{RP}}^\circ + \frac{0.059}{1} \log [\text{H}^+]$$

$$= 0 + 0.059 \times (-3) = -0.177 \text{ V.}$$

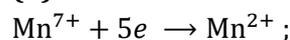
113 (a)

E_{op}° of K > E_{op}° of Al.

114 (d)

In concentration cell net redox change is zero and the decrease in free energy during transfer of matter is responsible for electrical work.

115 (b)



Thus, 5 mole electron = 5 faraday.

116 (b)

$$E_{\text{cell}} = E_{\text{cell}}^\circ + \frac{0.0592}{n} \log \frac{[\text{Cu}^{2+}]}{[\text{T}^{1+}]^2}$$

$$n = 2$$

\therefore According to above equation E_{cell} can be increased by increasing $[\text{Cu}^{2+}]$.

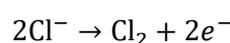
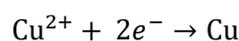
117 (c)

Only Zn and Fe are above H;

Also Fe^{3+} can be reduced to Fe^{2+} by H.

118 (b)

The cathode and anode reactions respectively are



The two moles of electrons have been transferred from anode to cathode to produce Cu and Cl_2 in a mole ratio of 1:1. Thus, 2F electricity is required

119 (b)

On dilution, ionic mobility increases but number of ions present in 1 mL decreases; Thus, only conductivity decreases and rest all increases.

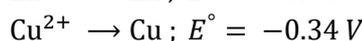
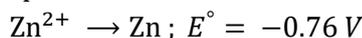
120 (d)

$$E_{\text{cell}} = E_{\text{OP}_{\text{anode}}} + E_{\text{RP}_{\text{cathode}}}$$

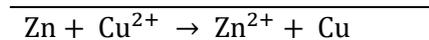
$$= E_{\text{OP}_{\text{anode}}} - E_{\text{OP}_{\text{cathode}}}$$

121 (b)

Electrode potential of cell must be positive for spontaneous reaction.



Redox reaction is



$$E_{\text{cell}} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$$

$$= -0.34 - (-0.76)$$

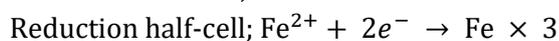
$$= +0.42 \text{ V}$$

E_{cell} is positive, so above reaction is feasible.

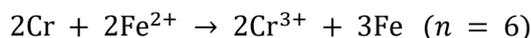
122 (c)

Among given elements, D has the minimum reduction potential (-2.37 V) hence, it can displace all other from their salts.

123 (a)



Net cell reaction;



$$E_{\text{cell}}^{\circ} = E_{\text{oxidation}}^{\circ} - E_{\text{reduction}}^{\circ}$$

$$= 0.72 - 0.42$$

$$= 0.30 \text{ V}$$

$$E_{\text{cell}}^{\circ} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log \frac{[\text{Cr}^{3+}]^2}{[\text{Fe}^{2+}]^3}$$

$$= 0.30 - \frac{0.0591}{6} \log \frac{(0.1)^2}{(0.01)^3}$$

$$= 0.30 - \frac{0.0591}{6} \log \frac{10^{-2}}{10^{-6}}$$

$$= 0.30 - \frac{0.0591}{6} \log 10^4$$

$$E_{\text{cell}} = 0.2606 \text{ V}$$

124 (c)

A thin film of Cr_2O_3 is formed on Cr Surface.

125 (b)

The unit of electrochemical equivalent (Z) is g/C.

$$w = Z \cdot i \cdot t$$

$$\therefore Z = \frac{w}{i \cdot t} \text{ g/C}$$

126 (d)

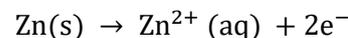
The elements which are below H_2 in electrochemical series, cannot displace H_2 .

\therefore Out of Li^+ , Sr^{2+} , Al^{3+} and Ag^+ , Ag^+ is below H_2 in electrochemical series, so Ag^+ cannot displace H_2 .

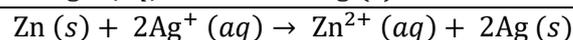
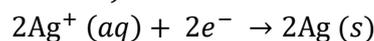
127 (b)

As the reduction potential of Zn is less than that of Ag, hence Zn will act as anode when a cell is made using them.

Hence, the correct reaction will be



(oxidation)



128 (a)

$$W \propto i \times t \text{ and } W = Z \times i \times t.$$

129 (b)

$$\text{Cell constant} = \frac{k}{C} = 0.0212 \times 55$$

$$= 1.166 \text{ cm}^{-1}$$

130 (c)

Reducing power, *ie*, the tendency to lose electrons increases as the reduction potential decreases

131 (a)

$$7. \quad \text{Reducing character} \propto \frac{1}{\text{reduction potentials}}$$

8. Oxidizing power of halogen decreases from F_2 to I_2 because their reduction potentials decrease from fluorine to iodine.

9. The reducing power of hydrogen halides increases from hydrogen chloride to hydrogen iodide since, the stability of the $\text{H}-\text{X}$ bond decreases in the same order. Hence, all statements are correct.

132 (d)

$$\text{If } E^{\circ} = 0, \text{ then } \Delta G^{\circ} = -nE^{\circ}F = 0.$$

133 (d)

$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$$

$$\therefore 2.46 = (+0.80) - E_{\text{Al}^{3+}/\text{Al}}^{\circ}$$

$$\text{Or } E_{\text{Al}^{3+}/\text{Al}}^{\circ} = 0.80 - 2.46 = -1.66 \text{ V}$$

134 (d)

$$E^{\circ} \text{ for reaction in (d)} = E_{\text{OP}_{\text{Br}}}^{\circ} + E_{\text{RP}_1}^{\circ} = -1.09 + (-0.54)$$

$$= -1.63 \text{ V}$$

Since, E° is negative and thus, reaction is non-spontaneous.

136 (b)

$$\Delta G^{\circ} = -nFE^{\circ}$$

$$\Delta G^\circ = -2.303 RT \log K_c$$

$$\therefore nFE^\circ = 2.303 RT \log K_c$$

$$\log K_c = \frac{nFE^\circ}{2.303 RT}$$

$$= \frac{2 \times 96500 \times 0.295}{2.303 \times 8.314 \times 298}$$

$$\log K_c = 9.97$$

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

137 (c)

The molar conductivity of potassium hexacyanoferrate (II) *i. e.*, $K_4[Fe(CN)_6]$ is highest because it gives maximum number of ions on ionization.



139 (a)

The metals having higher negative value of standard reduction potential are placed above hydrogen in electrochemical series. The metals placed above hydrogen has a great tendency to donate electrons or oxidising power. The metals having great oxidizing power are strongest reducing agent. Zn has higher negative value of standard reduction potential. Therefore, it is the strongest reducing agent.

140 (d)

$$w = 60 \text{ g}$$

$$i = 5A$$

$$\text{Equivalent weight of Ca} = \frac{\text{atomic weight}}{\text{valency}}$$

$$= \frac{40}{2} = 20$$

According to first law of Faraday electrolysis

$$w = Zit = \frac{\text{equivalent weight}}{96500} \times i \times t$$

$$\therefore 60 = \frac{20}{96500} \times 5 \times t$$

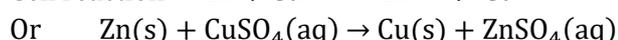
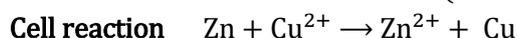
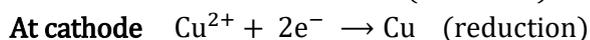
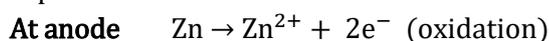
$$t = \frac{96500 \times 60}{20 \times 5} \text{ s}$$

$$= \frac{96500 \times 60}{20 \times 5 \times 60 \times 60} \text{ h}$$

$$= 16.08 \text{ h}$$

142 (b)

In Galvanic cell (Daniel cell) the electrical energy is produced from chemical reactions.



143 (b)

$$\Lambda_{AcOH}^\infty = \Lambda_{AcONa}^\infty + \Lambda_{HCl}^\infty - \Lambda_{NaCl}^\infty$$

$$= 91.0 + 426.2 - 126.5$$

$$= 390.7$$

144 (b)

The metal with more E_{OP}° is oxidised.

145 (d)

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.05915}{n} \log Q$$

For standard hydrogen electrode,

$$E_{\text{cell}}^\circ = 0.00V$$

$$\therefore E_{\text{cell}} = - \frac{0.05915}{n} \log Q$$

Given, pH = 1.0

$$\therefore [H^+] = 1 \times 10^{-1}$$

$$E_{\text{cell}} = - \frac{0.05915}{n} \log \frac{1}{[H^+]}$$

[\therefore The reaction occurring is $2H^+ + 2e^- \rightarrow H_2$]

$$= + \frac{0.05915}{1} \log(H^+)$$

$$= 0.05915 \log(10^{-1})$$

$$= -0.05915 V$$

$$= -59.15 \text{ mV}$$

146 (c)

$$\Lambda_{\text{eq}}^\circ = \kappa \times \frac{1000}{\text{normality}}$$

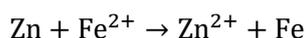
$$= \frac{0.005 \times 1000}{0.01} = 500 \text{ ohm}^{-1} \text{ cm}^2 \text{ equiv}^{-1}$$

147 (a)

$$E_{OP}^\circ \text{ of Mg} > E_{OP}^\circ \text{ of Al.}$$

148 (a)

For the given cell, reaction is



$$E = E^\circ - \frac{0.0591}{n} \log \frac{C_1}{C_2}$$

$$\text{or, } E^\circ = E + \frac{0.0591}{n} \log \frac{C_1}{C_2}$$

$$= 0.2905 + \frac{0.0591}{2} \log \frac{10^{-2}}{10^{-3}} = 0.32 V$$

$$E^\circ = \frac{0.0591}{2} \log K_c$$

$$\therefore \log K_c = \frac{0.32 \times 2}{0.0591} = \frac{0.32}{0.0295}$$

$$K_c = 10^{\frac{0.32}{0.0295}}$$

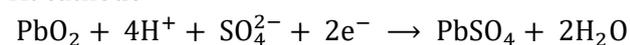
149 (d)

When Allead storage battery is discharged, the following cell reactions take place.

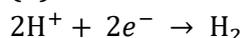
At anode



At cathode



150 **(d)**



According to Nernst equation,

$$E = E^\circ + \frac{0.0591}{n} \log \frac{1}{[\text{H}^+]^2}$$

$$E = 0 - \frac{0.0591}{2} \log [\text{H}^+]^2 \\ = -0.0591 \text{ pH}$$

151 **(a)**

$$E_{\text{Fe}^{2+}/\text{Fe}}^\circ = -0.441 \text{ V}$$

$$E_{\text{Fe}^{3+}/\text{Fe}}^\circ = -0.771 \text{ V}$$

$$E_{\text{cell}}^\circ = E_{\text{OP}_{\text{Fe}/\text{Fe}^{2+}}}^\circ + E_{\text{RP}_{\text{Fe}^{3+}/\text{Fe}^{2+}}}^\circ \quad (\text{See redox change}) \\ = +0.441 + 0.771 = 1.212 \text{ V}$$

152 **(b)**

$$E_{\text{OP}_{\text{Zn}}}^\circ > E_{\text{OP}_{\text{Cu}}}^\circ \text{ or } E_{\text{RP}_{\text{Zn}}}^\circ < E_{\text{RP}_{\text{Cu}}}^\circ$$

153 **(b)**

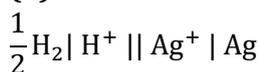
H_2SO_4 is strong electrolyte.

154 **(c)**

$$\Lambda_v = \frac{\Lambda^0}{100}$$

$$\therefore \alpha = \frac{\Lambda_v}{\Lambda^0} = \frac{\Lambda^0}{100\Lambda^0} = 0.01$$

155 **(b)**



$$E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ$$

$$= E_{\text{Ag}^+/\text{Ag}}^\circ - E_{\text{H}^+/\frac{1}{2}\text{H}_2}^\circ$$

$$= (0.80) - (0.0) = 0.80 \text{ V}$$

156 **(a)**

Ions move towards opposite electrodes due to coulombic forces of attraction.

157 **(c)**

More is E_{RP} , more is the tendency to get reduced.

E_{RP} for Ag is maximum.

158 **(d)**

E_{OP} for Li/Li⁺ is maximum in these.

159 **(b)**

$$250\text{mL of } 1 \text{ M AgNO}_3 \text{ contain} = \frac{250}{1000} \\ = 0.25 \text{ mole AgNO}_3$$

∴ Electricity required to liberate 1 g equivalent of metal

$$= 96500 \text{ C}$$

∴ Electricity required to liberate 0.25 g equivalent of metal

$$= \frac{96500 \times 0.25}{1} \\ = 24125 \text{ C}$$

160 **(b)**

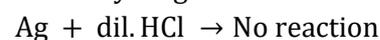
1 faraday = 1 eq. of Cu = 1/2 mole Cu

$$= N/2 \text{ atoms of Cu.}$$

161 **(a)**

The metals, present below hydrogen in the electrochemical series, cannot liberate hydrogen from the dilute acids.

Among the given metal only Ag is present below hydrogen in electrochemical series, so it does not evolve hydrogen with dil. HCl.



162 **(b)**

Any cell (like fuel cell), works when potential difference is developed.

163 **(a)**

$$t_c = \frac{u_c}{u_a + u_c}, t_a = \frac{u_a}{u_a + u_c}$$

Where, u_a and u_c are speed of ion and t_c and t_a are transport number of cation and anions respectively of an electrolyte.

$$\text{Thus, } t_c + t_a = 1$$

164 **(b)**

We know that 1 Faraday charge liberates one gram-equivalent of a metal, hence 0.5 F charge will liberate

$$= 0.5 \times 23$$

$$= 11.50 \text{ g of sodium (E} = 23)$$

165 **(b)**

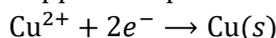
$$\text{Current (i)} = 1.5 \text{ A}$$

$$\text{Time (t)} 10 \text{ min} = 10 \times 60 = 600 \text{ s}$$

$$\text{Quantity of electricity passed } Q = i \times t \\ = (1.5 \text{ A}) \times (600 \text{ s})$$

$$= 900 \text{ C}$$

Copper is deposited as



2 moles of electrons or $2 \times 96500 \text{ C}$ of current deposit copper = 63.56 g

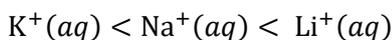
900 C of current will deposit copper

$$= \frac{63.56}{2 \times 96500} \times 900$$

$$= 0.296 \text{ g}$$

166 (b)

Ionic mobility depends upon the charge to size ratio of ion. The ionic size in case of hydrated cation is



167 (b)

Eq. of A = Eq. of B = Eq. of C

$$\text{or } \frac{2.1}{7/n_1} = \frac{2.7}{27/n_2} = \frac{7.2}{48/n_3}$$

$$0.3 n_1 = 0.1 n_2 = 0.15 n_3$$

$$\therefore n_1 = \frac{n_2}{3} = \frac{n_3}{2}$$

$$\text{If } n_1 = 1 \text{ then } n_2 = 3, \quad n_3 = 2$$

168 (d)

The electrode, which shows colour change during redox process is called indicator electrode.

169 (c)

Molar conductivity or molar conductance

$$(\Lambda_m) = \kappa \times V$$

$$\Lambda_m = \kappa \times \frac{1000}{C_m}$$

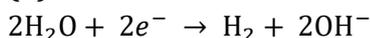
Where, C_m is molar concentration (mol L^{-1})

$$\therefore \text{Molar conductance } (\Lambda_m) \propto \left(\frac{1}{C}\right)$$

170 (b)

Rusting of iron is catalyzed by moist air.

171 (b)



For 0.01 mole H_2 , 0.02 mole of electrons are consumed charge required

$$= 0.02 \times 96500 \text{ C} = i \times t$$

$$\text{Time required} = \frac{0.02 \times 96500}{10 \times 10^{-3}} = 19.3 \times 10^4 \text{ s}$$

172 (b)

Metal having higher E_{OP}° replaces the other from its solution.

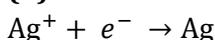
173 (a)

Eq. of Ag = Eq. of H_2 ;

$$\frac{W}{108} = \frac{5600 \times 2}{22400 \times 1}$$

$$\therefore W_{\text{Ag}} = 54 \text{ g}$$

174 (b)



$$9650 \text{ C} = 0.1 \text{ F} = 0.1 \text{ equivalent Ag}$$

$$= 0.1 \text{ mol Ag}$$

$$= 10.8 \text{ g Ag}$$

175 (d)

More or +ve is E_{OP}° for an electrode more is its reducing power and *vice - versa*.

177 (b)

BeCl_2 is predominantly more covalent among halides of alkaline earth metals.

178 (b)

In CuSO_4 , change is $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$;

In CuCN , change is $\text{Cu}^+ + e^- \rightarrow \text{Cu}$;

Thus, $W \propto E_{\text{Cu}}$, which is more in CuCN .

179 (c)

	E°	nE°
$\text{Mn}^{2+} + 2e^- \rightarrow \text{Mn}$	-1.18 V	-2.36
$\text{Mn}^{3+} + e^- \rightarrow \text{Mn}^{2+}$	1.51 V	1.51 V
$\text{Mn}^{3+} + 3e^- \rightarrow \text{Mn}$	-0.28	-0.85

180 (a)

$2\text{H}^- \rightarrow \text{H}_2 + 2e^-$; Hydrogen in CaH_2 is -ve.

181 (c)

1 mole of monovalent metal ion means charge of N electrons

i. e., 96500 C or 1 faraday.

182 (a)

For strong electrolytes Λ vs \sqrt{c} plots are straight line.

183 (d)

The metal should be capable of adsorbing H (*e.g.*, Pt).

184 (b)

Laws of electrolysis were proposed by Michael Faraday in 1833.

10. **Faraday's first law** "The mass of any substance deposited or liberated at any electrode is directly proportional to the quantity of electricity passed".

$$\text{i. e., } w \propto Q$$

where, w = mass of ions liberated in gram

Q = quantity of electricity passed in coulombs.

11. **Faraday's second law** "When the same quantity of electricity is passed through different electrolytes, the masses of different ions liberated at the electrodes are directly proportional to their chemical equivalents."

$$\text{i. e., } \frac{w_1}{w_2} = \frac{E_1}{E_2} \text{ OR } \frac{Z_1 it}{Z_2 it} = \frac{E_1}{E_2}$$

185 (d)

Strong electropositive metals (I group, II group and Al) cannot be obtained at cathode by electrolysing their aqueous salt solutions.

186 (b)

$$\frac{\text{Weight of Cu}}{\text{Weight of H}_2} = \frac{\text{Eq. wt. of Cu}}{\text{Eq. wt. of H}_2}$$

$$\frac{\text{Weight of Cu}}{0.504} = \frac{63.6/2}{1}$$

∴ Weight of Cu = 15.9 g

187 (b)

In presence of Hg electrode preferential discharge of Na⁺ (in comparison to H⁺) occurs.

188 (b)

Specific conductivity of a solution decreases with dilution.

189 (b)

Passage of current in electrolytic solution is due to migration of ions towards opposite electrodes.

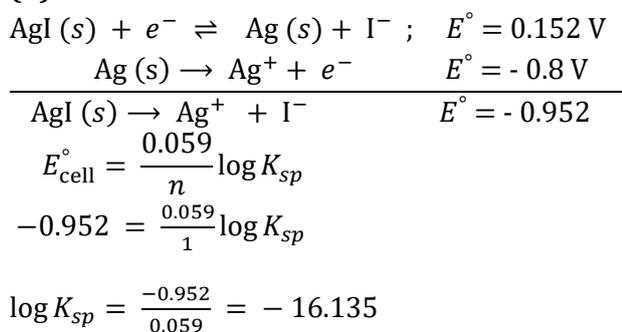
190 (c)

$$E_{\text{cell}} = E_{OP_L} + E_{RP_R} = -E_{RP_L} + E_{RP_R}$$

191 (a)

$E^\circ = E$, when $[Zn^{2+}] = 1M$;
Also process is $Zn^{2+}(aq) + 2e \rightarrow Zn(s)$.

192 (d)



193 (a)

$$\Lambda_{\text{CH}_3\text{COOH}}^\infty = \Lambda_{\text{CH}_3\text{COONa}}^\infty + \Lambda_{\text{HCl}}^\infty - \Lambda_{\text{NaCl}}^\infty$$

194 (a)

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

195 (d)

$$\frac{\text{Wt. of Cu deposited}}{\text{Wt. of H}_2 \text{ produced}} = \frac{\text{eq. wt. of Cu}}{\text{eq. wt. of H}}$$

$$\frac{0.16}{\text{wt. of H}_2} = \frac{64/2}{1} = \frac{32}{1}$$

$$\text{Wt. of H}_2 = \frac{0.16}{32} = 5 \times 10^{-3} \text{ g}$$

Volume of H₂ liberated at STP

$$= \frac{22400}{2} \times 5 \times 10^{-3} \text{ cc}$$

$$= 56 \text{ cc}$$

196 (a)

Faraday's laws are independent of external factors.

197 (a)

E_{OP}° for Li is more,
 $Li \rightarrow Li^+ + e$

Thus, Li is strong oxidant.

198 (d)

All are electrolytic cells.

199 (a)

96500 C or 1F will liberate 1 eq. of O₂ or 1/4 mole O₂ or 5.6 litre O₂ at NTP.

200 (a)

96500C or 1 Faraday charge is required for the deposition of 1 g-equivalent of a substance.

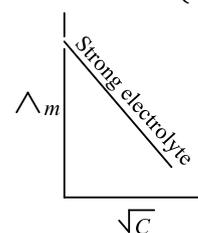
201 (c)

$$\text{Equivalent conductance} = \frac{1000 \times \text{conductance} \times \text{cell constant}}{\text{normality}}$$

So, units are, $\Omega^{-1} \text{ cm}^2 \text{equiv}^{-1}$ or $S \text{ cm}^2 \text{equiv}^{-1}$.

202 (b)

For strong electrolytes the plot of molar conductance (Λ_m) vs. \sqrt{C} is linear.



Variation of molar conductance

(Λ_m) with \sqrt{C} for strong electrolyte.

203 (b)

$$\begin{aligned} \Lambda_{\text{eq}}^\infty(\text{NH}_4 \text{OH}) &= \Lambda_{\text{eq}}^\infty(\text{NH}_4\text{Cl}) + \Lambda_{\text{eq}}^\infty(\text{NaOH}) - \Lambda_{\text{eq}}^\infty(\text{NaCl}) \\ &= (149.74 + 248.1 - 126.4) \\ &= 271.44 \Omega^{-1} \text{ cm}^2 \text{eq}^{-1} \end{aligned}$$

204 (c)

$$\begin{aligned} E_{\text{cell}}^\circ &= E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ \\ E_{\text{Ag}^+/\text{Ag}}^\circ - E_{\text{Cu}^+/\text{Cu}}^\circ \\ &= -0.80 - 0.34 \\ &= +0.46 \end{aligned}$$

205 (c)

$$\begin{aligned} E_{\text{cell}} &= E_{OP_{Ni/Ni^{2+}}} + E_{RP_{Au^{3+}/Au}} \\ &= E_{OP_{Ni}}^\circ - \frac{0.059}{2} \log[Ni^{2+}] + E_{RP_{Au}}^\circ + \frac{0.059}{3} \log[Au^{3+}] \\ &= 0.25 - \frac{0.059}{2} \log(1.0) + 1.50 + \frac{0.059}{3} \log 1.0 = 1.75 \text{ V} \end{aligned}$$

206 (d)

The metal placed below in electrochemical series does not react with that metal salt solution which metal is placed above in series.

208 (a)

Cell representation is done as follows

Anode | Anodic electrolyte || cathodic electrolyte | cathode

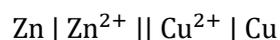
(i) Oxidation is loss of electron and it takes place at anode. Reduction is gain of electron and it takes place at cathode.

∴ For cell reaction,



Zn is anode and Cu is cathode.

∴ Cell representation is



209 (c)

2×96500 C electricity is used to liberate

= 22400 mL O_2 at STP

∴ 9.65×1000 C electricity will liberate

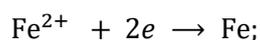
$$= \frac{22400 \times 9.65 \times 1000}{2 \times 96500}$$

= 1120 mL

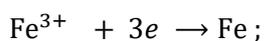
210 (d)

Oxidation half-cell

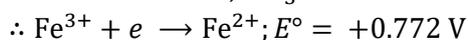
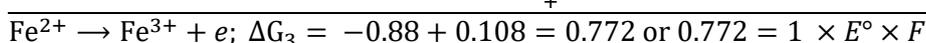
214 (d)



$$-\Delta G_1 = 2 \times (-0.44) \times F$$



$$-\Delta G_2 = 3 \times (-0.036) \times F$$

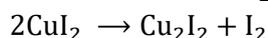
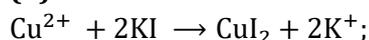


Above procedure should be used only when two half reactions on algebraic sum give a third half reaction.

215 (b)

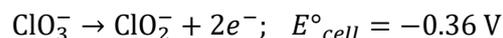
Chromium is more electropositive metal than iron. In stainless steel, chromium forms an oxide layer and thus it protects steel from corrosion.

216 (b)

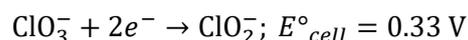


217 (d)

Salt bridge is used to remove or eliminate liquid junction potential arising due to different relative speed of ions of electrolytes at the junction of two



Reduction half-cell



$$E^\circ_{\text{cell}} = 0.33 - 0.36 = -0.03 = \frac{RT}{2F} \ln K$$

$$\text{or } -0.03 = \frac{0.059}{2} \log K \text{ or } K = 0.1$$



$$0.1 - 2x \quad x \quad x$$

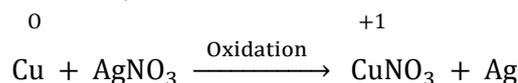
$$\frac{x^2}{(0.1 - 2x)^2} = 0.1$$

$$\text{or } x = 1.9 \times 10^{-2}$$

211 (a)

Cu is placed above Ag in electrochemical series, hence it can replace Ag from its salts solution.

Therefore, the reaction occurs as follows



212 (a)

E° does not depend on stoichiometry of change.

213 (d)

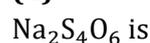
HCl is strong electrolyte and H^+ has highest conducting power due to Grothus conductance.

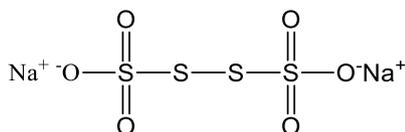
electrolytes in an electrochemical cell. Thus, a salt bridge such as KCl is placed in between two electrolytes. A salt used for this purpose should have almost same speeds of its cation and anion.

218 (b)

$$w_{\text{Ag}} = \frac{E_{\text{Ag}} \times Q}{96500} = \frac{108 \times 9.65}{96500} = 1.08 \times 10^{-2} \text{ g} \\ = 10.8 \text{ mg}$$

219 (d)





The two S atoms which are linked to each other have 0 oxidation number. The oxidation number of other two S-atoms can be calculated as

$$2x + 2 \times 0 + 6 \times -2 = -2$$

$$2x = 12 - 2 = 10$$

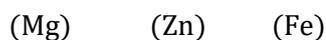
$$x = +5$$

220 (a)

Higher the negative value of E° , more is the reducing power.

The order of E° values (negative value) is

$$-2.37 > -0.76 > -0.44$$



\therefore Mg can reduce both Zn^{2+} and Fe^{2+} . Zn can reduce Fe^{2+} , but not Mg^{2+} . Fe cannot reduce Mg and Zn but can oxidize them.

221 (c)

It is definition of Kohlrausch's law.

222 (c)

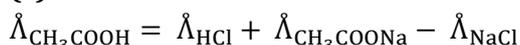


$$E_{\text{cell}} = E^\circ - \frac{0.059}{n} \log \frac{[\text{Sn}^{4+}]}{[\text{Sn}^{2+}]}$$

$$= E^\circ - \frac{0.059}{2} \log \left[\frac{0.01}{0.1} \right]$$

$$= E^\circ + \frac{0.059}{2}$$

224 (c)



225 (b)

Reduction always occurs at cathode during electrolysis.

226 (a)

$$E_{\text{cell}}^\circ = \frac{0.0591}{2} \log K_c$$

$$0.59 = \frac{0.0591}{2} \log K_c$$

$$\frac{0.59 \times 2}{0.059} = \log K_c$$

$$\therefore \log K_c = 20$$

$$K_c = \text{antilog } 20 = 10^{20}$$

227 (c)

Given, current = 241.25 C

We know that 1 C electricity will deposit 1.118×10^{-3} g of silver.

\therefore 241.25 C electricity will deposit

$$= (1.118 \times 10^{-3}) \times 241.25$$

$$= 0.27 \text{ g of silver.}$$

228 (a)

$$\text{CaF}_2 = k_{\text{solution}} - k_{\text{water}}$$

229 (a)

More negative is the standard reduction potential, greater is the tendency to lose electrons and hence, greater is the reactivity

230 (a)

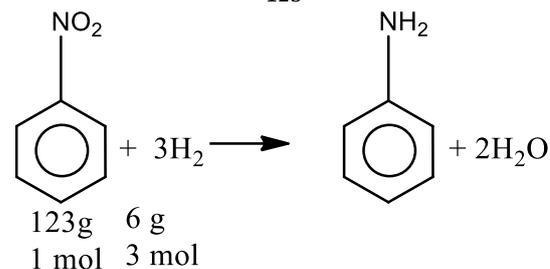
Follow Debye Hückel theory of strong electrolytes.

231 (d)

Given, mass of nitrobenzene = 12.3 g

Molecular mass of nitrobenzene = 123 g

$$\therefore \text{Moles} = \frac{12.3}{123} = 0.1$$



\therefore H_2 required for reduction of 0.1 mole nitrobenzene

$$= 0.1 \times 3 = 0.3 \text{ mol of hydrogen}$$

\therefore Amount of charge required to liberate 0.3 mol of hydrogen

$$= 2 \times 96500 \times 0.3 (\because \text{H}_2 \rightarrow 2\text{H}^+)$$

$$= 57900 \text{ C}$$

232 (a)

$$\text{Cell constant} = \frac{l}{a}$$

Where, l = distance between the electrode
 a = area of the electrode

233 (d)

More is solute-solute interaction, lesser is conductance. More is solute-solvent interaction, more is conductance. An increase in temperature also increases conductance due to increase in ionic mobility.

234 (a)

$$E_{\text{Zn}/\text{Zn}^{2+}}^\circ > E_{\text{H}/\text{H}^+}^\circ$$

235 (b)

H_2SO_4 is strong electrolyte and thus, ionises completely.

236 (a)

Standard electrode potential of hydrogen electrode is zero.

237 (c)

HCl is an electrolyte.

238 (b)

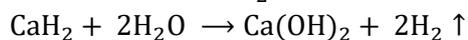
Use of electrolysis is not done in the production of water.

239 (a)

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

240 (a)

The ionic hydrides conduct electricity in molten state and liberate H_2 anode.



241 (d)

Specific conductance is defined as the conductance of one centimeter cube (cc) of the solution of an electrolyte. Upon dilution, the concentration of ions per cc decreases, so the specific conductance falls. Thus, specific conductance \propto concentration of electrolytes, hence specific conductance is least for the solution for 0.002 N concentration.

242 (c)

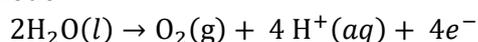
According to Faraday law, number of ions produced \propto quantity of electricity passed

243 (c)

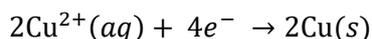
$$W = \frac{E \times i \times t}{96500}$$
$$\therefore 1.8 = \frac{E \times 3 \times 50 \times 60}{96500}$$
$$\therefore E = 19.3$$

244 (a)

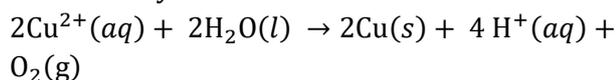
At anode



At cathode



Net electrolysis reaction is



So, H_2O is obtained.

245 (a)

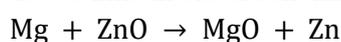
During the electrolysis of fused NaCl, chloride ions are oxidized at anode and this process is called oxidation

246 (d)

Since, reduction occurs at Ag electrode hence, increase in the concentration of Ag^+ or decrease in concentration of Cu^{2+} will increase the voltage.

247 (a)

Magnesium is more electropositive than zinc, hence it can reduce Zn^{2+} (magnesium is placed above zinc in electrochemical series).

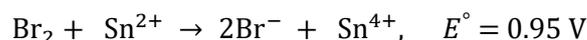


248 (c)

It is better to write unit as $kg\ coulomb^{-1}$ in place of $g\ coulomb^{-1}$.

249 (b)

Cell reaction will be



$$E_{cell}^\circ = \frac{0.059}{2} \log K_{eq}$$

$$0.95 = \frac{0.059}{2} \log K_{eq}$$

$$\frac{0.95 \times 2}{0.059} = \log K_{eq}$$

$$K_{eq} \approx 10^{32}$$

250 (b)

It is either $ohm^{-1}m^{-1}$ or Siemens m^{-1} , i. e., $S\ m^{-1}$.

251 (a)

Ionic mobility

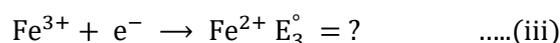
$$\mu_{Ag^+} = \frac{\text{ionic conductance}}{96500}$$
$$= \frac{5 \times 10^{-4}}{96500}$$
$$= 5.2 \times 10^{-9} cm/s$$

252 (c)

Given,



We need to calculate



We can obtain the (III) by subtracting II from I but E_3° , we can not obtain that way because electrode potential is intensive property. That's when we determine E_3° calculating

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

(ΔG is an extensive property)

$$\Delta G_3 = 3 \times 0.036\ F - 2 \times 0.439\ F$$

$$\Delta G_3 = 0.108\ F - 0.878\ F$$

$$-1 \times F \times E_3^\circ = -0.770\ F$$

$$E_3^\circ = 0.770\ F$$

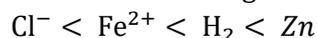
253 (b)

$$E_{cell}^\circ = E_{OP_{Sn}}^\circ + E_{RP_{Pb}}^\circ$$

or $E_{RP_{Sn}}^\circ < E_{RP_{Pb}}^\circ$, then Sn will oxidise to reduce Pb^{2+} .

254 (d)

The substances which have lower reduction potentials are stronger reducing agents. Hence, the order of strength of reducing agent is increases in the following order

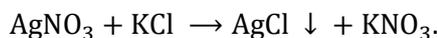


strength of reducing agent increases

255 (d)

$$\begin{aligned}\Lambda_m(\text{CH}_3\text{COOH}) &= \Lambda_m(\text{CH}_3\text{COONa}) + \Lambda_m(\text{HCl}) \\ &- \Lambda_m(\text{NaCl}) \\ &= 91.0 + 425.9 - 126.4 \\ &= 516.9 - 126.4 \\ &= 390.5 \text{ S cm}^2 \text{ mol}^{-1}\end{aligned}$$

256 (c)



257 (a)

$$\begin{aligned}\Lambda_m^\infty \text{ for BaCl}_2 &= \Lambda_m^\infty \text{ Ba}^{2+} + 2\Lambda_m^\infty \text{ Cl}^- \\ \therefore \Lambda_{\text{eq}}^\infty \text{ for BaCl}_2 &= 1/2\Lambda_m^\infty \text{ Ba}^{2+} + \Lambda_m^\infty \text{ Cl}^- \\ &= 127/2 + 76 \\ &= 139.5 \Omega^{-1} \text{ cm}^2\end{aligned}$$

258 (d)

$$\begin{aligned}-\Delta G &= nEF \\ -966 \times 10^3 &= 4 \times E \times 96500 \\ \therefore E &= -2.5 \text{ V}\end{aligned}$$

Thus, an e.m.f. of + 2.5 is needed to carry out the electrolytic reduction of Al_2O_3 .

259 (c)

In Galvanic cell the electrical energy is produced from chemical reaction. *i. e.*, chemical energy is transformed into electrical energy.

260 (d)

Faraday's laws are independent of all other external factors and $W \propto Q$.

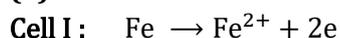
261 (b)

$$\begin{aligned}\Lambda &= k \times \frac{1000}{M} \\ &= C \times \frac{1}{a} \times \frac{1000}{M} \\ \therefore C &= \frac{\Lambda \times a \times M}{l \times 1000} \\ \text{i. e., } C &\propto \text{conc.} \\ &\propto a \\ &\propto \frac{1}{l} \\ \therefore \text{unit of constant } (\Lambda) &= \frac{C \times l \times 1000}{a \times M} \\ &= \frac{\text{S} \times \text{m}}{\text{m}^2 \times \text{mole m}^{-3}} \\ &= \text{S m}^2 \text{ mol}^{-1}\end{aligned}$$

262 (c)

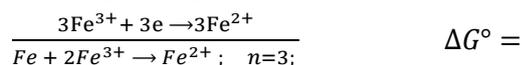
Cu can displace Ag from AgNO_3 but it cannot liberate H_2 from HCl because $E_{\text{OPCu}}^\circ > E_{\text{OPAg}}^\circ$ and $E_{\text{OPCu}}^\circ = -\text{ve}$.

263 (a)



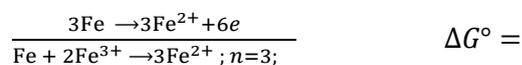
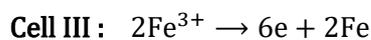
$$\Delta G^\circ =$$

$$-nE^\circ F$$



$$\Delta G^\circ =$$

$$-nE^\circ F$$



$$\Delta G^\circ =$$

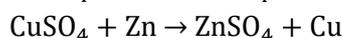
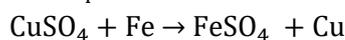
$$-nE^\circ F$$

n and E° are different for each cell.

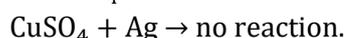
264 (c)

CuSO_4 will react with elements placed above it in electrochemical series and it would not react with elements placed below it in electrochemical series.

$\therefore \text{CuSO}_4$ reacts with Zn and Fe placed above it.



$\therefore \text{CuSO}_4$ does not react with Ag, placed below it.



265 (c)

Rust is Fe_2O_3 and $\text{Fe}(\text{OH})_3$.

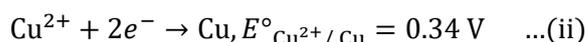
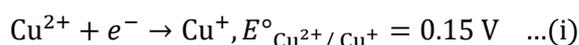
267 (c)

$$W = \frac{E \times i \times t}{96500}$$

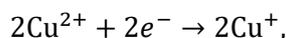
$$\therefore P = \frac{E \times c \times t}{96500}$$

268 (a)

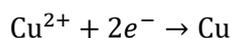
The required reaction ($\text{Cu}^{2+} + \text{Cu} \rightarrow 2\text{Cu}^+$) can be obtained by using the following reactions



Multiplying Eq. (i) by 2, we get



$$\Delta G_1 = -nFE = -2 \times F \times 0.15 \quad \dots(\text{iii})$$



$$\Delta G_2 = -nFE = -2 \times F \times 0.34 \quad \dots(\text{iv})$$

Subtract the Eq. (iv) from Eq. (iii)



$$\Delta G_3 = -nFE = -1 \times F \times E^\circ$$

$$\text{Also } \Delta G_3 = \Delta G_1 - \Delta G_2$$

$$\therefore -1FE^\circ = (-2F \times 0.15) - (-2F \times 0.34)$$

$$E^\circ = -0.38$$

This is the value for the reaction



But the given reaction is just reverse of it

$$\therefore E_{\text{cell}} \text{ for given reaction} = +0.38 \text{ V}$$

269 (d)

According to Kohlrausch's law



$$\Lambda_{\text{CH}_3\text{COOH}}^\circ = \Lambda_{\text{CH}_3\text{COO}^-}^\circ + \Lambda_{\text{H}^+}^\circ \dots \text{(i)}$$

$$\Lambda_{\text{HCl}}^\circ = \Lambda_{\text{H}^+}^\circ + \Lambda_{\text{Cl}^-}^\circ \dots \text{(ii)}$$

$$\Lambda_{\text{CH}_3\text{COONa}}^\circ = \Lambda_{\text{CH}_3\text{COO}^-}^\circ + \Lambda_{\text{Na}^+}^\circ \dots \text{(iii)}$$

Thus, on adding (ii) and (iii), if $\Lambda_{\text{Na}^+}^\circ$ and $\Lambda_{\text{Cl}^-}^\circ$ are subtracted we can obtain the value of $\Lambda_{\text{HOAc}}^\circ$.

Thus, additional value required is $\Lambda_{\text{NaCl}}^\circ$.

270 (c)

$$\Lambda_{\text{CH}_3\text{COOH}}^\circ = \Lambda_{\text{CH}_3\text{COONa}}^\circ + \Lambda_{\text{HCl}}^\circ - \Lambda_{\text{NaCl}}^\circ$$

271 (b)

$$\Lambda_m = \Lambda_{\text{eq}} \times \text{valency factor}$$

$$= 314.28 \times 2 = 628.56 \text{ mho cm}^2 \text{ mol}^{-1}$$

272 (a)

$$t_{\text{K}^+} + t_{\text{Cl}^-} = 1$$

$$\therefore t_{\text{K}^+} = 1 - 0.505 = 0.495$$

273 (a)

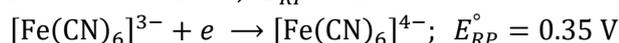
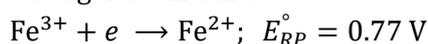
In voltaic cells net redox change brings in passage of current. In concentration cells transfer of matter brings in passage of current.

274 (c)

H^+ gets discharged at cathode and thus, $[\text{OH}^-]$ increases in solution.

275 (b)

Fe^{3+} and $\text{Fe}(\text{CN})_6^{3-}$ are oxidants. Higher is E_{RP}° stronger is oxidant.



276 (a)

$$\kappa = \frac{1}{R} \times \text{cell constant}$$

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

277 (c)

$$W = \frac{E \times i \times t}{96500}$$

$$\therefore 20 = \frac{36.5 \times i \times 360 \times 60}{96500}$$

$$= 2.45 \text{ ampere}$$

278 (b)

E_{cell}° for the reaction $M^+ + X^- \rightarrow M + X$ is given as follows

$$= E_{\text{cathode}}^\circ + E_{\text{anode}}^\circ$$

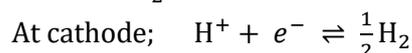
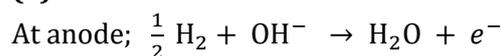
$$= 0.44 \text{ V} + 0.33 \text{ V}$$

$$= +0.11 \text{ V}$$

Since, E_{cell}° is positive for the above reaction.

Hence, this reaction is spontaneous.

279 (a)

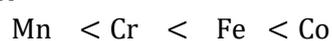


$$K = \frac{[\text{H}_2\text{O}]}{[\text{H}^+][\text{OH}^-]}$$

$$E^\circ = \frac{RT}{F} \ln K_w$$

280 (b)

$E_{\text{M}^{2+}/\text{M}}^\circ$ values follow the order with negative sign



$$E_{\text{RP}}^\circ \quad -1.18 \quad -0.74 \quad -0.44 \quad -0.27 \text{ V}$$

281 (c)

The number of ions present in solution as well as ionic mobility increase on dilution in case of weak electrolytes and thus, equivalent conductivity increases. Also in case of strong electrolytes only ionic mobility increases with dilution and thus, equivalent conductivity increases. For weak electrolytes, dilution causes increase in number of ions as well as increase in ionic mobility.

282 (b)

Specific conductance = conductance \times cell constant

$$1.3 \text{ Sm}^{-1} = \frac{1}{50} \text{ S} \times \text{cell constant}$$

\therefore Cell constant

$$= 1.3 \times 50 \text{ m}^{-1} = 65 \text{ m}^{-1} = (65/100) \text{ cm}^{-1}$$

$$\text{Molar conductivity} = \frac{1000 \times \text{conductance} \times \text{cell constant}}{\text{molarity}}$$

$$= \frac{1000}{0.4} \times \frac{1}{260} \times \frac{65}{100} = 6.25 \text{ Scm}^2 \text{ mol}^{-1}$$

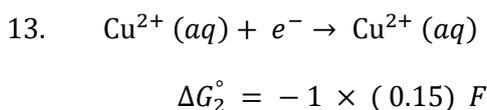
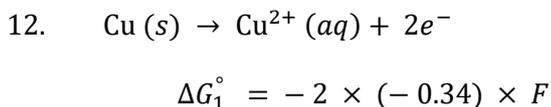
$$= 6.25 \times 10^{-4} \text{ Sm}^2 \text{ mol}^{-1}$$

283 (d)

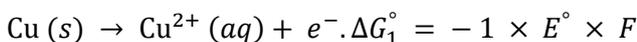
There will be no passage of current and ions will show simply diffusion.

284 (c)

From given data (from $\Delta G^\circ = -nE^\circ F$)



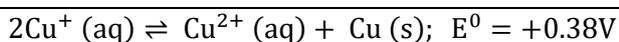
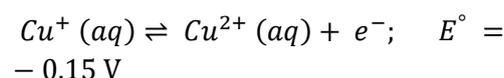
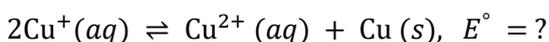
On addition,



$$\begin{aligned} \Delta G_3^{\circ} &= \Delta G_1^{\circ} + \Delta G_2^{\circ} \\ &= (-2 \times -0.34 \times F) + (-1 \times 0.15 \times F) \\ &= +0.68 F - 0.15 F = 0.53 F \end{aligned}$$

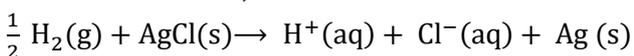
$$\text{Or } E^{\circ} = -0.53 \text{ V}$$

Reaction,



286 (d)

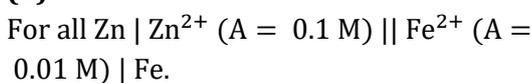
The cell reaction is,



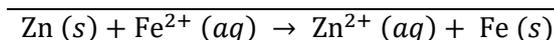
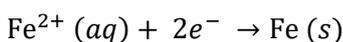
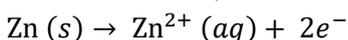
Obviously, here hydrogen is being oxidized and AgCl is getting reduced. Hence, the correct cell representation will be



287 (b)



The cell reaction,



On applying Nernst equation,

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log_{10} \frac{[\text{Zn}^{2+}]}{[\text{Fe}^{2+}]}$$

$$0.2905 = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log_{10} \frac{0.1}{0.01}$$

$$0.2905 = E_{\text{cell}}^{\circ} - 0.0295 \times \log_{10} 10$$

$$0.2905 = E_{\text{cell}}^{\circ} - 0.0295 \times 1$$

$$\therefore E_{\text{cell}}^{\circ} = 0.2905 + 0.0295 = 0.32 \text{ V}$$

At equilibrium ($E_{\text{cell}} = 0$)

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log_{10} K_c$$

$$\therefore 0 = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log_{10} K_c$$

$$\text{or } E_{\text{cell}}^{\circ} = \frac{0.0591}{2} \log_{10} K_c$$

$$0.32 = \frac{0.0591}{2} \log_{10} K_c$$

$$\text{or } K_c = 10^{0.32/0.0295}$$

288 (a)

$$\text{Cell constant} = \frac{\text{sp. conductance}}{\text{conductance}}$$

289 (a)

$$\text{Cell constant} = \frac{1}{a} = \frac{0.8}{1.2} = 0.66 \text{ cm}^{-1}.$$

290 (c)



$$\text{or } \frac{0.216}{108} = \frac{W_{\text{Hg}^{2+}}}{a/2} = \frac{W_{\text{Hg}_2^{2+}}}{a/1}; (\text{and } a = 200.6)$$

291 (a)

Among Cu, Al, Fe and Zn, Al occupies highest position in electrochemical series.

Al displaces Cu, Fe and Zn from their salts.

292 (a)

We know from Kohlrausch's law

$$\Lambda_{\text{CH}_3\text{COOH}}^{\circ} = \Lambda_{\text{CH}_3\text{COONa}}^{\circ} + \Lambda_{\text{HCl}}^{\circ} - \Lambda_{\text{NaCl}}^{\circ}$$

293 (d)



$$\text{Given, } E_{\text{Ni}^{2+}/\text{Ni}}^{\circ} = -0.25 \text{ V}$$

$$E_{\text{Au}^{3+}/\text{Au}}^{\circ} = +1.5 \text{ V}$$

Here, Ni is anode and Au is cathode.

$$\begin{aligned} \therefore E_{\text{cell}} &= E_C - E_A \\ &= 1.5 - (-0.25) \\ &= 1.5 + 0.25 \\ &= 1.75 \text{ V} \end{aligned}$$

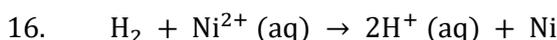
294 (c)

For the cell reaction, Fe acts as cathode and Sn as anode

Hence,

$$E_{\text{cell}} = 0.80 - (-0.76) = 1.56 \text{ V}$$

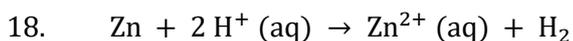
(maximum voltage)



$$E_{\text{cell}} = -0.25 \text{ V}$$



$$E_{\text{cell}} = 0.34 - (-0.76) = 1.00 \text{ V}$$



$$E_{\text{cell}} = 0 - (0.76 \text{ V}) = 0.76 \text{ V}$$

313 (c)

Cu voltameter or Cu or Ag coulometer are used to detect the amount deposited on an electrode during passage of known charge through solution.

314 (d)

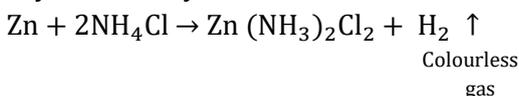
$$E_{\text{RP}_{\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}}} = 1.33 \text{ V and } E_{\text{RP}_{\text{Fe}^{3+}/\text{Fe}^{2+}}} = 0.77 \text{ V}$$

$\therefore E_{\text{OP}_{\text{Fe}^{2+}/\text{Fe}^{3+}}}$ is more thus it will oxidise or electron will flow from Fe electrode to Cr electrode. Also Fe electrode will be negative. Also.

$$E_{\text{cell}} = E_{\text{OP}_{\text{Fe}}} + E_{\text{RP}_{\text{Cr}}} = -0.77 + 1.33$$

315 (c)

H_2 gas is evolved, when Zn reacts with NH_4Cl in dry cell battery.



316 (d)

Reducing character is based upon higher negative value of reduction electrode potential.

Thus, order of reducing character is

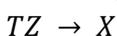
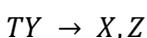
$$B > C > A$$

317 (a)

More is E_{OP} , more is the tendency to get itself oxidised and more is reducing power. Given values are of E_{OP} .

318 (a)

$$\text{Oxidizing tendency} \propto \frac{1}{\text{electrode potential}}$$



\Rightarrow Order of electrode potential is

$$\text{TY} < \text{TZ} < \text{TX}$$

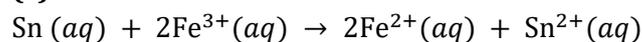
\Rightarrow Order of oxidation of the anion is

$$\text{Y}^- > \text{Z}^- > \text{X}^-$$

319 (a)

In the internal circuit of Agalvanic cell ions flow while in the external circuit, there is Aflow of electrons from zinc rod to copper rod.

321 (c)



$$E_{\text{cell}}^\circ = E_{\text{oxi}}^\circ + E_{\text{red}}^\circ$$

$$= E_{\text{Sn}/\text{Sn}^{2+}}^\circ + E_{\text{Fe}^{3+}/\text{Fe}^{2+}}^\circ$$

$$\text{Given, } E_{\text{Sn}^{2+}/\text{Sn}}^\circ = -0.14 \text{ V}$$

$$\therefore E_{\text{Sn}/\text{Sn}^{2+}}^\circ = +0.14 \text{ V}$$

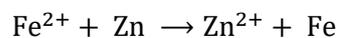
$$E_{\text{Fe}^{3+}/\text{Fe}^{2+}}^\circ = 0.77 \text{ V}$$

$$E_{\text{cell}}^\circ = 0.14 + 0.77 = 0.91 \text{ V}$$

322 (d)



Cell reaction is



$$E_{\text{cell}}^\circ = E_{\text{cathode}} - E_{\text{anode}}$$

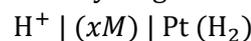
$$= -0.44 - (0.76)$$

$$= -0.44 + 0.76$$

$$= 0.32 \text{ V}$$

323 (c)

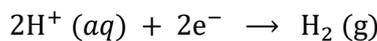
Reduction hydrogen half-cell is



+

Pressure p_{H_2}

Half - cell reaction is



Reaction quotient

$$= Q = P_{\text{H}_2} / [\text{H}^+]^2, n = 2$$

$$E_{\text{red}} = E_{\text{red}}^\circ - \frac{0.0591}{n} \log Q = 0 - \frac{0.0591}{2} \log Q$$

	P_{H_2}	$[\text{H}^+]$	Q	E_{red}
(a)	1atm	2.0 M	0.25	+ve
(b)	1atm	1.0 M	1.0	0
(c)	2atm	1.0 M	2.0	-ve
(d)	2atm	2.0 M	0.50	+ve

$E_{\text{red}} = 0.00 \text{ V}$ for standard hydrogen electrode

If $Q > 1$, then $E_{\text{red}} = -\text{ve}$.

Thus, correct answer is (c).

324 (a)

During the electrolysis of an aqueous solution of MgSO_4 between inert electrodes, the products available on the cathode and the anode are $\text{H}_2(\text{g})$ and $\text{O}_2(\text{g})$ respectively.

326 (d)

Follow Kohlrausch's law.

328 (a)

Because in it covalent bonding is present

329 (b)

Cations carrying positive charge move towards cathode where they get discharged.

330 (a)

The PH of 0.1 M HCl and 0.1 M acetic acid is not the same, because HCl is a strong acid, so its pH is more and CH₃COOH is a weak acid, so its pH is less

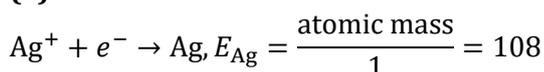
331 (d)

Electrochemical cell are based upon the reaction between various electrolytes. The reaction given in option (d) does not involve electrolytes, so it cannot be Abase for electrochemical cell.

333 (a)

The element having more E_{OP}° is oxidised and other is to be reduced.

334 (a)



$$\text{Number of faraday} = \frac{W_{\text{Ag}}}{E_{\text{Ag}}} = \frac{108}{108} = 1$$

335 (a)

$$[\text{H}^+] = c \cdot \alpha = 0.0133 \times 0.1 = 0.00133 \text{ M.}$$

336 (a)

$$E_1 = E^{\circ} - \frac{0.059}{1} \log[\text{H}^+]_1$$

$$E_2 = E^{\circ} - \frac{0.059}{1} \log[\text{H}^+]_2$$

on adding (also $E_{\text{H}}^{\circ} = 0$)

$$E_1 + E_2 = -\frac{0.059}{1} [\log(\text{H}^+)_1 + \log(\text{H}^+)_2]$$

Now for CH₃COOH = CH₃COO⁻ + H⁺

$$[\text{H}^+] = \frac{K_a [\text{CH}_3\text{COOH}]}{[\text{CH}_3\text{COO}^-]}$$

$$\therefore [\text{H}^+]_1 = K_a \cdot \frac{y}{x}$$

$$[\text{H}^+]_2 = K_a \cdot \frac{x}{y}$$

$$\therefore E_1 + E_2 = -\frac{0.059}{1} \left[\log \frac{K_a y}{x} + \log \frac{K_a x}{y} \right] = -0.059 [2 \log K_a]$$

$$\log K_a = \frac{E_1 + E_2}{2 \times (-0.059)}$$

$$\log K_a = \frac{E_1 + E_2}{0.118}$$

$$\text{or } \text{p}K_a = \frac{E_1 + E_2}{0.118}$$

337 (a)

$$W = \frac{E \times i \times t}{96500}$$

$$\therefore \frac{112 \times 2}{22400} = \frac{1 \times i \times 965}{96500}$$

$$\therefore i = 1 \text{ ampere}$$

338 (d)

$$\text{pH} = 3, [\text{H}^+] = 10^{-3}$$

$$E = E_{\text{red}}^{\circ} + 0.059 \log(\text{ion})$$

$$E = 0 + 0.059 \log(10^{-3})$$

$$E = +0.059(-3) = -0.177 \text{ V}$$

339 (d)

$$\text{Given, } E_{\text{cell}}^{\circ} = 0.34 \text{ V}$$

$$E_{\text{cell}}^{\circ} = E_{(\text{Cu}^{2+}/\text{Cu})}^{\circ} - E_{(\text{H}_2/\text{H}^+)}^{\circ}$$

$$0.34 = E_{(\text{Cu}^{2+}/\text{Cu})}^{\circ} - 0.00$$

$$\therefore E_{(\text{Cu}^{2+}/\text{Cu})}^{\circ} = +0.34$$

340 (d)

Reduction potential of hydrogen electrode,

$$E_{\text{H}} = -\frac{2.303RT}{F} \log \frac{1}{[\text{H}^+]}$$

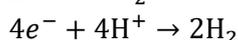
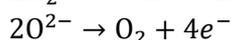
$$= -0.059 \text{ pH} = -0.059 \times 3 = -0.177 \text{ V}$$

341 (c)

Ag or Cu voltameters or coulometers are used to measure the total charge passed through solution in terms of weight of Ag or Cu deposited there on voltameter or coulometer cathode.

342 (b)

Reaction for electrolysis of water is



$$\therefore n = 4,$$

So, 4 F charge liberates = 1 mol = 22.4 dm³ oxygen

$$\therefore 1 \text{ F charge will liberate} = \frac{22.4}{4} = 5.6 \text{ dm}^3 \text{ oxygen}$$

343 (a)

$$Q = 2.5 \times 386 = 965 \text{ C}$$

$$2F(2 \times 96500 \text{ C}) \text{ deposited, Cu} = 63.5 \text{ g}$$

$$\text{Hence, } 965 \text{ C will deposit, Cu} = 0.3175 \text{ g}$$

344 (b)

Given, product = 0.1 M and reactant = 1 M

$$E_{\text{cell}}^{\circ} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{[\text{products}]}{[\text{reactants}]}$$

$$1.50 = E_{\text{cell}}^{\circ} - 0.02955 \log \left[\frac{0.1}{1} \right]$$

$$E_{\text{cell}}^{\circ} = 1.470 \text{ V}$$

$$E_{\text{cell}}^{\circ} = E_{\text{H}^+/\text{H}_2}^{\circ} - E_{\text{M}^{\circ}/\text{M}^{2+}}^{\circ}$$

$$E_{\text{M}^{\circ}/\text{M}^{2+}}^{\circ} = -1.470$$

$$\text{So, } E_{M^{2+}/M}^{\circ} = 1.470 \text{ V}$$

345 (c)

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} + \frac{0.059}{2} \log \frac{[\text{Cu}^{2+}]}{[\text{Zn}^{2+}]};$$

Thus, on doubling concentration of both Cu^{2+} and Zn^{2+} , there will be no effect on E_{cell} .

346 (a)

$$\Lambda_{\text{H}_2\text{O}}^{\infty} = \lambda_{\text{H}^+}^{\infty} + \lambda_{\text{OH}^-}^{\infty}$$

347 (a)

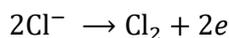
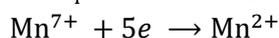
$$\kappa = \Lambda_{\text{eq}} \cdot C$$

$$= (91 \Omega^{-1} \text{ cm}^2 \text{ eq}^{-1}) \left(\frac{2.54}{159/2 \times 1000} \text{ eq} \cdot \text{cm}^{-3} \right)$$

$$= 2.9 \times 10^{-3} \Omega^{-1} \text{ cm}^{-1}$$

348 (a)

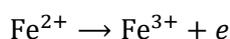
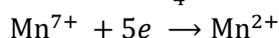
MnO_4^- will oxidise Cl^- ion according to equation



Thus, $E_{\text{cell}}^{\circ} = E_{\text{OP}_{\text{Cl}^-/\text{Cl}_2}}^{\circ} + E_{\text{RP}_{\text{MnO}_4^-/\text{Mn}^{2+}}}^{\circ}$

$$= -1.40 + 1.51 = 0.11 \text{ V or reaction is}$$

feasible MnO_4^- will oxidise Fe^{2+} to Fe^{3+}



$$\therefore E_{\text{cell}}^{\circ} = E_{\text{OP}_{\text{Fe}^{2+}/\text{Fe}^{3+}}}^{\circ} + E_{\text{MnO}_4^-/\text{Mn}^{2+}}^{\circ}$$

$$= -0.77 + 1.51$$

$$= 0.74 \text{ V or reaction is feasible}$$

Thus, MnO_4^- will not oxidise Fe^{2+} to Fe^{3+} in aqueous HCl medium but it will also oxidise Cl^- to Cl_2 . Suitable oxidant should not oxidise Cl^- to Cl_2 and should only oxidise Fe^{2+} to Fe^{3+} in redox titrations.

349 (c)

This is representation of standard hydrogen electrode.

350 (a)

$$E_{\text{cell}} = E_C - E_A$$

$$\text{Given, } E_{\text{Ag}^+/\text{Ag}}^{\circ} = 0.80 \text{ V}$$



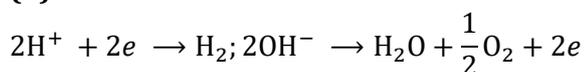
\therefore Hydrogen is anode and silver is cathode.

$$E_{\text{cell}} = E_C - E_A$$

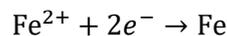
$$= 0.80 - 0 \quad (\because E_{\text{H}^+/\text{H}}^{\circ} = 0)$$

$$= 0.80 \text{ V}$$

351 (d)



352 (b)



$$E_{\text{Fe}} = \frac{56}{2} = 28$$

$$w_{\text{Fe}} = E_{\text{Fe}} \times \text{number of faraday}$$

$$= 28 \times 3 = 84 \text{ g}$$

353 (a)

$$w = Zit$$

$$\therefore w = \frac{32.69 \times 5 \times 60 \times 40}{96500} = 4.065 \text{ g}$$

354 (a)

Velocities of both K^+ and NO_3^- are nearly the same in KNO_3 , so it is used to make salt-bridge

355 (b)

It is secondary reference electrode.

356 (a)

$$E_{\text{cell}}^{\circ} = E_{\text{red (cathode)}}^{\circ} - E_{\text{oxi (anode)}}^{\circ}$$

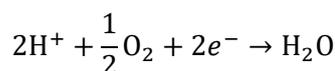
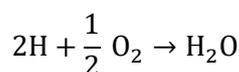
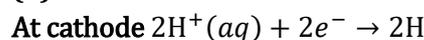
$$= E_{\text{Ag}^+/\text{Ag}}^{\circ} - E_{\text{Cu}^+/\text{Cu}^{2+}}^{\circ}$$

$$= 0.80 - (+0.34) = +0.46 \text{ V}$$

357 (c)

Because conductance increases when the dissociation is more

358 (a)



359 (b)

$$\text{Resistance of } \frac{N}{10} \text{ solution} = 2.5 \times 10^3 \Omega$$

$$\kappa = \frac{1}{\text{resistance}} \times \text{cell constant}$$

$$= \frac{1}{2.5 \times 10^3} \times 1.25$$

$$= \frac{1.25 \times 10^{-3}}{2.5} = 5 \times 10^{-4} \Omega^{-1} \text{ cm}^{-1}$$

Equivalent conductance

$$= \frac{\kappa \times 100}{M}$$

$$= \frac{5 \times 10^{-4} \times 1000}{1/10}$$

$$= 5 \Omega^{-1} \text{ cm}^2 \text{ equiv}^{-1}$$

360 (d)

In concentration cell, two electrolytic solutions of same electrolyte but having different concentrations (e.g., Pt H₂ | HCl | HCl | Pt H₂) are used. Moreover electrolyte of one strength but electrodes of two different concentrations are used.

361 (b)

$$E_{\text{cell}}^{\circ} = E_{\text{(RHS)}}^{\circ} - E_{\text{(LHS)}}^{\circ} = 1.61 - 1.51 = 0.10 \text{ V}$$

$$\Delta G^{\circ} = -nFE^{\circ}$$

$$= -5 \times 96500 \times 0.10 \text{ J}$$

$$\Delta G^{\circ} = -48.25 \text{ kJ}$$

362 (a)

$$W = \frac{E \times i \times t}{96500}$$

$$\text{or } 0.635 = \frac{63.5 \times i \times t}{2 \times 96500}$$

$$\therefore i \times t = 1930 \text{ C}$$

363 (c)

Given fact is : X is above Z and Y in electrochemical series and thus, Z cannot displace X from its salt.

364 (b)

The charging of lead storage battery involves the reverse reactions shown in answer 5.

365 (d)

$$\therefore R \propto l \text{ and } R \propto \frac{1}{a} \therefore R \propto \frac{l}{a} \text{ or } R = \rho \cdot \frac{l}{a}$$

Where R is resistance, ρ is specific resistance and l , a are length and area of cross-section of wire.

$$\therefore \kappa = \frac{1}{\rho} \text{ where } \kappa \text{ is conductivity}$$

$$\text{Then } \kappa = \frac{1}{R} \times \frac{l}{a} = C \times \frac{l}{a};$$

$$\text{Where } C \text{ is conductance} = \frac{1}{R}$$

Also, Equivalent conductivity = Conductivity $\times V_{\text{in mL}}$ where V is solution containing 1 equivalent in it or

$$\text{Eq. conductivity} = \kappa \times \frac{1000}{N}$$

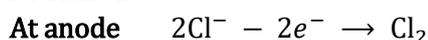
366 (b)

1 g atom of Mg = 2g eq. Mg = 2 faraday.

367 (b)



molten



When one mole of NaCl is electrolysed, 1 mole Na and $\frac{1}{2}$ mole Cl₂ is obtained.

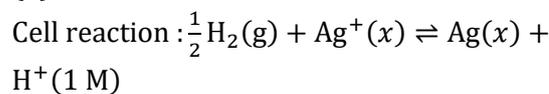
Thus, when 230 g ($\frac{230}{23} = 10\text{mol}$) Na is formed, the moles of

$$\text{Cl}_2 \text{ obtained will be} = \frac{\text{moles of Na}}{2}$$

$$= \frac{10}{2}$$

$$= 5.0 \text{ mol}$$

368 (a)



$$E = E^{\circ} = -\frac{0.0591}{n} \log \frac{[\text{H}^+]}{[\text{Ag}^+]}$$

$$0.503 = 0.80 - \frac{0.0591}{1} \log \frac{1}{x}$$

$$x = 9.43 \times 10^{-6} \text{ M}$$

Number of moles of Ag⁺ in 350 mL

$$= \frac{9.43 \times 10^{-6} \times 350}{1000} = 3.3 \times 10^{-6}$$

$$\text{Mass of Ag} = 3.3 \times 10^{-6} \times 108 = 3.56 \times 10^{-4} \text{ g}$$

$$\% \text{ of Ag in the ore} = \frac{3.56 \times 10^{-4}}{1.05} \times 100$$

$$= 0.0339\%$$

369 (a)

$$\alpha = \frac{\Lambda_v}{\Lambda_{\infty}} = \frac{15.8}{350} = 0.04514$$

370 (c)

$$E_{\text{cell}} = \frac{0.059}{1} \log \frac{[\text{H}^+]_{\text{RHE}}}{[\text{H}^+]_{\text{LHE}}}$$

$$= 0.059 \log \frac{10^{-3}}{10^{-8}} = 0.059 \times 5 = 0.295 \text{ V}$$

371 (a)

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.059}{n} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

$$= 1.10 - \frac{0.059}{2} \log \frac{0.1}{0.1}$$

$$= 1.10 \text{ V}$$

372 (c)

Ag is below Cu in electrochemical series. Metal placed above replaces other placed below in series.

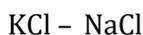
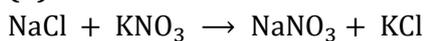
373 (b)

Eq. of Ni = Eq. of Cr

$$\therefore \frac{0.3}{59/2} = \frac{W}{52/3}$$

$$\therefore W_{\text{Cr}} = 0.176 \text{ g}$$

374 (b)



Molar conductivities 152 128

Difference = 24 S cm² mol⁻¹

Molar conductivities of KNO₃ = 111

Molar conductivities of NaNO₃ = 111 - 24

= 87 S cm² mol⁻¹

375 (d)

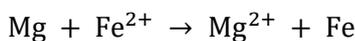
The degree of ionisation of an electrolyte increases with dilution and temperature as well as in presence of polar solvent.

376 (a)

$$\begin{aligned} E_{\text{cell}}^{\circ} &= E_C - E_A \\ &= -0.45 - (-2.37) \\ &= -1.92 \text{ V} \end{aligned}$$



The cell reaction is



$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.059}{n} \log \frac{[\text{Mg}^{2+}]}{[\text{Fe}^{2+}]}$$

$$1.92 = 1.92 - \frac{0.059}{n} \log \frac{x}{0.01}$$

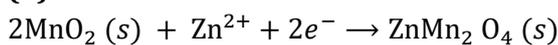
$$0 = \frac{-0.059}{2} \log \frac{x}{0.01}$$

$$\therefore x = 0.01 \ \text{M}$$

377 (d)

Follow theory of corrosion.

378 (b)



2 × 87 g MnO₂ required 2 × 96500 F charge

$$8 \text{ g MnO}_2 = \frac{2 \times 96500 \times 8}{2 \times 87}$$

$$= 8873.56 \ \text{F}$$

$$Q = it$$

$$8873.56 = 2 \times 10^{-3} \times t$$

$$t = \frac{8873.56 \times 10^3}{2}$$

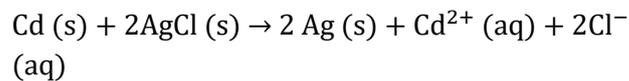
$$= 4436781.5 \ \text{s}$$

$$= \frac{4436781.5}{24 \times 60 \times 60}$$

$$= 51.35 \ \text{days}$$

379 (d)

The cell reaction is



$$E_1 = 0.6915 \ \text{V} \quad \text{at } 0^{\circ}\text{C}$$

$$E_2 = 0.6753 \ \text{V} \quad \text{at } 25^{\circ}\text{C}$$

$$\begin{aligned} \text{Now, } \frac{\partial E_{\text{cell}}}{\partial T} &= \frac{E_2 - E_1}{T_2 - T_1} \\ &= \frac{0.6753 - 0.6915}{298 - 273} \\ &= -6.48 \times 10^{-4} \end{aligned}$$

$$\Delta S = nF \left[\frac{\partial E_{\text{cell}}}{\partial T} \right]$$

Now, we put the value

$$\Delta S = 2 \times 96500 (-6.48 \times 10^{-4}) = -125.064$$

We know that,

$$\begin{aligned} \Delta G &= -nFE_{\text{cell}} \\ &= -2 \times 96500 \times 0.6753 \\ &= -1.303 \times 10^5 \end{aligned}$$

As, $\Delta G = \Delta H - T\Delta S$

For calculating $\Delta H = \Delta G + T\Delta S$

$$= -1.303 \times 10^5 + 298 (-125.064 \ \text{kJ})$$

$$\Delta H = -1.6726 \times 10^5 \ \text{J}$$

$$= -167.26 \ \text{kJ}$$

381 (c)

$$\begin{aligned} E_{\text{red}} &= 0.78 + \frac{0.0591}{1} \log(10^{-7})^2 \\ &= 0.78 - 0.059 \times 7 = 0.367 \ \text{V} \end{aligned}$$

382 (b)

$$\frac{\text{wt. of Al deposited}}{\text{wt. of Cu deposited}} = \frac{\text{eq. wt. of Al}}{\text{eq. wt. of Cu}}$$

$$\frac{2.7}{\text{wt. of Cu}} = \frac{27/3}{63.5/2}$$

$$\text{wt. of Cu} = 9.525 \ \text{g}$$

383 (d)

$E = Q \cdot V = 1\text{C} \times 1\text{V} = 1\text{J}$. It is unit of energy. Also $1\text{J} = 10^7 \ \text{erg}$.

384 (a)

$$E_{\text{Cu}/\text{Cu}^{2+}} = E_{\text{Cu}/\text{Cu}^{2+}}^{\circ} - \frac{0.059}{2} \log[\text{Cu}^{2+}]$$

if $\log[\text{Cu}^{2+}] = 0$, i.e., $[\text{Cu}^{2+}] = 1$, then $E_{\text{Cu}/\text{Cu}^{2+}} = E_{\text{Cu}/\text{Cu}^{2+}}^{\circ}$

$$\text{or } OA = E_{\text{Cu}/\text{Cu}^{2+}}^{\circ} = -E_{\text{Cu}/\text{Cu}^{2+}}^{\circ} = -0.34$$

$$\begin{aligned} \text{Now, } E_{\text{Cu}/\text{Cu}^{2+}} &= -0.34 - \frac{0.059}{2} \log 0.1 \\ &= -0.34 + \frac{0.059}{2} \text{ V} \end{aligned}$$

385 (a)

$$E_{\text{cell}}^{\circ} = E_{\text{OP}_L}^{\circ} + E_{\text{RP}_R}^{\circ} = -\phi_L + \phi_R$$

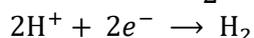
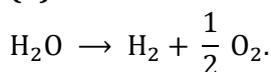
386 (b)

Pb is above Ag and below Ni in electrochemical series.

387 (b)

Smaller is the size of ion, more is hydration, lesser is mobility, i.e., mobility of ion in aqueous solution decreases with decreases with decrease in size of hydrated ion.

388 (d)



∴ 1 mole of H_2 is liberated from 2 moles of electrons.

$$\begin{aligned} 2 \text{ moles electrons} &= 2 \text{ F charge} \\ &= 2 \times 96500 \text{ C charge} \end{aligned}$$

Given, $i = 4 \text{ A}$, $t = 30 \text{ min}$

$$\therefore 2 \times 96500 \text{ C charge liberates} \\ = 22400 \text{ mL H}_2 \text{ at NTP}$$

$$\begin{aligned} \therefore 4 \times 30 \times 60 \text{ C charge liberates} \\ &= \frac{22400}{2 \times 96500} \times 4 \times 30 \times 60 \text{ mL H}_2 \\ &= 835.6 \text{ mL} \\ &= 0.836 \text{ L of H}_2 \end{aligned}$$

389 (d)

$$R = 100 \Omega$$

$$\kappa = \frac{1}{R} \left(\frac{l}{a} \right)$$

$$\frac{l}{a} \text{ (cell constant)} = 1.29 \times 100 \text{ m}^{-1}$$

Given, $R = 520 \Omega$; $C = 0.2 \text{ M}$

μ (molar conductivity) = ?

$$\mu = \kappa \times V$$

(κ can be calculated as $\kappa = \frac{1}{R} \left(\frac{l}{a} \right)$ now cell constant is known)

$$\begin{aligned} \text{Hence, } \mu &= \frac{1}{520} \times 129 \times \frac{1000}{0.2} \times 10^{-6} \text{ m}^3 \\ &= 12.4 \times 10^{-4} \text{ S m}^2 \text{ mol}^{-1} \end{aligned}$$

390 (a)

$$u^0 = \frac{\lambda^0}{96500} = \frac{5 \times 10^{-3}}{96500} = 5.18 \times 10^{-8}$$

391 (b)

$$Q = 10^{-6} \times 1 \text{ C};$$

Also, $96500 \text{ C} = 6.023 \times 10^{23} \text{ electrons}$.

392 (c)

According to Faraday Ist law

$$w = Zit$$

Where, w = amount of substance

Z = Faraday constant or electrochemical equivalent

i = current in ampere

t = time in second

If $i = 1$, $t = 1$, $w = Z$

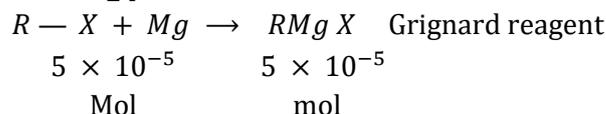
393 (c)

96500 C current produces 12 g Mg

9.65 C current produces

$$= \frac{12 \times 9.65}{96500} = 1.2 \times 10^{-3} \text{ g Mg}$$

$$= \frac{1.2 \times 10^{-3}}{24} = 5 \times 10^{-5} \text{ mol Mg}$$



Hence, number of moles of $RMgX$ produced = 5×10^{-5} .

394 (a)

$$\Delta G = \Delta H - T\Delta S, \Delta G = -nEF \text{ and}$$

$$\Delta G = \Delta H + T \left[\frac{\partial(\Delta G)}{\partial T} \right]_P$$

$$\therefore \Delta H = -nF \left[E - T \left(\frac{\partial E}{\partial T} \right)_P \right]$$

$$\therefore \Delta S = nF \left(\frac{\partial E}{\partial T} \right)_P$$

395 (a)

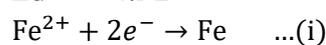
Equal equivalent of each are liberated.

Eq. of X = Eq. of Y

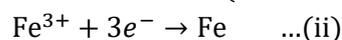
$$\frac{W_1}{2M/2} = \frac{W_2}{M/1} \quad \therefore W_1 = W_2$$

396 (a)

$$\Delta G^\circ = -nFE^\circ$$

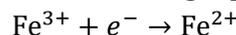


$$\Delta G^\circ = -2 \times F \times (-0.440 \text{ V}) = 0.880 \text{ F}$$



$$\Delta G^\circ = -3 \times F \times (-0.036) = 0.108 \text{ F}$$

On subtracting Eqs. (i) from (ii), we get



$$\Delta G^\circ = 0.108 \text{ F} - 0.880 \text{ F} = -0.772 \text{ F}$$

$$\begin{aligned} E^\circ \text{ for the reaction} &= -\frac{\Delta G^\circ}{nF} = -\frac{(-0.772 \text{ F})}{1 \times F} \\ &= +0.772 \text{ V} \end{aligned}$$

397 (c)

$$\begin{aligned} E_{\text{cell}}^\circ &= E_{\text{OP}_{\text{Zn}}}^\circ + E_{\text{RP}_{\text{Cu}}}^\circ \\ &= 0.76 + 0.34 = +1.10 \text{ V}. \end{aligned}$$

398 (d)

$$\lambda_c^\infty = \mu_c^\infty \times \text{Faraday};$$

$$\lambda_a^\infty = \mu_a^\infty \times \text{Faraday};$$

$$\text{Also, } \lambda_{\text{electrolyte}}^\infty = \lambda_a^\infty \times \lambda_c^\infty.$$

399 (a)

Metal undergoes oxidation during corrosion and thus, act as anode.

400 (c)

MnO_4^- is itself a strong oxidant.

401 (c)

Eq. of Al = Eq. of Cu = Eq. of Na

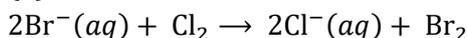
$$\text{or } \frac{1}{3} \text{ mole Al} = \frac{1}{2} \text{ mole Cu} = 1 \text{ mole Na}$$

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

402 (c)

In Agalvanic cell, oxidation (*i. e.*, removal of e^-) occurs at anode. These electrons flow through external circuit from anode to cathode. Therefore, the direction of current in external circuit is from cathode (-ve) to anode (+ve).

403 (a)



Bromine is released by chlorine as more reactive halogen displaces less reactive halogen from their salt solutions.

404 (a)

E_{RP}° for H is maximum in these.

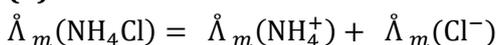
405 (d)

Eq. of $\text{Fe}^{2+} = \text{Eq. of Fe}^{3+}$

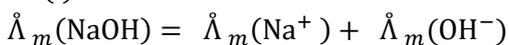
$$\text{or } \frac{W_1}{A/2} = \frac{W_2}{A/3}$$

$$\text{or } \frac{W_{\text{Fe}^{2+}}}{W_{\text{Fe}^{3+}}} = \frac{3}{2}$$

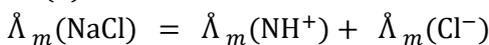
406 (b)



.....(i)

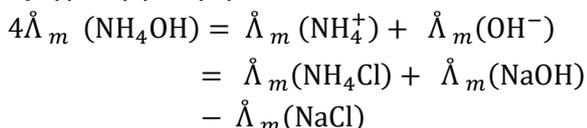


.....(ii)

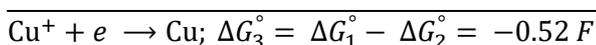


.....(iii)

By (i) + (ii) + (iii)



407 (b)



$$\therefore -1 \times E^\circ \times F = 0.52 F \Rightarrow E^\circ = 0.52 V$$

408 (c)

Only oxidation occurs at anode. Also discharge potential of H_2O is less than discharge potential of SO_4^{2-} .

409 (c)

Calculate E_{cell}° for each. For (C) $E_{\text{cell}}^\circ = 3.6 V$.

410 (c)

$$6 \times 10^{23} \text{ electron} = 1 \text{ eq.}$$

411 (d)

$E_{\text{cell}}^\circ = E_{OP_I}^\circ + E_{RP_{II}}^\circ$; I is oxidised, II is reduced.

424 (d)

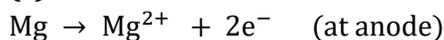
412 (a)

Charge = 2 × charge of electron = 2 × 1.602 × 10⁻¹⁹ C.

414 (a)

The substances which have lower reduction potentials are stronger reducing agent. The reduction potential of zinc is lowest among these hence, it is the strongest reducing agent.

416 (c)



$$E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ$$

$$= +0.34 - (-2.37)$$

$$= +2.71 V$$

417 (a)

As ammonia is added the concentration of H^+ decreases, thus oxidation potential increases by 0.65

418 (d)

In presence of attacked electrodes (Ni electrodes with Ni salt), the metal gets dissolved at anode and metal ions get discharged at cathode.

419 (a)

A cation having highest reduction potential will be reduced first and so on. However, Mg^{2+} in aqueous solution will not be reduced

($E_{\text{Mg}^{2+}/\text{Mg}}^\circ < E_{\text{H}_2\text{O}/\frac{1}{2}\text{H}_2 + \text{OH}^-}$). Instead water would be reduced in preference

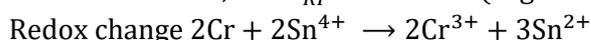
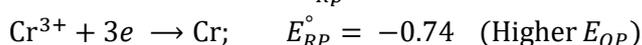
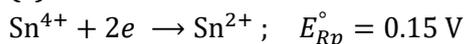
420 (d)

Mercury cannot displace hydrogen from acid. The reason can be explained on the basis of electrochemical series. The metal, which is placed above hydrogen in this series, can displace hydrogen from an acid. Mercury is placed below hydrogen in ECS, hence, it cannot displace hydrogen from an acid.

421 (a)

This is Kohlrausch law for A_xB_y .

422 (c)

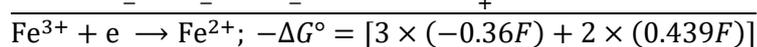
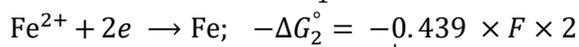
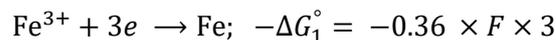


$$\therefore E_{\text{cell}}^\circ = E_{OP_{\text{Cr}}}^\circ + E_{RP_{\text{Sn}}}^\circ = 0.74 + 0.15 = +0.89 V$$

423 (c)

E_{cell}° for $\text{Zn}|\text{Zn}^{2+}||\text{Ni}^{2+}|\text{Ni}$ is positive

$$E^\circ = E_{OP_{\text{Zn}}}^\circ + E_{RP_{\text{Ni}}}^\circ = 0.76 - 0.23 = 0.53 V$$



$$\therefore +1 \times F \times E^\circ = [3 \times (-0.36F) + 2 \times (0.439F)]$$

$$\text{or } E^\circ = [3 \times (-0.36) + 2 \times 0.439]V$$

425 (d)

$$E = E^\circ - \frac{0.059}{n} \log \left[\frac{M}{M^{n+}} \right]$$

$$= E^\circ - \frac{0.059}{n} \log \left[\frac{\text{products}}{\text{reactants}} \right]$$

E = 0 at equilibrium

$$0 = E^\circ - \frac{0.059}{n} \log K_{eq}$$

$$\text{Or } \log K_{eq} = \frac{nE^\circ}{0.059}$$

$$\text{Or } \log K_{eq} = \frac{1 \times 0.36}{0.059} = 6.09$$

$$K_{eq} = 1.2 \times 10^6$$

426 (c)

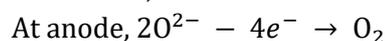
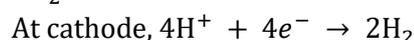
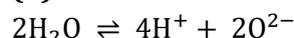
Electrolytes on dissolution in water furnish ions and these ions carry charge towards opposite electrodes.

427 (a)

$$W_{\text{Na}} = \frac{E \times i \times t}{96500} = \frac{23 \times 5 \times 10 \times 69}{96500}$$

$$= 0.715 \text{ g}$$

429 (b)



\therefore 4 Faraday of charge liberates = 1 mol

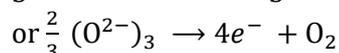
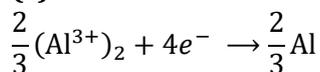
$$= 22.4 \text{ dm}^3 \text{O}_2$$

\therefore 1 Faraday of charge liberates

$$= \frac{22.4}{4} \times 1$$

$$= 5.6 \text{ dm}^3 \text{O}_2$$

430 (c)



$-\Delta G = nFE$ (E is potential required)

$$E = -\frac{960 \times 1000}{4 \times 96500} = -2.5 \text{ V}$$

431 (c)

Cl is placed below Br in electrochemical series; the non-metal placed below in series, replaces other from its solution.

432 (a)

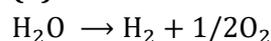
In a galvanic cell, the electrons flow from anode to cathode through the external circuit. At anode

(-ve pole) oxidation and at cathode (+ pole) reduction takes place

433 (b)

H^+ is lightest ion and thus, possesses maximum velocity among monovalent ions.

434 (b)



\therefore Mole ratio of $\text{H}_2\text{O}:\text{O}_2 :: 1:1/2$

435 (c)

$E_{OP}^\circ = +\text{Ve}$ thus, $E_{RP}^\circ = -\text{ve}$.

436 (c)

$$\text{Eq. of Al formed } \frac{W}{E} = \frac{it}{96500}$$

$$\therefore W_{\text{Al}} = \frac{4 \times 10^4 \times 6 \times 60 \times 60 \times 27}{96500} \left(E = \frac{27}{3} \right)$$

$$\therefore W_{\text{Al}} = 8.05 \times 10^4 \text{ g}$$

437 (c)

H_2 undergoes oxidation and $\text{AgCl}(\text{Ag}^+)$ undergoes reduction

438 (c)

The ions which lies below H^+ in electrochemical series, displace H^+ from solution.

$\therefore \text{Cu}^{2+}$ lies below H^+ in electrochemical series.

$\therefore \text{Cu}^{2+}$ ions displace H^+ when H_2 gas is bubbled in solution containing these ions.

439 (b)

Anode at which oxidation occurs is represented at left hand side.

440 (a)

Ionic conductance is ionic mobility \times Faraday.

441 (d)

$$1F = N \times e = E \text{ g} = 96500 \text{ C}$$

442 (b)

Standard electrode potential is measured by voltmeter.

443 (a)

$\text{Cl}_2 + 2\text{KI} \rightarrow 2\text{KCl} + \text{I}_2$; iodine will be liberated first to impart violet colour to CHCl_3 layer.

444 (b)

$$w_{\text{metal}} = \frac{E \times i \times t}{96500} = \frac{E \times 3 \times 50 \times 60}{96500}$$

$$\therefore E = \frac{96500 \times w}{3 \times 50 \times 60} = \frac{96500 \times 1.8}{3 \times 50 \times 60} = 19.3$$

445 (a)

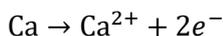
Given, current (i) = 0.25 mA = 0.025 A

Time (t) = 60 s

$$Q = i \times t = 60 \times 0.025 = 1.5 \text{ C}$$

$$\text{No. of electrons} = \frac{1.5 \times 6.023 \times 10^{23}}{96500}$$

$$e^- = 9.36 \times 10^{18}$$



$2e^-$ are required to deposit one Ca atom

$9.36 \times 10^{18} e^-$ will be used to deposit

$$= \frac{9.36 \times 10^{18}}{2} = 4.68 \times 10^{18}$$

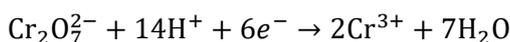
446 (b)

If F is Faraday and N is Avogadro number, charge of

$$\text{electron} = \frac{F}{N}$$

447 (b)

Half-cell reaction is

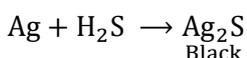
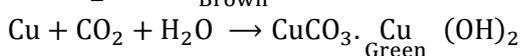
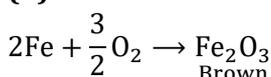


$$E_{\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}} - E_{\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}}^\circ$$

$$= -\frac{0.0591}{n} \log \frac{[\text{Cr}^{3+}]^2}{[\text{Cr}_2\text{O}_7^{2-}][\text{H}^+]^{14}}$$

$$= \frac{0.0591}{6} \log [10^{-3}]^{14} = -0.414 \text{ V}$$

448 (d)



450 (c)



On adding



$$\therefore 2E_3^\circ = 0.65 \text{ V} \quad \text{or} \quad E_3^\circ = 0.325 \text{ V}$$

451 (c)

When the solution of a weak electrolyte is diluted, the volume of the solution increases, hence equivalent conductivity (λ_c) increases. However, during this process, the number of current carrying particles per cm^2 decreases, hence specific conductivity (K_c) decreases.

452 (b)

$$E_{\text{cell}} = E_{\text{OPH}}^\circ + E_{\text{RPH}}^\circ + \frac{0.059}{2} \log \frac{[\text{H}^+]_{\text{RHE}P_1}^2}{[\text{H}^+]_{\text{LHE}P_2}^2}$$

$$= \frac{0.059}{2} \log \frac{(10^{-3})^2}{(10^{-2})^2} \quad (E_{\text{OP}}^\circ = E_{\text{RP}}^\circ)$$

$$= -0.059 \text{ V}$$

Thus, cell reaction is non-spontaneous.

453 (c)

From Kohlrausch's law

$$\Lambda_m^\infty = \nu_+ \lambda_+^\infty + \nu_- \lambda_-^\infty$$

For CaCl_2

$$\Lambda_m^\infty(\text{CaCl}_2) = \Lambda_{\text{Ca}^{2+}}^\infty + 2\Lambda_{\text{Cl}^-}^\infty$$

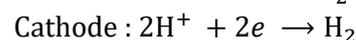
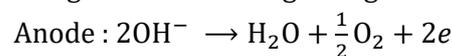
$$= 118.88 \times 10^{-4} + 2 \times 77.33 \times 10^{-4}$$

$$= 118.88 \times 10^{-4} + 154.66 \times 10^{-4}$$

$$= 273.54 \times 10^{-4} \text{ m}^2 \text{ mho mol}^{-1}$$

455 (b)

In case of very dilute solution of NaCl, electrolysis brings in the following changes :



456 (b)

Mg lie above Cu in electrochemical series and hence, Cu electrode acts as cathode

$$E_{\text{cell}}^\circ = E_{\text{Cu}^{2+}/\text{Cu}}^\circ = E_{\text{Mg}^{2+}/\text{Mg}}^\circ$$

$$\therefore 2.70 \text{ V} = 0.34 - E_{\text{Mg}^{2+}/\text{Mg}}^\circ$$

$$\therefore E_{\text{Mg}^{2+}/\text{Mg}}^\circ = -2.36 \text{ V}$$

457 (a)

This, is Hückel-Onsager equation.

458 (c)

Number of equivalents of silver formed

= number of equivalents of copper formed

In AgNO_3 , Ag is in +1 oxidation state

In CuSO_4 , Cu is in +2 oxidation state

$$\text{Equivalent weight of Ag} = \frac{108}{1} = 108$$

$$\text{Equivalent weight of Cu} = \frac{63.6}{2} = 31.8$$

$$\frac{w_1}{w_2} = \frac{E_1}{E_2}$$

$$\therefore \frac{10.79}{w_{\text{Cu}}} = \frac{108}{31.8}$$

$$\text{Or } w_{\text{Cu}} = \frac{10.79 \times 31.8}{108} = 3.2 \text{ g}$$

459 (b)

Energy = charge \times potential

$$= 1 \times 100 \times 115 \text{ J}$$

$$= 11.5 \text{ kJ}$$

460 (d)

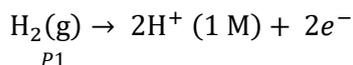
Cu is below Fe in electrochemical series.

461 (c)

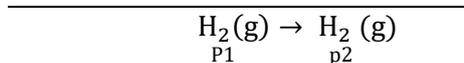
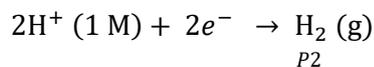
The process of zinc-plating on iron-sheet is known as galvanization.

462 (b)

LHS half cell



RHS half cell



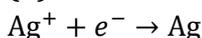
$$E_{\text{cell}}^{\circ} = 0.00 \text{ V}, K = \frac{p_2}{p_1}, n = 2$$

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{RT}{nF} \log_e K$$

$$= 0 - \frac{RT}{2F} \log_e \frac{p_2}{p_1}$$

$$E_{\text{cell}} = \frac{RT}{2F} \log_e \frac{p_1}{p_2}$$

463 (b)



96500 C charge liberates silver = 108 g

∴ 96500 C will liberate silver = 10.8 g

464 (b)

Given 125 mL of 1 M AgNO₃ solution. It means that

∴ 1000 mL of AgNO₃ solution contains

$$= 108 \text{ g Ag}$$

∴ 125 mL of AgNO₃ solution contains

$$= \frac{108 \times 125}{1000} \text{ g Ag}$$

$$= 13.5 \text{ g Ag}$$

∴ 108 g of Ag is deposited by 96500 C

∴ 13.5 g of Ag is deposited by

$$= \frac{96500}{108} \times 13.5$$

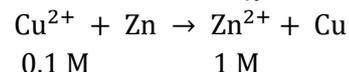
$$= 12062.5 \text{ C}$$

$$Q = it$$

$$\text{Or } t = \frac{Q}{i} = \frac{12062.5}{241.25} = 50$$

465 (c)

$$E_{\text{cell}}^{\circ} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log Q$$



$$Q = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} = \frac{1}{0.1} = 10$$

$$E_{\text{cell}} = 1.10 - \frac{0.0591}{2} \log 10$$

$$= 1.10 - 0.0295$$

$$= 1.0705 \text{ V}$$

466 (a)

$$(126 \text{ Scm}^2 \text{ mol}^{-1}) \Lambda_{\text{NaCl}}^{\infty} = \Lambda_{\text{Na}^+}^{\infty} + \Lambda_{\text{Cl}^-}^{\infty} \quad \dots(\text{i})$$

$$(152 \text{ Scm}^2 \text{ mol}^{-1}) \Lambda_{\text{KBr}}^{\infty} = \Lambda_{\text{K}^+}^{\infty} + \Lambda_{\text{Br}^-}^{\infty} \quad \dots(\text{ii})$$

$$(150 \text{ Scm}^2 \text{ mol}^{-1}) \Lambda_{\text{KCl}}^{\infty} = \Lambda_{\text{K}^+}^{\infty} + \Lambda_{\text{Cl}^-}^{\infty} \quad \dots(\text{iii})$$

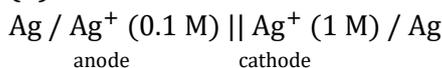
By Eqs. (i) + (ii) - (iii)

$$\therefore \Lambda_{\text{NaBr}}^{\infty} = \Lambda_{\text{Na}^+}^{\infty} + \Lambda_{\text{Br}^-}^{\infty}$$

$$= 126 + 152 - 150$$

$$= 128 \text{ Scm}^2 \text{ mol}^{-1}$$

467 (b)



$$E_{\text{cell}} = \frac{2.303RT}{nF} \log \frac{C_1}{C_2}$$

Here, n = number of electrons in cell reaction = 1

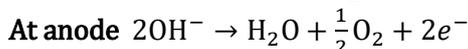
C_1 = concentration of cathodic electrolyte = 1 M

C_2 = concentration of anodic electrolyte = 0.1 M

$$\therefore E_{\text{cell}} = \frac{0.059}{1} \log \frac{1}{0.1}$$

$$E_{\text{cell}} = 0.059 \text{ V}$$

468 (a)



469 (c)

$$E_{\text{cell}}^{\circ} = E_{\text{OP}_{\text{Sn}}}^{\circ} + E_{\text{RP}_{\text{Fe}}}^{\circ} = 0.14 + 0.77 = 0.91 \text{ V}$$

470 (a)

An important application of fuel cell used in space programmes to provide power for heat and light as well as drinking water to astronauts.

471 (a)

$$\text{Conductance} = \frac{1}{R} = \frac{1}{210}$$

$$= 4.76 \times 10^{-3} \text{ mho.}$$

472 (c)

According to Faraday law's

$$m = Z \times Q = \frac{E \times Q}{96500}$$

$$m = \frac{108}{96500} \times 2 \times 96500 = 216 \text{ g}$$

473 (a)

Kohlraush gave the standard value of conductivity for different concentration of KCl solution. Thus, since,

$$\kappa = C \times \frac{l}{a};$$

By finding conductance of same concentration KCl solution, one finds l/a .

474 (c)

$$E_{\text{Cu}^{2+}/\text{Cu}}^{\circ} = +0.34 \text{ V} = E_{\text{RP}}^{\circ}$$

(Above H if arranged in decreasing E_{RP}°)

475 (c)

Given, weight of hydrogen liberated

$$= 5.04 \times 10^{-2} \text{ g}$$

Eq. wt. of hydrogen = 1.008

Eq. wt. of silver = 108

Weight of silver deposited, $w = ?$

According to Faraday's second law of electrolysis,

$$\frac{\text{weight of silver deposited}}{\text{weight of hydrogen liberated}} = \frac{\text{eq.wt. of silver}}{\text{eq.wt. of hydrogen}}$$

$$\frac{w}{5.04 \times 10^{-2}} = \frac{108}{1.008}$$

$$w = \frac{108 \times 5.04 \times 10^{-2}}{1.008} = 5.4 \text{ g}$$

476 (d)

At infinite dilution an electrolyte is 100% ionised or $\alpha = 1$ for weak electrolytes. At this point all interionic effect disappears.

477 (b)

$$\lambda_{\text{K}^+}^{\infty} = \mu_{\text{K}^+}^{\infty} \times F$$

$$\text{or } \mu_{\text{K}^+}^{\infty} = \frac{64.35}{96500}$$

$$= 6.67 \times 10^{-4} \text{ cm}^2 \text{ sec}^{-1} \text{ volt}^{-1}$$

478 (a)

Zn has more +ve E_{OP}° and thus possesses more tendency to get oxidised and act as reducing agent.

479 (a)

In Daniell cell, oxidation occurs at anode and reduction occurs at cathode. Thus, the element, which has higher negative value of reduction potential, is used as anode and that with lower negative value of reduction potential is used as cathode.

\therefore Reduction potential of Zn is more negative than Cu.

\therefore Zn | Zn²⁺ is anode and Cu²⁺ | Cu is cathode.

480 (d)

To prevent rusting or oxidation of Fe by the action of O₂ in presence of H⁺, galvanization of iron object is made.

481 (b)

It is definition of equivalent conductivity.

482 (a)

$$\Delta G^{\circ} = -nE^{\circ}F \quad (\text{for } \text{Cu} + 2\text{Ag}^+ \rightarrow \text{Cu}^{2+} + 2\text{Ag}) \quad n = 2$$

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

$$= -88788 \text{ J} = -89.0 \text{ kJ}$$

483 (d)

It does not depend upon mass.

484 (c)

The number of ions per cc decreases with dilution and therefore, specific conductance decreases with dilution.

485 (d)

Aluminium is more electropositive than Zn, hence Al replace it from its salt solution.

486 (c)

Blocks of magnesium metal provide cathodic protection and protect oxidation of steel.

487 (a)

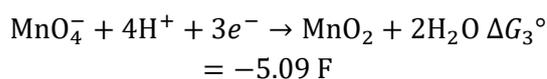


$$\Delta G_1^{\circ} = -5(1.51)F = -7.55F$$



$$\Delta G_2^{\circ} = -2(1.23)F = -2.46F$$

On substrating



$$E_{\text{MnO}_4^- | \text{MnO}_2}^{\circ} = \frac{\Delta G_3^{\circ}}{-nF} = \frac{-5.09F}{-3F} = 1.70 \text{ V}$$

488 (c)

Salt bridge is used to remove or eliminate liquid junction potential arising due to different relative speed of ions of electrolytes at the junction of two electrolytes in an electrochemical cell. Thus, a salt bridge such as KCl is placed in between two electrolytes. A salt used for this purpose should have almost same speeds of its cation and anion.

489 (d)

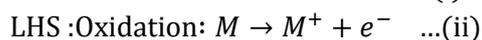
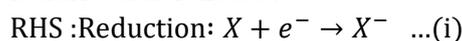
As AgNO₃ is added to solution, KCl will be displaced according to following reaction

$$\text{AgNO}_3(aq) + \text{KCl}(aq) \rightarrow \text{AgCl}(s) + \text{KNO}_3(aq)$$

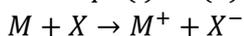
For every mole of KCl displaced from solution, one mole of KNO₃ comes in solution resulting in almost constant conductivity. As the end point is reached, added AgNO₃ remain in solution increasing ionic concentration, hence conductivity increases.

490 (c)

For the given cell, $M|M^+||X^-|X$, the cell reaction is derived as follows :



Add Eqs. (i) and (ii)



The cell potential = -0.11 V

Since, $E_{\text{cell}} = -ve$, the cell reaction derived above is non-spontaneous. In fact, the reverse reaction will occur spontaneously

491 (d)

Equivalent conductance (Λ) = specific conductance (κ) \times ϕ

where, ϕ = volume in cm^3 containing 1 g equivalent of electrolyte

1. g equivalent is dissolved in = 1000 cm^3

$$1 \text{ g equivalent is dissolved in} = \frac{1000}{0.1} \text{ cm}^3$$

$$= 10000 \text{ cm}^3$$

$$\text{So, } \Lambda = 10000 \kappa$$

492 (d)

All are same.

493 (c)

1 g atom of Al = 3 eq. of Al = 3 faraday charge
3 mole electrons = $3N$ electron.

494 (d)

E_{RP}° Cu^{2+}/Cu is more;

Thus, Cu^{2+} gets reduced easily to oxidise others.

495 (c)

In electrolytic cell, flow of electron is possible from cathode to anode through internal supply.

496 (a)

Eq. of Mg = Eq. of Al

$$\text{or } \frac{W_{\text{Mg}}}{12} = \frac{W_{\text{Al}}}{9}$$

497 (c)

For the change, $2\text{Fe}^{3+} + 3\text{I}^- \rightleftharpoons 2\text{Fe}^{2+} + \text{I}_3^-$

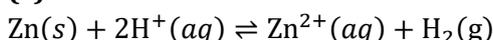
$$E_{\text{cell}}^{\circ} = 0.77 - 0.54 = 0.23 \text{ V}$$

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

$$0.23 = \frac{0.0591}{2} \log K_c$$

$$K_c = 6.26 \times 10^7$$

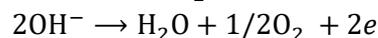
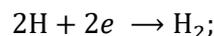
498 (c)



$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.059}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{H}^+]^2}$$

When H_2SO_4 is added then $[\text{H}^+]$ will increase therefore E_{cell} will also increase and equilibrium will shift towards right

499 (a)



\therefore Wt. ratio of $\text{H}_2 : \text{O}_2 = 2 : 16$ or $1 : 8$

500 (d)

Ionic mobilities increase on dilution.

501 (b)

Cell is completely discharged, it means equilibrium gets established,

$$E_{\text{cell}} = 0$$



Cell reaction: $\text{Zn} + \text{Cu}^{2+} \rightleftharpoons \text{Zn}^{2+} + \text{Cu}$

$$K_{\text{eq}} = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

We know,

$$E_{\text{cell}}^{\circ} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log K_{\text{eq}}$$

$$E_{\text{cell}}^{\circ} = \frac{0.0591}{n} \log K_{\text{eq}}$$

$$\text{Or } 1.10 = \frac{0.0591}{2} \log K_{\text{eq}}$$

$$K_{\text{eq}} = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} = \text{antilog } \frac{2.20}{0.0591}$$

$$= \text{antilog } 37.3$$

502 (c)

$$E_{\text{cell}} = E_{\text{OP}} + E_{\text{RP}}$$

503 (a)

$$E_{\text{cell}}^{\circ} = E_{\text{OP}_{\text{Zn}}}^{\circ} + E_{\text{RP}_{\text{Ag}}}^{\circ} = 0.76 + 0.77 = +1.53 \text{ V.}$$

504 (c)

Rust is $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$ or mixture of $\text{Fe}_2\text{O}_3 + \text{Fe}(\text{OH})_3$.

505 (a)

Metal with $-ve$ E_{RP}° or $+ve$ E_{OP}° possesses the tendency to get itself oxidised.

506 (a)

$$\begin{aligned} \Lambda &= k \times \frac{1000}{M} = \frac{1}{R} \times \frac{1}{a} \times \frac{1000}{M} \\ &= \frac{1}{45} \times \frac{2.2}{3.8} \times \frac{1000}{0.5} = 25.73 \end{aligned}$$

507 (a)

$$\text{Ionic mobility} = \frac{\text{speed of ions}}{\text{potential gradient}} = \frac{\text{m sec}^{-1}}{\text{volt m}^{-1}}$$

509 (a)

$$\text{Charge} = 10 \times 1 = 10 \text{ C}$$

Also, $96500 \text{ C} = 6 \times 10^{23}$ electrons.

510 (c)

Cryolite is used to lower the m. p. of alumina as well as to make it good conductor of current.

511 (c)

The degree of dissociation of all electrolyte increases with increase in dilution (or decrease in concentration).

512 (d)

Complex formation enhances the electrolytic deposition of Au.

513 (b)

Pure water is almost unionised so, it does not conduct electricity.

514 (a)

Galvanic cell is an electrochemical cell that converts the chemical energy of a spontaneous reaction into electrical energy.

515 (b)

Standard hydrogen electrode (SHE) is reference electrode. It is used to determine the electrode potential of any half cell. The electrode potential of any standard hydrogen electrode is arbitrarily taken as zero.

516 (a)

$$\Lambda_{\text{AgCl}}^{\infty} = \lambda_{\text{Ag}^+}^{\infty} + \lambda_{\text{Cl}^-}^{\infty} \\ = \lambda_{\text{AgNO}_3}^{\infty} + \lambda_{\text{NaCl}}^{\infty} - \lambda_{\text{NaNO}_3}^{\infty}$$

517 (d)

These are characteristics of fuel cells.

518 (a)

$$Q = N \cdot e \text{ or } e = \frac{Q}{N} = \frac{\text{faraday}}{\text{Av. no.}}$$

519 (a)

For strong electrolytes, $\alpha = 1$ at normal dilution, but $\Lambda_v \neq \Lambda_{\infty}$.

520 (b)

$E_{\text{cell}} = +ve$ thus redox changes are
 $A \rightarrow A^+ + e$ LHS shows oxidation.
 $B^+ + e \rightarrow B$ RHS shows reduction.
 $A + B^+ \rightarrow A^+ + B$

521 (b)

The laws of electrolysis were given by Faraday.

522 (a)

Current $i = ?$

$$\text{Equivalent weight of Al} = \frac{27}{3} = 9$$

$$w = Z \cdot i \cdot t \text{ or } i = \frac{w}{Z \cdot t} = \frac{0.09 \times 96500}{9 \times 96.5} \\ = 10 \text{ A}$$

523 (a)

Half cell reaction occurs only when it is coupled with other electrode.

524 (a)

Faraday discovered electrolysis.

525 (a)

1 faraday charge = E g deposition.

526 (a)

(i) The electrode with higher oxidation potential acts as anode and electrode with lower oxidation potential acts as cathode.

(ii) Cell reaction is spontaneous when E_{cell} is positive. In case of A and B ; oxidation potential of $A = -0.03$ oxidation potential of $B = +0.108$

Because oxidation potential of B is more than A so oxidation of A is not possible. The cell having A as anode and B as cathode is not possible.

Therefore, non-spontaneous cell reaction takes place between A and B .

527 (b)

$$\text{Cell constant} = \frac{k}{C} = 0.0212 \times 55 \\ = 1.166 \text{ cm}^{-1}$$

528 (a)

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.059}{n} \log \left[\frac{\text{product}}{\text{reactant}} \right]$$

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

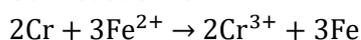
$$E_{\text{Fe}^{2+}/\text{Fe}}^{\circ} = -0.44 \text{ V}$$



$\therefore \text{Cr}^{3+}/\text{Cr}$ is anode and Fe^{2+}/Fe is cathode.

$$E_{\text{cell}}^{\circ} = E_{\text{C}}^{\circ} - E_{\text{A}}^{\circ} \\ = (-0.44) - (0.74) \\ = -0.44 + 0.74 \\ = +0.30 \text{ V}$$

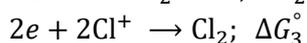
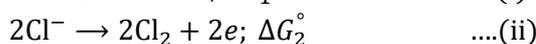
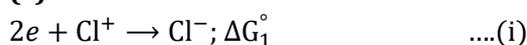
Cell reaction is



number of electrons in cell reaction = 6

$$E_{\text{cell}}^{\circ} = E_{\text{cell}}^{\circ} - \frac{0.059}{n} \log \left[\frac{\text{product}}{\text{reactant}} \right] \\ = +0.30 \text{ V} - \frac{0.059}{6} \log \left[\frac{(\text{Cr}^{3+})^2}{(\text{Fe}^{2+})^3} \right] \\ = 0.30 - \frac{0.059}{6} \log \left[\frac{(0.1)^2}{(0.01)^3} \right] \\ = 0.30 - \frac{0.059}{6} \log 10^4 \\ = 0.30 - \frac{0.059}{6} \times 0.60 \\ = 0.30 - 5.9 \times 10^{-3} \\ = 0.2941 \text{ V}$$

529 (c)



$$\therefore \Delta G_3^{\circ} = 2 \times \Delta G_1^{\circ} + \Delta G_2^{\circ}$$

$$-2 \times E_3^\circ \times F = -2 \times 2 \times 0.94 F - 2 \times (-1.36) \times F$$

$$\therefore E_3^\circ = 0.52 \text{ V}$$

530 (a)

$$E_{\text{cell}} = E_{\text{OPH}}^\circ + E_{\text{RPH}}^\circ + \frac{0.059}{2} \log \frac{[\text{H}^+]^2 \cdot P_1}{[\text{H}^+]^2 \cdot P_2}$$

$$= \frac{0.059}{+2} \log \frac{P_1}{P_2}$$

E_{cell} is +ve when $P_1 > P_2$.

532 (c)

$$\text{Eq. of O}_2 = \text{Eq. of Ag or } \frac{1.6}{8} = \frac{W}{108};$$

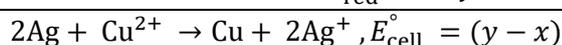
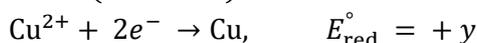
$$\therefore W_{\text{Ag}} = 21.6 \text{ g}$$

533 (c)

At LHS (oxidation) $2 \times (\text{Ag} \rightarrow \text{Ag}^+ + e^-)$,

$$E_{\text{oxi}}^\circ = -x$$

At RHS (reduction)



Note E° values remain constant when half - cell equation is multiplied / divided.

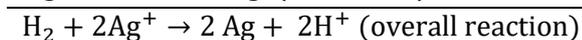
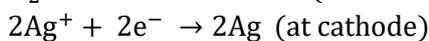
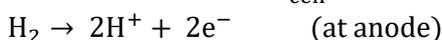
535 (a)

$$\Lambda_{\text{eq}}^\infty = \lambda_a^\infty + \lambda_c^\infty = 315 + 35 = 350.$$

536 (d)

$\text{Pt}(s), \text{H}_2(g) | \text{H}^+(1 \text{ M}) || \text{Ag}^+(aq) | \text{Ag}(s)$

$$\text{EMF of cell} = 0.62 \text{ V}, \quad E_{\text{cell}}^\circ = 0.80 \text{ V}$$



$$E_{\text{cell}} = E^\circ - \frac{2.303RT}{2F} \log \frac{[\text{H}^+]^2}{[\text{Ag}^+]^2 [\text{H}_2]}$$

$$E_{\text{cell}} = E^\circ - \frac{2.303RT}{2F} \log \frac{1}{[\text{Ag}^+]^2}$$

$$0.62 = 0.80 + \frac{2.303 \times 0.06}{2} \log \frac{1}{[\text{Ag}^+]^2}$$

$$0.62 = 0.80 + \frac{2 \times 2.303 \times 0.06}{2} \log [\text{Ag}^+]$$

$$-0.18 = 0.1382 \log [\text{Ag}^+]$$

$$[\text{Ag}^+] = 0.05 \text{ m}$$

$$\therefore \text{Mole of Ag}^+ \text{ in } 100 \text{ mL} = 0.05 \times \frac{100}{1000}$$

$$\text{Wt. of Ag}^+ \text{ in } 100 \text{ mL} = 0.05 \times \frac{100}{1000} \times 108$$

% of Ag in 1.08 g alloy

$$= \frac{0.05 \times 100 \times 108}{1000 \times 1.08} \times 100$$

$$= 50 \%$$

537 (b)

$$Q = it$$

$$= 1 \times 60$$

$$= 60 \text{ C}$$

538 (c)

During electrolysis, volumes of O_2 and H_2 liberated are in the ratio of 1 : 2 Hence, volume of H_2 liberated will be 4.48 dm^3 .

539 (b)

When 1 F electricity is passed through the solution. 1 g-equivalent of Cu is liberated.

540 (d)

Zn is placed above Sn in electrochemical series.

541 (a)

1 faraday involves charge of 1 mole electrons.

542 (c)

Oxidation of Cl^- at anode and reduction of Na^+ at cathode.

543 (d)

These are characteristic of conductivity water.

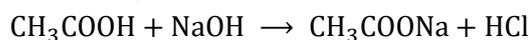
544 (c)

These are the facts about use of Mg in protecting iron against corrosion.

545 (b)

Sum of molar conductivity of reactants = sum of molar conductivity of products

Therefore, for the reaction



$$\Lambda_m^0 \text{CH}_3\text{COOH} = \Lambda_m^0 \text{CH}_3\text{COONa} + \Lambda_m^0 \text{HCl} - \Lambda_m^0 \text{NaCl}$$

$$= 91 + 425.9 - 126.4$$

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

546 (b)

$$E_{\text{cell}} = E_{\text{cell}}^\circ + \frac{0.059}{n} \log [M^{n+}]$$

$$E_{\text{cell}} = 0.34 + \frac{0.059}{2} \log 10^{-2}$$

$$= 0.34 + \frac{0.059}{2} \times -2$$

$$= +0.281 \text{ V}$$

547 (a)

We know that, $\Delta G^\circ = -nF \cdot E^\circ$

Where, $n = 1$ (number of transferred electron in the cell reaction)

$F = 96500 \text{ C}$ (Faraday's constant)

$E^\circ = ?$ (the standard electrode potential of cell)

$\Delta G^\circ = -21.20 \text{ kJ}$ (standard free energy)

$$\therefore -21.20 \times 1000 \text{ J} = -1 \times 96500 \times E^\circ$$

$$\text{Or } E^\circ = \frac{21200}{96500} = 0.220 \text{ V}$$

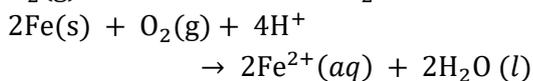
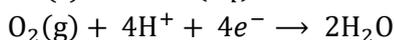
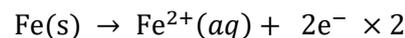
548 (b)

$$-\Delta G = nEF = nE^\circ F + RT \ln \frac{[\text{Cu}^{2+}]}{[\text{Zn}^{2+}]}$$

$\therefore \Delta G$ is function of $\ln(c_2/c_1)$.

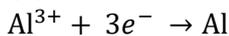
549 (d)

The half reactions are



$$E = E^\circ - \frac{0.059}{4} \log \frac{(10^{-3})^2}{(10^{-3})^4(0.1)} = 1.57 \text{ V}$$

550 (d)



$$w = ZQ$$

Where, w = amount of metal

$$w = 5.12 \text{ kg}$$

$$= 5.12 \times 10^3 \text{ g}$$

 Z = electrochemical equivalent

$$Z = \frac{\text{equivalent weight}}{96500} = \frac{\text{atomic mass}}{\text{electrons} \times 96500}$$

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

$$5.12 \times 10^3 = \frac{27}{3 \times 96500} \times Q$$

$$Q = \frac{5.12 \times 10^3 \times 3 \times 96500}{27} \text{ C}$$

$$= 5.49 \times 10^7 \text{ C}$$

551 (c)

 E_{OP}° of Mg > E_{OP}° of Zn

552 (b)

$$E_{\text{cell}} = E_{OP_{\text{Cl}}}^\circ + E_{RP_{\text{Cl}}}^\circ + \frac{0.059}{2} \log \frac{[\text{Cl}^-]_{\text{LHE}}^2 P_2}{[\text{Cl}^-]_{\text{RHE}}^2 P_1}$$

$$= \frac{0.059}{2} \log \frac{P_2}{P_1}$$

 E_{cell} is +ve when $P_2 > P_1$.

553 (c)

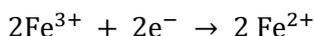
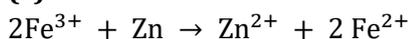
3 eq. of Ag and 3 eq. of Cu will be deposited. Na and Al will not deposit during electrolysis of their aqueous salt solution. Because both have higher E_{OP}° than E_{OP}° of H. These metals can be extracted by electrolysis of their fused salts.

554 (a)

$$t_{\text{K}^+}^\infty = t_{\text{KCl}}^\infty \times t_{\text{K}^+}$$

$$130 \times 0.495 = 64.35$$

555 (c)



According to Nernst's equation,

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{RT}{nF} \ln \frac{[\text{Fe}^{2+}]^2}{[\text{Fe}^{3+}]^2}$$

 E_{cell} is depend on concentration of Fe^{2+} .Therefore on increasing the concentration of $[\text{Fe}^{2+}] E_{\text{cell}}$ decreases.

556 (b)

$$96500 \text{ C} = 6 \times 10^{23} \text{ electrons.}$$

557 (a)

Because fluorine is most powerful oxidizing agent than other halogens

558 (d)

$$\text{Equivalent conductivity, } \Lambda_{\text{eq}} = \frac{\kappa \times 1000}{\text{normality}}$$

$$= \text{ohm}^{-1} \text{ cm}^2 (\text{g} - \text{equiv})^{-1}$$

559 (b)

$$E_{\text{cell}}^\circ = \frac{0.059}{2} \log K_c$$

$$\log K_c = \frac{E_{\text{cell}}^\circ \times 2}{0.059}$$

$$\log K_c = \frac{0.47 \times 2}{0.059}$$

$$K_c = 8.5 \times 10^{15}$$

560 (d)

Mg will not deposit on cathode during electrolysis.

561 (a)

$$\Lambda_{\text{NH}_4\text{OH}}^\infty = \Lambda_{\text{NH}_4\text{Cl}}^\infty + \Lambda_{\text{NaOH}}^\infty - \Lambda_{\text{NaCl}}^\infty$$

562 (b)

At cathode, cations are reduced. Also discharge potential of H^+ is less than Na^+ .

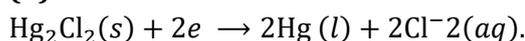
563 (a)

Conductance of a solution = conductance of ions present in solution = $\sum C_i + C_{\text{water}}$.

564 (d)

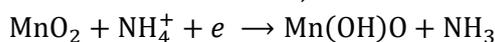
$$\Lambda_{\text{NaBr}} = \Lambda_{\text{NaCl}} + \Lambda_{\text{KBr}} - \Lambda_{\text{KCl}}$$

565 (d)



566 (b)

Zn acts as anode and gets oxidised. At graphite cathode the reaction is,

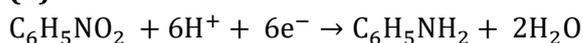


567 (a)

$$\text{Eq. of H}_2 = 8/1 = 8$$

 \therefore 8 faraday are needed.

568 (d)



1 mol = 123 g nitrobenzene requires 6 mol electrons

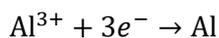
$$= 6 \times 96500 \text{ C charge}$$

$$\therefore 12.3 \text{ g nitrobenzene will require} = \frac{6 \times 96500 \times 12.3}{123}$$

$$= 6 \times 9650 = 57900 \text{ C}$$

569 (c)

At cathode,

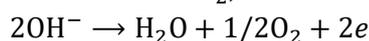
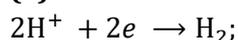


$$E_{\text{Al}} = \frac{27}{3} = 9$$

$$w_{\text{Al}} = E_{\text{Al}} \times \text{no. of faradays}$$

$$= 9 \times 0.1 = 0.9 \text{ g}$$

570 (a)



571 (c)

Weak electrolytes are 100% ionized at infinite dilution.

572 (a)

1 faraday deposits 1 equivalent which is also 1 g atom for Na.

574 (b)

$$E_1 = E_0 - \frac{0.0591}{2} \log \frac{0.01}{1} = E_0 + \frac{0.0591}{2} \times 2$$

$$E_2 = E_0 - \frac{0.0591}{2} \log \frac{1}{0.01} = E_0 - \frac{0.0591}{2} \times 2$$

$$\therefore E_1 > E_2$$

575 (a)

Sulphuric acid (H_2SO_4) used in lead storage battery.

576 (b)

It is the definition of conductivity.

577 (a)

By Kohlrausch's law

$$\Lambda_{\text{NaBr}}^\circ = \Lambda_{\text{NaCl}}^\circ + \Lambda_{\text{KBr}}^\circ - \Lambda_{\text{KCl}}^\circ$$

$$= 126 + 152 - 150$$

$$= 128 \text{ S cm}^2 \text{ mol}^{-1}$$

578 (d)

Λ vs \sqrt{c} curves are hyperbolic for weak electrolytes because on dilution their no. of ions as well as ionic mobility both increase.

579 (a)

A characteristic of electrode.

580 (a)

The process is called cathode protection where iron acts as cathode and thus, not oxidised.

581 (c)

1 faraday deposits E g of species.

582 (b)

E_{OP}° of Li is high enough.

583 (a)

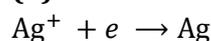


The overall reaction is



$\text{Fe}(\text{OH})_2$ may be dehydrated to iron oxide FeO , or further oxidized to $\text{Fe}(\text{OH})_3$ and then dehydrated to iron rust, Fe_2O_3 .

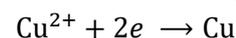
584 (b)



585 (b)

E_{OP}° of Fe $>$ E_{OP}° of Cu;

Thus, Fe gets oxidised or $\text{Fe} \rightarrow \text{Fe}^{2+} + 2e^-$;



586 (a)

Because, barring Ag, other come after the Fe in electrochemical series

587 (b)

The electrochemical cell stops working after sometime because electrode potential of both the electrodes become equal.

588 (a)

$$k = \frac{1}{R} \times \frac{1}{a} = \frac{1}{210} \times 0.66$$
$$= 3.14 \times 10^{-3} \text{ mho cm}^{-1}$$

589 (d)

$$\alpha = \frac{\Lambda_v}{\Lambda^\infty} = \frac{5.2}{390.7} = 0.0133$$

Or 1.33 %.

591 (c)

For gold plating the electrolyte $\text{K}[\text{Au}(\text{CN})_2]$ is used.

592 (c)

$$Q = i \times t = 100 \times 10^{-3} \times 30 \times 60 = 180 \text{ C}$$

594 (a)



$$E_{\text{cathode}} = -\frac{RT}{2F} \ln \frac{p_2}{[\text{H}^+]^2}$$

$$E_{\text{anode}} = -\frac{RT}{2F} \ln \frac{[\text{H}^+]^2}{p_1}$$

$$E_{\text{cell}} = E_{\text{anode}} + E_{\text{cathode}}$$

$$= -\frac{RT}{2F} \ln \frac{[\text{H}^+]^2}{p_1} - \frac{RT}{2F} \ln \frac{p_2}{[\text{H}^+]^2}$$

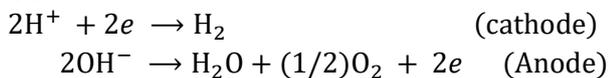
$$= -\frac{RT}{2F} \ln \frac{p_2}{p_1} = \frac{RT}{2F} \ln \frac{p_1}{p_2}$$

595 (b)

$$\begin{aligned} \text{Weight of Ag required} &= 80 \times 5 \times 10^{-3} \times \\ &1.05 (\text{wt.} = v \times d) \\ &= 0.42 \text{ g} \end{aligned}$$

$$\begin{aligned} \therefore W &= \frac{E i t}{96500} \\ \therefore 0.42 &= \frac{108 \times 3 \times t}{96500} \\ \therefore t &= 125 \text{ sec} \end{aligned}$$

596 (c)



597 (b)

$$\begin{aligned} W &= Z \times i \times t \\ \therefore W &= 0.0011180 \times 0.5 \times 200 = 0.11180 \text{ g} \end{aligned}$$

598 (c)

Zn is coated to protect iron from rusting.

599 (c)

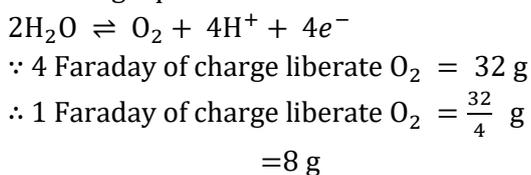
Ionisation depends upon concentration, temperature, nature of solute and on nature of solvent. Ionisation increases on increasing dilution.

600 (b)

$$\begin{aligned} \text{C}_5\text{H}_{12} + 8\text{O}_2 &\rightarrow 5\text{CO}_2 + 6\text{H}_2\text{O} \\ \Delta G^\circ &= 5 \times G^\circ\text{CO}_2 + 6 \times G^\circ\text{H}_2\text{O} - G^\circ_{\text{C}_5\text{H}_{12}} - 8 \times G^\circ_{\text{O}_2} \\ &= 5 \times (-394.4) + 6 \times (-237.2) + 8.2 \\ &= -3387 \text{ kJ} \\ \Delta G^\circ &= nE^\circ F \\ 3387 \times 10^3 &= 32 \times E^\circ \times 96500 [(C^{-2.4})_5 \rightarrow \\ &5(C^{4+}) + 32e] \\ \therefore E^\circ &= 1.0968 \text{ V} \quad n = 32 \end{aligned}$$

601 (c)

Hydrolysis of water can be represented by the following equations



602 (a)

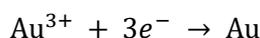
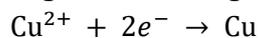
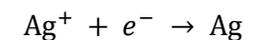
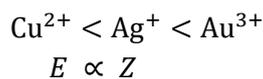
$$\begin{aligned} \text{Given, } \frac{l}{a} &= 0.5 \text{ cm}^{-1} \\ R &= 50 \Omega \\ N &= 1.0 \end{aligned}$$

$$\text{Specific conductance } (\kappa) = \frac{1}{\rho} = \frac{l}{R.a} = \frac{0.5}{50}$$

$$\begin{aligned} \Lambda &= \kappa \times \frac{1000}{N} \\ &= \frac{0.5}{50} \times \frac{1000}{0.1} \\ &= 10 \Omega^{-1} \text{ cm}^2 \text{ g eq}^{-1} \end{aligned}$$

603 (d)

The increasing order of deposition of cations at the cathode is



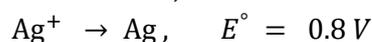
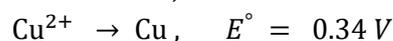
3 Faradays liberate 1 mole of Au, 3 moles of Ag and 3/2 moles of Cu. Thus, molar ratio of Ag:Cu: Au is 3:3/2:1 or 6:3:2.

604 (a)

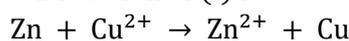
$$\begin{aligned} E^\circ_{\text{cell}} &= E^\circ_{\text{OP}_{\text{Mg}}} + E^\circ_{\text{RP}_{\text{Cu}}} \\ &= 2.37 + 0.34 = 2.71 \text{ V.} \end{aligned}$$

605 (b)

Given,

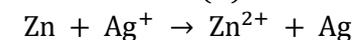


Cell reaction of (I) is



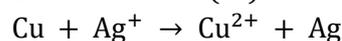
$$\begin{aligned} E^\circ_{\text{cell}} &= E_{\text{oxidation}} + E_{\text{reduction}} \\ &= +0.76 + 0.34 \\ &= +1.10 \text{ V} \end{aligned}$$

Cell reaction of (II) is



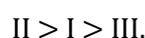
$$\begin{aligned} E^\circ_{\text{cell}} &= 0.76 + 0.8 \\ &= +1.56 \text{ V} \end{aligned}$$

Cell reaction of (III) is



$$\begin{aligned} E^\circ_{\text{cell}} &= -0.34 + 0.8 \\ &= +0.46 \end{aligned}$$

So, the correct order of E°_{cell} of these cell is



606 (c)

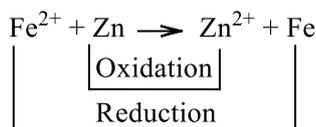
The equivalent conductance of strong electrolyte is increased by dilution, because its value is equal to the multiple of K_v and the volume of solution. By making dilution the volume of solution increases which also increase the value of equivalent conductivity.

607 (b)

Λ vs \sqrt{c} curves on extrapolating to zero concentration gives Λ^∞ for strong electrolytes. NH_4OH is weak electrolyte its Λ vs \sqrt{c} curves are shown in fig. (d) of problem 41.

608 (c)

In this reaction,



$$\text{EMF} = E_{\text{cathode}} - E_{\text{anode}}$$

$$= -0.41 - (-0.76)$$

$$= +0.35 \text{ V}$$

609 (a)

H₂ is anode because oxidation takes place. Cu is cathode because reduction takes place

610 (b)

Solid NaCl does not conduct electricity due to absence of free ions.

611 (b)

The hydrated ion size is $\text{Li}_{\text{aq}}^+ > \text{Na}_{\text{aq}}^+ > \text{K}_{\text{aq}}^+ > \text{Rb}_{\text{aq}}^+$. Larger is ion, lesser is its mobility.

612 (d)

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} + \frac{0.059}{2} \log \frac{[\text{Ni}^{2+}]}{[\text{Zn}^{2+}]}$$

$$\therefore 0.5105 = E_{\text{cell}}^{\circ} + \frac{0.059}{2} \log \frac{1}{1} \text{ or } E_{\text{cell}}^{\circ} = 0.5105 \text{ V}$$

613 (a)

The variation of e.m.f. of cell with temperature at constant pressure $(\partial E/\partial T)_P$ is referred as temperature coefficient of cell.

614 (a)

Wt. of Ag deposited = eq. wt. of Ag = 108 g

Wt. of Ni deposited = eQ. wt. of Ni = 29.5 g

Wt. of Cr deposited = eq. wt. of Cr = 17.3 g

615 (a)

The correct relation of E_{cell}° with free energy of cell reaction and equilibrium constant are as follows

$$\Delta G^{\circ} = nFE_{\text{cell}}^{\circ}$$

$$\text{and } \Delta G^{\circ} = -2.303 RT \log K_{\text{eq}}$$

616 (a)

Anode : $\text{Fe}(s) + 2\text{OH}^- \rightarrow \text{FeO}(s) + \text{H}_2\text{O}(l) + 2e$

Cathode : $\text{Ni}_2\text{O}_3 + \text{H}_2\text{O}(l) + 2e \rightarrow 2\text{NiO}(s) + 2\text{OH}^-$

$$E_{\text{cell}} = E_{\text{OP}_{\text{Fe}/\text{FeO}}}^{\circ} - \frac{0.059}{2} \log_{10} \frac{[\text{H}_2\text{O}]}{[\text{OH}^-]^2}$$

$$+ E_{\text{RP}_{\text{Ni}_2\text{O}_3/\text{NiO}}}^{\circ}$$

$$+ \frac{0.059}{2} \log_{10} \frac{[\text{H}_2\text{O}]}{[\text{OH}^-]^2}$$

$$E_{\text{cell}} = E_{\text{OP}_{\text{Fe}/\text{FeO}}}^{\circ} + E_{\text{RP}_{\text{Ni}_2\text{O}_3/\text{NiO}}}^{\circ}$$

617 (b)

$K_a = c\alpha^2$ for weak acid;



618 (d)

C is also calomel electrode.

619 (c)



At cathode $\text{Cu}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Cu}(s)$

At anode $2\text{H}_2\text{O}(l) \rightarrow \text{O}_2(g) + 4\text{H}^+(\text{aq}) + 4e^-$

Thus, for the production of one mole of copper from copper sulphate 2F of electricity is required.

620 (a)

Eq. of Al = Eq. of Ag

$$\therefore \frac{W_{\text{Al}}}{9} = \frac{W_{\text{Ag}}}{108} \text{ or } \frac{W_{\text{Al}}}{W_{\text{Ag}}} = \frac{9}{108}$$

621 (d)

$W \propto E$; if $Q = \text{constant}$.

622 (a)

From solution of CuSO_4 , Cu can be recovered by Fe metal. Because Fe is more reactive than Cu, it replace Cu easily.

623 (d)

Large negative RP or more positive oxidation potential and thus, more is the tendency to get oxidized.

624 (c)

According to Faraday's second law of electrolysis, we have

$$\frac{w_1}{w_2} = \frac{E_1}{E_2}$$

$$\therefore \frac{1.08}{x} = \frac{108}{1}$$

\therefore Weight of hydrogen (x) = 0.01 g

Hence, the volume of hydrogen at

$$\text{STP} = \frac{22400 \times 0.01}{2} = 112 \text{ cm}^3$$

625 (d)

By Nernst equation,

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{2.303 RT}{nF} \log_{10} K$$

At equilibrium $E_{\text{cell}} = 0$

Given that,

$$\therefore R = 8.315 \text{ J K}^{-1} \text{ mol}^{-1}$$

$$T = 25^\circ\text{C} + 273 = 298\text{K}$$

$$F = 96500 \text{ C and } n = 2$$

$$\therefore E_{\text{cell}}^{\circ} = \frac{2.303 \times 8.314 \times 298}{2 \times 96500} \log_{10} K$$

$$= \frac{0.0591}{2} \log_{10} K$$

\therefore Given that $E_{\text{cell}}^{\circ} = 0.295 \text{ V}$

$$\therefore 0.295 = \frac{0.0591}{2} \log_{10} K$$

$$\log_{10} K = \frac{0.295 \times 2}{0.0591} = 10$$

$$\log_{10} K = \text{antilog } 10$$

$$K = 1 \times 10^{10}$$

626 (a)

Follow E_{OP}° values in electrochemical series.

627 (c)

$$\begin{aligned}\Lambda_{\text{CH}_3\text{COOH}} &= \Lambda_{\text{CH}_3\text{COONa}} + \Lambda_{\text{HCl}} - \Lambda_{\text{NaCl}} \\ &= 91.0 + 426.2 - 126.5 \\ &= 390.7 \text{ S cm}^2 \text{ mol}^{-1}\end{aligned}$$

628 (a)

Fe is more electropositive than copper

Hence, Cu^{2+} can oxidize Fe

629 (a)

Eq. of Al = Eq. of Na;

$$\therefore \frac{1.8}{27/3} = \frac{W}{23}$$

$$\therefore W_{\text{Na}} = 4.6 \text{ g}$$

630 (a)

X coulomb deposits 1 mole Al or 3 eq. of Al and thus, it will deposit 3 mole or 3 eq. of Ag is monovalent,

$$\left[\therefore \frac{W}{E} (\text{for Ag}) = \frac{W}{E} (\text{for Al}) \right]$$

631 (d)

During electrolysis of CuSO_4 , Cu^{2+} gets discharged at cathode and OH^- at anode. Thus, solution becomes acidic due to excess of H^+ and SO_4^{2-} or H_2SO_4

Session : 2025-26

AS PER NEW NTA SYLLABUS

Total Questions : 642

CHEMISTRY (QUESTION BANK)

3.ELECTROCHEMISTRY

Assertion - Reasoning Type

This section contain(s) 0 questions numbered 1 to 0. Each question contains STATEMENT 1(Assertion) and STATEMENT 2(Reason). Each question has the 4 choices (a), (b), (c) and (d) out of which **ONLY ONE** is correct.

- a) Statement 1 is True, Statement 2 is True; Statement 2 is correct explanation for Statement 1
- b) Statement 1 is True, Statement 2 is True; Statement 2 is **not** correct explanation for Statement 1
- c) Statement 1 is True, Statement 2 is False
- d) Statement 1 is False, Statement 2 is True

1

Statement 1: In a Daniel cell, If concentrations of Cu^{2+} and Zn^{2+} ions are doubled. The emf of the cell will be doubled

Statement 2: If the concentration of ions in contact with the metal is doubled the electrodes potential will be doubled.

2

Statement 1: If $\lambda_{\text{Na}}^\circ = \lambda_{\text{Cl}^-}^\circ$ are molar limiting conductivity of sodium and chloride ions respectively, then the limiting molar conductivity of sodium chloride is given by this equation.

$$\lambda_{\text{Na}}^\circ = \lambda_{\text{Cl}^-}^\circ$$

Statement 2: This is according to Kohlrusch law of independent migration of ions.

3

Statement 1: For a galvanic cell, oxidation takes place at anode and reduction at cathode

Statement 2: Electrons in the external circuit move from anode to cathode

4

Statement 1: Galvanized iron does not rust

Statement 2: Zinc has a more negative electrode potential than iron

5

Statement 1: The cell potential of mercury cell is 1.35 V, which remains constant.

Statement 2: In mercury cell, the electrolyte is a paste of KOH and ZnO.

6

Statement 1: If an aqueous solution of NaCl is electrolysed using inert electrodes, chlorine gas is produced at the anode, and H₂ gas is produced at the cathode

Statement 2: Electrolysis of aqueous solutions containing metal ion that is a weaker oxidizing agent than H⁺(aq), will produce H₂ gas at cathode

7

Statement 1: Copper metal with HCl liberates hydrogen

Statement 2: Hydrogen is present above Cu in the reactivity series

8

Statement 1: An electrochemical cell can be set up only if the redox reaction is spontaneous.

Statement 2: A reaction is spontaneous, if free energy change is negative.

9

Statement 1: A current of 96.5 A is passed into aqueous AgNO₃ solution for 100 s. The weight of silver deposited is 10.8 g (At. Wt. of Ag = 108)

Statement 2: The mass of a substance deposited during the electrolysis of an electrolyte is inversely proportional to the quantity of electricity passing through the electrolyte.

10

Statement 1: Lead storage battery is a secondary cell.

Statement 2: Lead storage battery is used in automobiles and invertors.

11

Statement 1: On increasing dilution, the specific conductance keep on increasing

Statement 2: On increasing dilution, degree of ionization of weak electrolyte increases and molality of ions also increases

CHEMISTRY (QUESTION BANK)

3.ELECTROCHEMISTRY

: ANSWER KEY :

- | | | | | | | | | | | | | | |
|----|---|----|---|----|---|----|---|----|---|-----|---|-----|---|
| 1) | d | 2) | a | 3) | b | 4) | a | 9) | c | 10) | b | 11) | d |
| 5) | b | 6) | a | 7) | d | 8) | b | | | | | | |

CHEMISTRY (QUESTION BANK)**3.ELECTROCHEMISTRY****: HINTS AND SOLUTIONS :**

- 1 **(d)**
On doubling the concentration. The emf will remain unchanged because $[Zn^{2+}]/[Cu^{2+}]$ will remain same.
- 2 **(a)**
According to Kohlrausch law, 'Limiting molar conductivity of an electrolyte can be represented as the sum of the individuals contributions of the anion and cation of the electrolyte.
- 3 **(b)**
In a galvanic cell, anode is the negative terminal and the reaction occurring here is oxidation. Cathode is the positive terminal and the reaction occurring here is reduction. Electrons move from anode to cathode in the external circuit
- 4 **(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 oxidized, the iron start to rust
- 5 **(b)**
 $Zn(l) + Hg(s) \rightarrow ZnO(l) + Hg(l)$
The cell potential remains constant during its life as the overall reaction does not involve any ion in solution whose concentration change during its life time.
- 6 **(a)**
Na metal is not produced at the cathode (instead H_2 is formed) because water contains H_3O^+ ions
- 7 **(d)**
Copper is present below hydrogen, therefore hydrogen from HCl can't be liberated by treating it with copper
- 8 **(b)**
If redox reaction is spontaneous, ΔG is -ve and hence E° is positive.
 $-\Delta G^\circ = nFE^\circ_{cell}$
- 9 **(c)**
 $i = 96.5 A, t = 100s$
 $w = 10.8$
 $Q = it$
 $Q = 96.5 \times 100 = 9650 C$
 $\therefore 96500 C$ deposited wt. of Ag = 108 g
 $\therefore 9650 C$ deposited wt. of Ag = $\frac{108 \times 9650}{96500}$
- So, Assertion is true but Reason is false. Reason is false because mass of substance deposited is directly proportional to the quantity of electricity.
- 10 **(b)**
Leads storage battery is a secondary cell because after use it can be recharged by passing current through it in opposite direction so that it can be used again.
- 11 **(d)**
The specific conductivity decreases while equivalent and molar conductivities increase with dilution