DCAM classes
Dynamic Classes for Academic Mastery

3.ELECTROCHEMISTRY

Single Correct Answer Type

1. The desired amount of charge for obtaining one mole of Al from Al^{3+} is
a) 96500 C
b) 2×96500 C
c) 3×96500 C
d) $\frac{96500}{2}$ C

2. A certain current liberates 0.504 g of hydrogen in 2 hr. How many gram of copper can be liberated by the same current flowing for the same time in $CuSO_4$solution?
a) 12.7
b) 16
c) 31.8
d) 63.5

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

4. The Edison storage cell is represented as :
$Fe(s) +\ FeO(s)|\ KOH(aq)|\ Ni_2O_3(s)|\ Ni_2O_3(s)|\ Ni(s)$
The half reactions are $Ni_2O_3(s) + H_2O(l) + 2e^- \rightarrow 2NiO(s) + 2OH^-;\ E^\circ = +0.40$ V
$FeO(s) + H_2O(l) + 2e^- \rightarrow Fe(s) + 2OH^-;\ E^\circ = -0.87$ V
Choose the incorrect statement
a) E_{anode} increases with increase in concentration of OH^-
b) $E_{cathode}$ decreases with increase in concentration of OH^-
c) $E^\circ_{cell} = 1.27$ V
d) E_{cell} increases with increase in concentration of FeO

5. Standard reduction potentials of the half reactions are given below :
$F_2(g) + 2e^- \longrightarrow 2F^-(aq);\quad E^\circ = +2.85$ V
$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq);\quad E^\circ = +1.36$ V
$Br_2(l) + 2e^- \longrightarrow 2Br^-(aq);\quad E^\circ = +1.06$ V
$I_2(s) + 2e^- \longrightarrow 2I^-(aq);\quad E^\circ = +0.53\ V$
The strongest oxidising and reducing agents respectively are :
a) F_2 and I^-
b) Br_2 and Cl^-
c) Cl_2 and Br^-
d) Cl_2 and I_2

6. The standard reduction potential for $Fe^{2+}|Fe$ and $Sn^{2+}|Sn$ electrodes are -0.44 V and -0.14 V respectively. For the cell reaction, $Fe^{2+} + Sn \longrightarrow Fe + Sn^{2+}$, the standard e.m.f. is:
a) + 0.30 V
b) 0.58 V
c) + 0.58 V
d) − 0.30 V

7. Electrolytes when dissolved in water dissociates into ions because
a) They are unstable
b) The water dissolves it
c) The force of repulsion increases
d) The force of electrostatic attraction are broken down by water

8. Which ion has exceptionally higher Λ^∞ values?
a) H^+
b) K^+
c) NH_2^-
d) OH

9. Limiting molar ionic conductivities of a uni-univalent electrolyte are 57 and 73. The limiting molar conductivity of the solution will be :
a) 130 $S\ cm^2\ mol^{-1}$
b) 65 $S\ cm^2\ mol^{-1}$
c) 260 $S\ cm^2\ mol^{-1}$
d) 187 $S\ cm^2\ mol^{-1}$

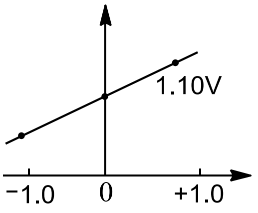
10. Molten NaCl conducts electricity due to the presence of :
a) Free electrons
b) Free molecules
c) Free ions
d) Atoms of Na and Cl

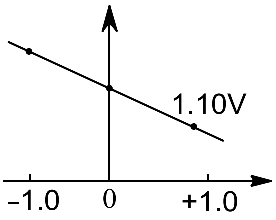
11. The emf of the cell, $(E_{Zn^{2+}/Zn} = -0.76\ V)$
Zn / Zn^{2+} (1 M) || Cu^{2+} (1 M) | Cu
$(E_{Cu^{2+}/Cu} = +0.34\ V)$ will be
a) +1.10 V
b) -1.10 V
c) +0.42 V
d) -0.42 V

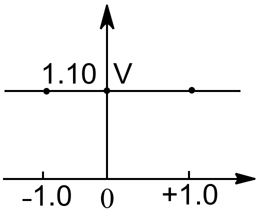
12. Which represents a concentration cell?
a) $PtH_2|HCl||\ HCl\ |PtH_2$
b) $PtH_2|HCl||Cl_2|Pt$
c) $Zn|Zn^{2+}||Cu^{2+}|Cu$
d) $Fe|Fe^{2+}||Cu^{2+}|Cu$

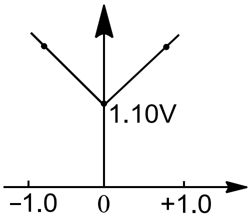
c_1 c_2 c_1

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, $M^{n+}(aq) + ne \longrightarrow M^0\ (s)$. The standard reduction potential values of the metals M_1, M_2 and M_3 are -0.34 V, -3.05 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 $CuSO_4(aq)$ 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 Mg| Mg^{2+}(0.01 M)|| Sn^{2+}(0.1 M)|Sn at 298 K is (Given, $E^{\circ}_{Mg^{2+},Mg} = -2.34$ V, -2.34 V, $E^{\circ}_{Sn^{2+},Sn} = -0.14$ 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')?
    $\text{Zn}(s) + \text{Cu}^{2+}(M) \longrightarrow \text{Zn}^{2+}(M') + \text{Cu}(s)$;
    $E^{\circ}_{\text{Cell}} = 1.10\ V$
    X – axis : $\log_{10}\frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$, Y – axis : E_{Cell}

a)


b)


c)


d)


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 $E°_{M^{3+}/M^{2+}}$ 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. Ni / Ni^{2+} [1.0 M] || Au^{3+} [1.0 M] / Au where E°
for Ni^{2+} /Ni is -0.250 V; and E° for
Au^{3+} / 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) $H_2S_2O_8$ 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. Zn (s) | Zn^{2+} (0.1 M) || Cu^{2+} (1 M) | Cu (s)
II. Zn (s) | Zn^{2+} (1 M) || Cu^{2+} (1 M) | Cu (s)
III. Zn (s) | Zn^{2+} (1 M) || Cu^{2+} (0.1 M) | Cu (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

$2H^+ + 2e^- + \frac{1}{2} O_2 \longrightarrow H_2O(l);$

$E^\circ = +1.23\ V$

$Fe^{2+} + 2e^- \longrightarrow Fe(s); \quad E^\circ = -0.44\ V$

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 cm^2 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 $Fe^{2+} + 2e^- \rightarrow Fe$ is -0.44 V and E° for $Zn^{2+} + 2e^- \rightarrow Zn$ 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 $AgNO_3$ and $CuSO_4$ 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. $I_2(s)|\, I^-(0.1\ M)$ half-cell is connected to a $H^+(aq)|\, H_2(1\ bar)|$ Pt half-cell and emf is found to be 0.7714 V. If $E^\circ_{I_2/I^-} = 0.535$ V, find the pH of H^+/H_2 half-cell

a) 1 b) 2 c) 3 d) 5

48. The $E^\circ_{M^{3+}/M^{2+}}$ 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) $Pt/H_{2(P_1)}|HCl|Pt/H_2(P_2)$

b) $Pt/H_2|\,\underset{c_1}{HCl}|\,\underset{c_2}{HCl}|\,Pt/H_2$

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. $E^\circ_{Cu} = 0.34\ V$, $E^\circ_{Zn} = 0.76\ V$. ADaniel 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 $E^\circ$of Fe^{2+} / Fe and Sn^{2+} /Sn are -0.44 V and -0.14 V respectively. If cell reaction is

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

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 $Ba(NO_3)_2$ c) 1 M $La(NO_3)_3$ d) 1 M $Th(NO_3)_4$

56. Dipping iron article into a strongly alkaline solution of sodium phosphate

a) Does not affect the article b) Forms $Fe_2O_3.xH_2O$ 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

$Cu\ (s)\ |\ Cu^{2+}\ (aq)\ ||\ Hg^{2+}\ (aq)\ |\ Hg\ (l)$ is

a) $Hg + Cu^{2+} \longrightarrow Hg^{2+} + Cu$ b) $Hg + Cu^{2+} \longrightarrow Cu^{+} + Hg^{+}$

c) $Cu + Hg \rightarrow CuHg$ d) $Cu + Hg^{2+} \longrightarrow Cu^{2+} + Hg$

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 ($F = 96500\ C\ mol^{-1}$)

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

($E_{RP} = -2.36\ V$) 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 25^0C is 0.002765 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 :

$Zn(s) + Fe^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Fe(s);$

$E^{\circ}_{cell} = 2905$ V at 298 K

a) $e^{0.32/0.0295}$ b) $10^{0.595/0.76}$ c) $10^{0.0250/0.32}$ 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_3 COOH 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 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 $Pb(NO_3)_2$ 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 $Cu(NO_3)_2$ 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 Zn | Zn^{2+} (1 M)|| Cu^{2+}|Cu(1M) is 1.1 volt. If the standard reduction potential of Zn^{2-}| Zn is – 0.78 volt, what is the oxidation potential of Cu | 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) amphere^{-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, 1120mL 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 $Cr_2O_7^{2-}$ 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 mho $cm^2 eq.^{-1}$ b) 3.14 mho $cm^2 eq.^{-1}$ c) 314.28 mho^{-1} $cm^2 eq.^{-1}$ d) 3.14 mho^{-1} $cm^2 eq.^{-1}$

82. IV. $Cu + 2HCl \rightarrow CuCl_2 + H_2(g)$

$$[E^\circ_{Cu^{2+}/Cu} = +0.34\ V]$$

V. $Zn + 2HCl \rightarrow ZnCl_2 + H_2(g)$

$$[E^\circ_{Zn^{2+}/Zn} = -0.76\ V]$$

VI. $Ag + 2HCl \rightarrow AgCl + \frac{1}{2}H_2(g)$

$[E^{\circ}_{Ag^+/Ag} = +0.80\ V]$

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 $Zn^{2+} + 2e^- \rightarrow Zn$, $E^{\circ} = -0.762$ V

$Mg^{2+} + 2e^- \rightarrow Mg, E^{\circ} = -2.37$ 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 $AgCl + e^- \rightarrow Ag + Cl^-$ 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. $\Lambda^{\infty}_{ClCH_2COONa} = 224\Omega^{-1}\ cm^2\ g\ equiv^{-1}$

$\Lambda^{\infty}_{NaCl} = 38.5\Omega^{-1}\ cm^2\ g\ equiv^{-1}$

$\Lambda^{\infty}_{HCl} = 203\Omega^{-1}\ cm^2\ g\ equiv^{-1}$

What is the value of λ_{ClCH_2COOH} =?

a) $288.5\Omega^{-1}\ cm^2\ g\ equiv^{-1}$ b) $289.5\Omega^{-1}\ cm^2\ g\ equiv^{-1}$

c) $388.5\Omega^{-1}\ cm^2\ g\ equiv^{-1}$ d) $59.5\Omega^{-1}\ cm^2\ g\ equiv^{-1}$

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 4.23×10^{-4} mho cm^{-1}. If the degree of dissociation of acid at this dilution is 0.0612, then equivalent conductivity at infinite dilution is ……mho cm^2 $eq.^{-1}$:

a) 172.8 b) 180 c) 190 d) 160

91. The highest electrical conductivity of the following aqueous solutions is of

a) 0.1 M difluoroacetic acid b) 0.1 M fluoroacetic acid

c) 0.1 M chloroacetic acid d) 0.1 M acetic acid

92. Which one is correct?

a) Ni displaces zinc from its solution

b) Zn displaces iron from its solution

c) Ag displaces copper from its solution

d) Cu displaces nickel from its solution

93. In an electrolytic cell of Ag | $AgNO_3$ | Ag, when current is passed, the concentration of $AgNO_3$:

a) Increases b) Decreases c) Remains same d) None of these

94. The resistance of 1N solution of acetic is 250Ω ,when measured in a cell having a cell constant of

$1.15 cm^{-1}$.The equivalent conduction (in $ohm^{-1}\ cm^2 equiv^{-1}$) of 1N acetic acid is

a) 2.3 b) 4.6 c) 9.2 d) 18.4

95. The standard reduction potential E° for the half reactions are as

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

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

The emf for the cell reaction,

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

a) 0.42 V b) -0.42 V c) -1.1 V d) 1.1 V

96. Ionic mobility is equal to:

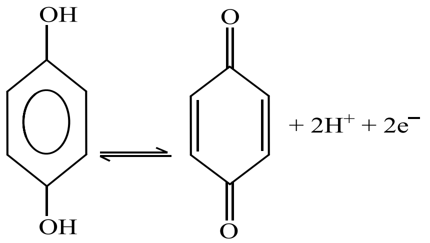
a) Speed of ions

b) Speed of ion under potential difference of 1 volt

c) Speed of ions under unit potential gradient

d) None of the above

97. At pH = 2, $E^\circ_{\text{Quinhydrone}} = 1.30\ V, E_{\text{Quinhydrone}}$ will be :



a) 1.36 V b) 1.30 V c) 1.42 V d) 1.20 V

98. The equilibrium constant for the reaction : $Cu + 2Ag^+(aq) \longrightarrow Cu^{2+}(aq) + 2Ag; E^\circ = 0.46$ V at 298 K is :

a) 2.0×10^{10} b) 4.0×10^{10} c) 4.0×10^{15} d) 2.4×10^{10}

99. For a given cell reaction; $Cr + 3H_2O + OCl^- \longrightarrow Cr^{3+}\ 3Cl^- + 6OH^-$, the species undergoing reduction is :

a) Cr b) Cr^{6+} c) OCl^- d) Cl^-

100. If the H^+ concentration is decreased from 1 M to 10^{-4} M at 25 °C for the couple MnO_4^- / Mn^{2+}, then the oxidising power of the MnO_4^- / Mn^{2+} couple decreases by

a) -0.18 V b) 0.18 V c) 0.38 V d) -0.38 V

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

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

102. If an iron rod is dipped in $CuSO_4$ solution, then :

a) Blue colour of the solution turns red

b) Brown layer is deposited on iron rod

c) No change occurs in the colour of the solution

d) None of the above

103. Which of the following liberates hydrogen on reaction with dilute H_2SO_4?

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

104. Agalvanic 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* CH_3COOH c) 0.1 *M* H_3PO_4 d) 0.1 *M* H_2SO_4

106. Which metal does not give the following reaction?

$M + \text{water or steam} \longrightarrow \text{oxide} + H_2 \uparrow$

a) Iron b) Sodium c) Mercury d) Magnesium

107. 4.5 g of Al (at. mass 27 amu) is deposited at cathode from Al^{3+} solution by a certain quantity of charge. The volume of H_2 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) $S^2\ m^2 mol$ b) $S^2\ m^2 mol^{-2}$ c) $S\ m^2 mol^{-1}$ 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 MnO_4^- of Mn^{2+}?
a) 4 b) 5 c) 3 d) 2

116. For the cell,
$T1\ |\ T1^+\ (0.001\ M)\ ||\ Cu^{2+}\ (0.1\ M)\ |\ Cu$
E_{cell} at 25°C is 0.83 V. E_{cell} can be increased
a) By decreasing $[Cu^{2+}]$
b) By increasing $[Cu^{2+}]$
c) By increasing $[T1^+]$
d) None of these

117. In an aqueous solution, hydrogen (H_2) will not reduce :
a) Fe^{3+} b) Cu^{2+} c) Zn^{2+} d) Ag^+

118. How many faradays of electricity are required to electrolyse 1 mole $CuCl_2$ 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) $E_{cell} = E_{OP}$ anode $- E_{RP}$ cathode
ii) $E_{cell} = E_{OP}$ anode $+ E_{RP}$ cathode
iii) $E_{cell} = E_{RP}$ anode $+ E_{RP}$ cathode
iv) $E_{cell} = E_{OP}$ anode $- E_{OP}$ cathode
a) (iii) and (i) b) (i) and (ii) c) (iii) and (iv) d) (ii) and (iv)

121. $Zn^{2+} \longrightarrow Zn\ (s);\ E^\circ = -0.76\ V$
$Cu^{2+} \longrightarrow Cu\ (s);\ E^\circ = -0.34\ V$
Which of the following is spontaneous?
a) $Zn^{2+} + Cu \longrightarrow Zn + Cu^{2+}$
b) $Cu^{2+} + Zn \longrightarrow Cu + Zn^{2+}$
c) $Zn^{2+} + Cu^{2+} \longrightarrow Zn + Cu$
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,
$E^\circ_{Cr^{3+}/Cr} = 0.72\ V$, $E^\circ_{Fe^{2+}/Fe} = 0.42\ V$.
The potential for the cell
Cr/Cr^{3+} (0.1 M) || Fe^{2+} (0.01 M) | Fe is
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,
$Zn^{2+} + 2e^- \rightleftharpoons Zn;\ E^\circ = -0.76$ V and
$Ag^+ + e^- \rightleftharpoons Ag;\quad E^\circ = +0.80\ V$. Which of the following reactions take place?
a) $Zn^{2+}(aq) + 2Ag(s) \rightarrow 2Ag^+(aq) + Zn(s)$ b) $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$
c) $Zn^{2+}(aq) + Ag^+(aq) \rightarrow Zn(s) + Ag(s)$ d) $Zn(s) + Ag(s) \rightarrow Zn^{2+}(aq) + Ag^+(aq)$

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^{0}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 Ametal M for which E° for the half reaction
$M^{n+} + ne^- = M$, is very negative will be Agood 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 $(NH_4)_2SnCl_6$ 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 Acell 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 $I_2 + 2e \longrightarrow 2I^-$, standard reduction potential = + 0.54 volt. For $2Br^- \longrightarrow Br_2 + 2e^-$, standard oxidation potential = − 1.09 volt. For $Fe \longrightarrow Fe^{2+} + 2e^-$, stabdard oxidation potential = + 0.44 volt. Which of the following reactions is non-spontaneous?

a) $Br_2 + 2I^- \longrightarrow 2Br^- + I_2$

b) $Fe + Br_2 \longrightarrow Fe^{2+} + 2Br^-$

c) $Fe + I_2 \longrightarrow Fe^{2+} + 2I^-$

d) $I_2 + 2Br^- \longrightarrow 2I^- + Br_2$

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) Hexaaquochromium (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

$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$; $-0.762\ V$

$Cr^{3+}(aq) + 3e^- \rightleftharpoons Cr(s)$; $-0.74\ V$

$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$; $+0.00\ V$

$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$; $+0.77\ V$

Which one of the following is the strongest reducing agent?

a) Zn (s) b) Cr(s) c) H_2 (s) d) Fe^{2+} (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)?

a) $Cu(s) + ZnSO_4(aq) \longrightarrow CuSO_4(aq) + Zn(s)$

b) $Zn(s) + CuSO_4(aq) \rightarrow Cu(s) + ZnSO_4(aq)$

c) $Ni(s) + ZnSO_4(aq) \longrightarrow NiSO_4(aq) + Zn(s)$

d) $2Na(s) + CdSO_4(aq) \longrightarrow Na_2SO_4(aq) + Cd(s)$

143.

Electrolyte	**KCl**	**KNO_3**	**HCl**	**NaOAc**	**NaCl**
Λ^{∞} (S cr mol^{-1})	149.9	145.0	426.2	91.0	126.5

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

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

144. Is the reaction, $2Al + 3Fe^{2+} \longrightarrow 2Al^{3+} + 3Fe$ possible?

a) No, because standard oxidation potential of Al $< Fe$

b) Yes, because standard oxidation potential of Al $> Fe$

c) Neither (a) nor (b)

d) Data are unpredictable

145. What will be the electrode potential of that hydrogen electrode is filled with HCl solution of pH value 1.0?
a) -59.15 V　b) +59.15　c) +59.15 mV　d) -59.15 mV

146. The conductivity of a 0.01 N solution is found to be 0.005 $\text{ohm}^{-1}\ \text{cm}^{-1}$. The equivalent conductivity of the solution will be
a) $5 \times 10^{-2}\ \text{ohm}^{-1}\ \text{cm}^2\ \text{equiv}^{-1}$　b) $5.00 \times 10^{-3}\ \text{ohm}^{-1}\ \text{cm}^2$
c) $500\ \text{ohm}^{-1}\ \text{cm}^{-2}\ \text{equiv}^{-1}$　d) $0.5\ \text{ohm}^{-1}\ \text{cm}^2\ \text{equiv}^{-1}$

147. A correct electrochemical series can be obtained from K, Ca, Na, Al, Mg, Zn. Fe, Pb, H, Cu, Hg, Ag, Au by interchanging :
a) Al and Mg　b) Zn and Fe　c) Zn and Pb　d) Pb and H

148. The emf of the cell Zn | Zn^{2+}(0.01 M) || Fe^{2+}(0.001 M) | Fe at 298 K is 0.2905. The value of equilibrium constant for the cell reaction is
a) $10^{10\frac{0.32}{0.0298}}$　b) $e^{\frac{0.32}{0.0295}}$　c) $10^{\frac{0.32}{0.0591}}$　d) $10^{\frac{0.26}{0.0295}}$

149. When Alead storage battery is discharged
a) Lead sulphate is consumed　b) SO_2 is evolved
c) Lead is formed　d) Sulphuric acid is consumed

150. EMF of hydrogen electrode in term of pH is (at 1 atm pressure)
a) $E_{H_2} = \frac{RT}{F} \times pH$　b) $E_{H_2} = \frac{RT}{F} \cdot \frac{1}{\text{p}H}$
c) $E_{H_2} = \frac{2.303RT}{F} .\text{p}H$　d) $E_{H_2} =$ - 0.0591 pH

151. If $E^{\circ}_{Fe^{2+}/Fe} = -0.441$ V and $E^{\circ}_{Fe^{3+}/Fe^{2+}} = 0.771$ V, the standard e.m.f. of the reaction $Fe + 2Fe^{3+} \longrightarrow 3Fe^{2+}$ will be :
a) 1.212 V　b) 0.111 V　c) 0.330 V　d) 1.653 V

152. When Zn piece is kept in $CuSO_4$ solution, copper gets precipitated because:
a) Standard reduction potential of zinc is more than copper
b) Standard reduction potential of zinc is less than copper
c) Atomic number of zinc is larger than copper
d) Atomic number of zinc is lower than copper

153. Ionic mobility of of electricity is 1 *M* solution of :
a) CH_3COOH　b) H_2SO_4　c) H_3PO_4　d) Boric acid

154. The equivalent conductivity of 0.1 *M* weak acid is 100 times less than that at infinite dilution. The degree of dissociation of weak electrolyte at 0.1 *M* is :
a) 100　b) 10　c) 0.01　d) 0.001

155. Standard electrode potential of cell $H_2|H^+||Ag^+|Ag$ is (Given, $E°_{Ag^+/Ag} = 0.80$ V)
a) 0.4 V　b) 0.8 V　c) 1.4 V　d) 1.8 V

156. If the current is passed into the solution of an electrolyte:
a) Anions move towards anode, cations towards cathode
b) Anions and cations both move towards anode
c) Anions move towards cathode, cations towards anode
d) No movement of ions takes place

157. The element that is easiest to be reduced is :
a) Fe　b) Cu　c) Ag　d) Sn

158. Standard reduction potential for, $Li^+|$ Li, $Zn^{2+}|Zn$, $H^+|H_2$ and $Ag^+ |$ Ag is $-3.05, -0.762, 0.00$ and $+$ 80 V. Which has highest reducing capacity?
a) Ag　b) H_2　c) Zn　d) Li

159. What is the quantity of electricity (in Coulombs) required to deposit all the silver from 250mL of 1 $MAgNO_3$ solution?
a) 2412.5　b) 24125　c) 4825.0　d) 48250

160. When 1 faraday of electricity is passed through $CuSO_4$ solution, number of atoms formed is :
a) 6.02×10^{23}　b) 3.01×10^{23}　c) 2　d) 6.02×10^{23}

161. Hydrogen gas is not liberated when the following metal is added to dil. HCl

a) Ag b) Zn c) Mg d) Sn

162. In Ahydrogen-oxygen fuel cell, combustion of hydrogen occurs to

a) Generate heat
b) Create potential difference between the two electrodes
c) Produce high purity water
d) Remove adsorbed oxygen from electrode surfaces.

163. The sum of the two transport number of ions for an electrolyte is always equal to :

a) 1 b) 2 c) 1/2 d) None of these

164. On passing 0.5 F electricity through molten sodium chloride, sodium deposited at cathode will be

a) 29.25 g b) 11.50 g c) 58.50 g d) 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?

a) 2.096 g b) 0.296 g c) 3.029 g d) 2.906 g

166. The correct order of molar conductivity at infinite dilution of LiCl, NaCl and KCl is

a) LiCl $> KCl > NaCl$ b) KCl $> NaCl > LiCl$ c) LiCl $> NaCl > KCl$ d) NaCl $> KCl > LiCl$

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 :

a) 3,1 and 2 b) 1, 3 and 2 c) 3,1 and 3 d) 2,3 and 2

168. Indicator electrode is :

a) SHE
b) Calomel electrode
c) Ag/AgCl electrode
d) Quinhydrone electrode

169. Molar conductance of electrolytic solution Λ_m is

a) $\propto l$ b) $\propto (1/A)$ c) $\propto (1/C)$ d) $\propto (\sqrt{C})$

170. Which metal is most readily corroded in moist air?

a) Copper b) Iron c) Silver d) 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 ($1F = 96500$ C mol^{-1})

a) $9.65 \times 10^4\ s$ b) $19.3 \times 10^4\ s$ c) $28.95 \times 10^4\ s$ d) $38.6 \times 10^4\ s$

172. The oxidation potential of Mg and Al are $+2.37$ and $+1.66$ volt respectively. The Mg in chemical reactions :

a) Will be replaced by Al
b) 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 Ag | Ag^+, Sn | Sn^{2+}, Ca | Ca^{2+}, Pb | 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
$Mn^{2+} + 2e^- \longrightarrow$ and $Mn^{3+} + e^- \longrightarrow M^{2+}$ are -1.18 V and 1.51 V respectively. What is the redox potenital for the reaction
$Mn^{3+} + 3e^- \longrightarrow Mn$?
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 dilition 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) $Zn^{2+}(aq) + 2e \longrightarrow Zn(s)$; $[Zn^{2+}] = 1M$
b) $Zn(g) \longrightarrow Zn^{2+} + 2e$; $[Zn^{2+}] = 1M$
c) $Zn^{2+}(aq) \longrightarrow Zn(s) + 2e$; $[Zn^{2+}] = 1M$
d) $Zn^{2+}(g) \longrightarrow Zn(s) - 2e$; $[Zn^{2+}] = 1M$

192. Given, the data at 25 °C ,
$Ag + I^- \longrightarrow AgI + e^-$; $E^\circ = 0.152\ V$
$Ag \rightarrow Ag^+ + e^-$; $E^\circ = -0.800V$
What is the value of $\log K_{sp}$ for AgI?

$\left(2.303\ \dfrac{RT}{F} = 0.059\ V\right)$

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 $S\ cm^2\ mol^{-1}$ at 298 K. The molar conductivity of CH_3COOH at the same temperature is .... $S\ cm^2\ mol^{-1}$.

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 liberates ....litre of O_2at 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) $\Omega\ cm^2 equiv^{-1}$　b) $\Omega\ cm^2$ equiv　c) $\Omega^{-1}\ cm^2 equiv^{-1}$　d) $\Omega\ cm^2$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 $126.4\Omega^{-1}\ cm^2 equiv^{-1}$. 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

a) Copper electrode acts as Acathode then E°_{cell} is +0.46 V

b) Silver electrode acts as anode then E°_{cell} is -0.34 V

c) Copper electrode acts as anode then E°_{cell} is +0.46 V

d) Silver electrode acts as Acathode then E°_{cell} is -0.34 V

205. e.m.f. of cell $Ni|Ni^{2+}(0.1M)||\ Au^{3+}(1.0\ M)|Au$ is ….., if $E°$ for $Ni^{2+}|Ni$ is -0.25 V, $E°$ for $Au^{3+}|Au$ is 1.50 V.

a) + 1.25 V　b) −1.75 V　c) + 1.75 V　d) + 4.0 V

206. The position of some metals in the electrochemical series in decreasing electropositive character is given as $Mg > Al > Zn > Cu > Ag$. What will happen, if a copper spoon is used to stir a solution of aluminium nitrate?

a) The spoon will get coated with aluminium

b) An alloy of copper and aluminium is formed

c) The solution becomes blue

d) There is no reaction

207. Which of the following statements is correct? Galvanic cell converts
  a) Chemical energy into electrical energy
  b) Electrical energy into chemical energy
  c) Metal from its elemental state to the combined state
  d) Electrolyte into individual ions

208. For cell reaction,
  $Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$
  Cell representation is
  a) $Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$ b) $Cu \mid Cu^{2+} \parallel Zn^{2+} \mid Zn$
  c) $Cu \mid Zn^{2+} \parallel Zn \mid Cu^{2+}$ d) $Cu^{2+} \mid Zn \parallel Zn^{2+} \mid Cu$

209. By passing 9.65 A current for 16 min 40 s , the volume of O_2 liberated at STP will be
  a) 280 mL   b) 560 mL   c) 1120 mL   d) 2240 mL

210. Consider the following disproportionation
  $2ClO_3^- \rightleftharpoons ClO_2^- + ClO_4^-$
  If the initial concentration of perchlorate ion is 0.1 M what it would be at equilibrium at 298 K?
  $\left(E^\circ_{ClO_4^-/ClO_3^-} = 0.36 \text{ V and } E^\circ_{ClO_3^-/ClO_2^-} = 0.33 \text{ V}\right)$
  a) 0.1 M   b) 0.05 M   c) 0.07 M   d) 0.19 M

211. When Cu reacts with $AgNO_3$ solution, the reaction takes place is
  a) Oxidation of Cu   b) Reduction of Cu   c) Oxidation of Ag   d) Reduction of NO_3^-

212. E° for $F_{2+} + 2e = 2F^-$ is 2.8 V, E° for $1/2\ F_2 + e = F^-$ is:
  a) 2.8 V   b) 1.4 V   c) $-$ 2.8 V   d) $-$ 1.4 V

213. Which one of the following solutions has highest conductance power?
  a) 0.1 M CH_3COOH b) 0.1 M NaCl   c) 0.1 M KNO_3 d) 0.1 M HCl

214. Standard electrode potentials of $Fe^{2+} + 2e \longrightarrow Fe$ and $Fe^{+} + 3e \longrightarrow Fe$ are $-$ 440 V and -0.036 V respectively. The standard electrode potential (E°) for $Fe^{3+} + e \longrightarrow Fe^{2+}$ is:
  a) $-$ 0.476 V   b) $-$ 0.404 V   c) + 0.404 V   d) + 0.772 V

215. Stainless steel does not rust because
  a) Chromium and nickel combine with iron
  b) Chromium forms an oxide layer and protects iron from rusting
  c) Nickel present in it, does not rust
  d) 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^{+}?
  a) With KCl   b) With KI   c) With both (a) and (b)   d) None of these

217. The main function of the salt bridge is :
  a) 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 $Na_2S_4O_6$ 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 Mg^{2+} / Mg is -2.37 V, of Zn^{2+} / Zn is -0.76 V and Fe^{2+} / 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 Asolution 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° - 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 :

$\mathring{\Lambda}_{CH_3COONa(aq.)} = 91.0\ S\ cm^2\ eq.^{-1}$

$\mathring{\Lambda}_{HCl(aq.)} = 426.2\ S\ cm^2\ eq.^{-1}$

What additional information one needs to calculate $\mathring{\Lambda}$ of an aqueous CH_3COOH?

a) $\mathring{\Lambda}$ of $ClCH_2COOH$
b) $\mathring{\Lambda}$ of CH_3COOK
c) $\mathring{\Lambda}$ of NaCl
d) $\mathring{\Lambda}$ of H^+

225. In the electrolysis of $CuSO_4$, the reaction $Cu^{2+} + 2e^- \longrightarrow Cu$, 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 3.86×10^{-5} 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 molility 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?
$Cu\,(s) + 2Ag^+\,(aq) \rightleftharpoons Cu^{2+}\,(aq) + 2\,Ag\,(s)$
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. $Sn^{4+} + 3e^- \longrightarrow Sn^{2+}$, $E^\circ = 0.13\,V$
$Br_2 + 2e^- \longrightarrow 2Br^-$, $E^\circ = 1.08\,V$
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) $ohm^{-1}\,cm^{-1}$ b) $ohm^{-1}\,cm^{-1}$ or Sm^{-1} c) $ohm\,m^{-1}$ d) $ohm\,cm^{-1}$

251. Ionic mobility of Ag^+ is
$(\lambda_{Ag^+} = 5 \times 10^{-1}\,\Omega^{-1}\,cm^2\ equiv^{-1})$
a) 5.2×10^{-9} b) 2.4×10^{-9} c) 1.52×10^{-9} d) 8.25×10^{-9}

252. $E^{\circ}_{Fe^{3+}/Fe} = -0.036\text{ V}$, $E^{\circ}_{Fe^{2+}/Fe} = -0.439\text{ V}$. The value of standard electrode potential for the charge, $\text{Fe}^{3+}\,(aq) + e^- \longrightarrow \text{Fe}^{2+}\,(aq)$ 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

$$\text{Zn}^{2+}/\text{Zn} = -0.74\,V,\ \text{Cl}_2\,/\,\text{Cl}^- = 1.36\,V$$

$$\text{H}^+/\frac{1}{2}\text{H}_2 = 0\,V \text{ and } \text{Fe}^{2+}\,/\,\text{Fe}^{3+} = 0.77\,V$$

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 ($\mathring{\Lambda}_m$) at infinite dilution of NaCl, HCl and CH_3COONa are 126.4, 425.9 and 91.0 S $\text{cm}^2\text{ mol}^{-1}$ respectively. $\mathring{\Lambda}_m$ for CH_3COOH will be :
a) 425.5 S $\text{cm}^2\text{ mol}^{-1}$
b) 180.5 S $\text{cm}^2\text{ mol}^{-1}$
c) 290.8 S $\text{cm}^2\text{ mol}^{-1}$
d) 390.5 S $\text{cm}^2\text{ mol}^{-1}$

256. KCl(aq) cannot be used as a salt bridge for the cell $\text{Cu}(s)|\text{CuSO}_4(aq)||\,\text{AgNO}_3(aq)|\text{Ag}(s)$ 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 :

$$\frac{2}{3}\text{Al}_2\text{O}_3 \longrightarrow \frac{4}{3}\text{Al} + \text{O}_2, \Delta_\text{r}G = +966\text{ kJ mol}^{-1}.$$

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) $\Delta G = -\dfrac{nF}{E_\text{cell}}$
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 :
a) Temperature is increased
b) Inert electrodes are used
c) A mixture of electrolytes is used
d) In none of the above cases

261. Conductance ‘C’ (in S) in 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:
a) S m mol^{-1}
b) $\text{S m}^2\text{ mol}^{-1}$
c) $\text{S}^{-2}\text{ m}^2\text{ mol}$
d) $\text{S}^2\text{ m}^2\text{ mol}^2$

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?
a) Mg
b) Fe
c) Cu
d) Cd

263. For the cell reaction $\text{Fe} + 2\text{Fe}^{3+} = 3\text{Fe}^{2+}$, which is not possible?
a) One cell can be constructed
b) Three different cells with different E°_cell are possible
c) Three different cells with different number of electrons used in redox reaction are possible
d) Three different cells with same ΔG° value are possible

264. Copper sulphate solution does not react with

a) Zinc b) Iron c) Silver d) All of these

265. Rust is
a) $FeO + Fe(OH)_2$ b) Fe_2O_3
c) $Fe_2O_3 + Fe(OH)_2$ d) Fe_2O_3 and $Fe(OH)_3$

266. The conductivity of strong electrolyte
a) Increases on dilution slightly b) Decreases on dilution
c) Does not change with dilution d) 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 :
a) $E = \frac{c \times t}{P \times 96500}$ b) $E = \frac{c \times P}{t \times 96500}$ c) $E = \frac{96500 \times P}{c \times t}$ d) $E = \frac{c \times t \times 9650}{P}$

268. Cu^+ ion is not stable in aqueous solution because of disproportionation reaction. E° value for disproportionation of Cu^+ is $\left(\text{given}, E^\circ_{Cu^{2+}/Cu^+} = 0.15,\ E^\circ_{Cu^{2+}/Cu} = 0.34 \text{ V}\right)$
a) $+0.38$ V b) -0.38 V c) $+0.49$ V d) -0.49 V

269. The molar conductivities Λ°_{NaOAc} and Λ°_{HCl} at infinite dilution in water at 25°C are 91.0 and 426.2 S cm^2/mol respectively. To calculate Λ°_{HOAc}, the additional value required is
a) $\Lambda^\circ_{H_2O}$ b) Λ°_{KCl} c) Λ°_{NaOH} d) 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 :
a) 201.28 $ohm^{-1}\ cm^2\ mol^{-1}$
b) 698.28 $ohm^{-1}\ cm^2\ mol^{-1}$
c) 390.71 $ohm^{-1}\ cm^2\ mol^{-1}$
d) 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 :
a) 628.56 $mho^{-1}\ cm^2$ molb) 628.56 mho $cm^2\ mol^{-1}$c) 6.28 mho $cm^2\ mol^{-1}$ d) 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 :
a) 0.495 b) 0.505 c) 0.0495 d) Cannot be predicted

273. A cell in which electric current is produced by net oxidation and reduction process is called :
a) Voltaic cell b) Electrolytic cell c) Concentration cell d) None of these

274. Electrolysis of aq. Solution of LiCl shows :
a) $pH < 7$ b) $pH = 7$ c) $pH > 7$ d) No change

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

276. The specific conductance of 0.1 N KCl solution at 23°C is 0.012 $ohm^{-1}\ cm^{-1}$. The resistance of cell containing the solution at the same temperature was found to be 55 ohm. The cell constant will be
a) 0.66 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, $M \mid M^+ \| X^- \mid X$,
$E^\circ\ (M^+ \mid M) = 0.44\ V, E^\circ = (X \mid X^-) = 0.33\ V$. From this datAone can deduce that
a) $E^\circ_{cell} = -0.77\ V$
b) $M^+ + X^- \rightarrow M + X$ is the spontaneous reaction
c) $M + X \rightarrow M^+ + X^-$ is the spontaneous reaction
d) $E^\circ_{cell} = 0.77\ V$

279. The standard reduction potential of the reaction,
$H_2O + e^- \longrightarrow 1/2\ H_2 + OH$ at 298 K is

a) $E^\circ = \frac{RT}{F}$ Iln K_w
b) $E^\circ = -\frac{RT}{F}$ In$[P_{H_2}]^{1/2}[OH^-]$
c) $E^\circ = -\frac{RT}{F}$ In$\frac{[p_{H_2}]^{1/2}}{[H^+]}$
d) $E^\circ = -\frac{RT}{F}$ In K_w

280. The correct order $\overset{\circ}{E}_{M^{2+}/M}$ values with negative sign for the four successive elements Cr, Mn, Fe and Co is :
a) Cr > Mn > Fe > Co
b) Mn > Cr > Fe > Co
c) Cr > Fe > Mn > Co
d) Fe > Mn > Cr > 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 50 Ω . 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 $Sm^2 mol^{-1}$
b) 6.25×10^{-4} $Sm^2 mol^{-1}$
c) 625×10^{-4} $Sm^2 mol^{-1}$
d) 62.5 $Sm^2 mol^{-1}$

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
$2Cu^+ (aq) \rightleftharpoons Cu^{2+} (aq) + Cu (s)$
choose correct E° for the above reaction if
$\overset{\circ}{E}_{Cu^{2+}/Cu} = 0.34$ V , $\overset{\circ}{E}_{Cu^{2+}/Cu^+} = 0.15$ V
a) -0.38 V
b) +0.49 V
c) +0.38 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. Agalvanic cell is constructed using the redox reaction,
$\frac{1}{2}H_2 (g) + AgCl(s) \rightleftharpoons H^+ (aq) + Cl^- (aq) + Ag (s)$
It is represented as
a) Pt | H_2(g) | HCl solution || $AgNO_3$ solution | Ag
b) Ag | AgCl(s)|KCl solution | |HCl solution| | H_2(g) | Pt
c) Pt | H_2(g) | KCl solution || AgCl(s) | Ag
d) Pt | H_2(g) , HCl solution || AgCl(s) | Ag

287. Zn | Zn^{2+} (A = 0.1 M) || Fe^{2+} (A = 0.01 M) | 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^2area, 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

$\Lambda°_{CH_3COONa} = 91.0$ S cm^2/equiv

$\Lambda°_{HCl} = 426.2$ S cm^2/equiv

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

a) $\Lambda°$ of NaCl

b) $\Lambda°$ of CH_3 COOK

c) The limiting equivalent conductance of $H^+ (\lambda°_{H^+})$

d) $\Lambda°$ of chloroacetic acid ($ClCH_2$ COOH)

293. The emf of the cell

Ni | Ni^{2+} (1.0 M) || Au^{3+} (1.0 M) | Au

is [E° (Ni^{2+}/ Ni) =-0.25 V and

$E°$ (Au^{3+}/ Au) =+1.5 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, $Fe^{2+} + Sn \rightarrow Fe + Sn^{2+}$, the standard emf is

a) 0.42 V b) – 0.42 V c) – 0.30 V d) – 1.10 V

295. In Acell that utilises the reaction,

$Zn\ (s) + 2H^+\ (aq) \rightarrow Zn^{2+}(aq) + H_2\ (g)$

addition of H_2SO_4 to cathode compartment, will

a) 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 $Sn(s) + 2Ag^+(aq) \longrightarrow Sn^{2+}(aq) + 2Ag(s)$?

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

$Fe^{2+} + 2e^- \longrightarrow Fe$ $E°$ = -0.440 V

$Fe^{3+} + 3e^- \longrightarrow Fe$ $E°$ = -0.036 V

The standard electrode potential ($E°$) for

$Fe^{3+} + e^- \longrightarrow Fe^{2+}$ is

a) +0.772 V b) -0.772 V c) +0.417 V d) -0.414 V

298. Normal Al – $AlCl_3$ 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 NH_4OH at infinite dilution is .... S cm^2 eq.$^{-1}$ .Given $\lambda^\infty_{OH^-} = 174, \lambda^\infty_{Cl^-} = 66$ and $\lambda^\infty_{NH_4Cl} =$ 130 S cm^2 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 Cu^{2+} 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 Al^{3+} 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) $E^\circ = \dfrac{2.303RT}{nF} \log K_{eq.}$
c) $-\Delta G^\circ = RT \ln K_{eq\cdot}$
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 A_2B is given by :
a) $\lambda^\infty_{A^+} + \lambda^\infty_{B^{-2}}$ b) $\frac{1}{2}\lambda^\infty_{A^+} + \lambda^\infty_{B^{-2}}$ c) $\lambda^\infty_{A^+} + \frac{1}{2}\lambda^\infty_{B^{-2}}$ d) $2\lambda^\infty_{A^+} + \lambda^\infty_{B^{-2}}$

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) O_2 d) H^+

309. The equivalent conductivity of monobasic acid at infinite dilution is 348 $ohm^{-1}\ cm^2\ eq^{-1}$. 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
$Zn = Zn^{2+} + Ze$ $E^\circ = +0.76$ V
$Fe = Fe^{2+} + Ze$ $E^\circ = +0.41$ V
The emf of the cell reaction
$Fe^{2+} + Zn = Zn^{2+} + Fe$ 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) $Cu + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag$ b) $Zn + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag$
c) $H_2 + Ni^{2+}(aq) \rightarrow 2H^+(aq) + Ni$ d) $Zn + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu$

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 : $Cr_2O_7^{2-} | Cr^{3+}, E^\circ_{red} = +1.33$ V and Electrode B : $Fe^{3+}|Fe^{2+}, E^\circ_{red} = 0.77$ 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 :
$Fe^{3+}(aq) + e \longrightarrow Fe^{2+}(aq); E^\circ = +0.77$ V
$Al^{3+}(aq) + 3e \longrightarrow Al(s); E^\circ = -1.66$ V
$Br_2\,(aq) + 2e \longrightarrow 2Br^-(aq); E^\circ = +1.08$ V
a) $Br^- < Fe^{2+} < Al$ b) $Fe^{2+} < Al < Br^-$ c) $Al < Br^- < Fe^{2+}$ d) $Al < Fe^{2+} < Br^-$

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 Agalvanic 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
$E^\circ_{Fe^{3+}/Fe^{2+}} = +0.77\,V$
$E^\circ_{Sn^{2+}/Sn} = -0.14\,V$
Under standard conditions the potential for the reaction
$Sn\,(s) + 2Fe^{3+}(aq) \rightarrow 2Fe^{2+}(aq) + Sn^{2+}(aq)$ is
a) 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
$Zn^{2+} + 2e^- \longrightarrow Zn;\ E^\circ = -0.76$ V
$Fe^{2+} + 2e^- \longrightarrow Fe;\ E^\circ = -0.44$ V
The emf of the cell reaction,
$Fe^{2+} + Zn \longrightarrow Zn^{2+} + Fe$ is
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 negative if
a) $p(H_2) = 1$ atm and $[H^+] = 2.0$ M b) $p(H_2) = 1$ atm and $[H^+] = 1.0$ M
c) $p(H_2) = 2$ atm and $[H^+] = 1.0$ M d) $p(H_2) = 2$ atm and $[H^+] = 2.0$ 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) $H_2(g)$ and $O_2(g)$ b) $O_2(g)$ and $H_2(g)$ c) $O_2(g)$ and Mg(s) d) $O_2(g)$ and $SO_2(g)$

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 $Al_2(SO_4)_3$. Given that $\Lambda^\infty_{Al^{3+}}$ and $\Lambda^\infty_{SO_4^{2-}}$ are the equivalent conductances at infinite dilution of the respective ions?
a) $2\Lambda^\infty_{Al^{3+}} + 3\Lambda^\infty_{SO_4^{2-}}$ b) $2\Lambda^\infty_{Al^{3+}} + \Lambda^\infty_{SO_4^{2-}}$ c) $(\Lambda^\infty_{Al^{3+}} + \Lambda^\infty_{SO_4^{2-}}) \times 6$ d) $\frac{1}{3}\Lambda^\infty_{Al^{3+}} + \frac{1}{2}\Lambda^\infty_{SO_4^{2-}}$

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
$Pt(H_2, 1\text{ atm}) \mid 0.1\text{ M HCl} \mid \mid 0.1\text{ M acetic acid} \mid$
$(H_2, 1\text{ atm})Pt$ 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 Abase for electrochemical cell?
a) $H_2 + O_2 \longrightarrow H_2O$
b) $AgNO_3 + Zn \rightarrow Zn(NO_3)_2 + Ag$
c) $AgNO_3 + NaCl \rightarrow AgCl\downarrow + NaNO_2$
d) $KMnO_4 + FeSO_4 + H_2SO_4 \longrightarrow K_2SO_4 + Fe_2(SO_4)_3 + MnSO_4 + H_2O$

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,
$Zn(s) \longrightarrow Zn^{2+} + 2e;\ E° = 0.76\text{ V}$
$Ag(s) \longrightarrow Ag^{+} + e;\ E° = -0.80\text{ V}$
Which reaction actually takes place?
a) $Zn(s) + 2Ag^{+}(aq) \longrightarrow Zn^{2+}Ag(s)$
b) $Zn^{2+} + 2Ag^{+}(s) \longrightarrow 2Ag^{+}(aq) + Zn(s)$
c) $Zn(s) + 2Ag(s) \longrightarrow Zn^{2+}(aq) + Ag^{+}(aq)$
d) $Zn^{2+}(aq) + Ag^{+}(aq) \longrightarrow Zn(s) + Ag(s)$

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 $[H^+]$ 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 pKa 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 Asolution 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.

Pt | H_2(1 atm) | H^+ (1 M) || Cu^{2+} (1 M) | 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 $\left(\text{the value of } \frac{2.303RT}{T} \text{ is } 0.059 \text{ V}\right)$

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) Voltameter 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,

$M\ (s) + 2H^+\ (1M) \longrightarrow H_2\ (g)\ (1\ atm) + M^{2+}\ (0.1\ M)$'

is 1.500 V. The standard reduction potential for $M^+/\ M\ (s)$ couple is

a) 0.1470 V b) 1.470 V c) 14.70 V d) None of these

345. For the cell Zn|Zn^{2+}|| Cu^{2+}|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 mho cm^2eq^{-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 Ω^{-1} cm^2 eq^{-1}. Its conductivity would be

a) $2.9 \times 10^{-3}\ \Omega^{-1}\ cm^2$ b) $1.8 \times 10^{-2}\ \Omega^{-1}\ cm^2$ c) $2.4 \times 10^{-4}\ \Omega^{-1}\ cm^2$ d) $3.6 \times 10^{-3}\ \Omega^{-1}\ cm^2$

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 ;

$MnO_4^-(aq) + 8H^+(aq) + 5e^- \longrightarrow$
$Mn^{2+}(aq) + 4H_2O; E° = 1.51\ V$

$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \longrightarrow$
$2Cr^{3+}(aq) + 7H_2O; E° = 1.38\ V$

$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq); E° = 0.77\ V$

$Cl_2(g) + 2e^- \longrightarrow 2Cl^-\ (aq); E° = 1.40\ V$

Identify the only incorrect statement regarding the quantitative estimation of aqueous $Fe(NO_3)_2$:

a) MnO_4^- can be used in aqueous HCl

b) $Cr_2O_7^{2-}$ can be used in aqueous HCl

c) MnO_4^- can be used in aqueous H_2SO_4

d) $Cr_2O_7^{2-}$ can be used in aqueous H_2SO_4

349. The standard H electrode is written as :

a) $Pt, H_2, H^+(a = 1)$

b) $PtH_2/H^+(a = 1)$

c) $PtH_2(g)(1\ atm)/H^+(a = 1)$

d) None of the above

350. Standard electrode potential of cell H_2 | H^+ ||Ag^+ | 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

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

Given,

$E^\circ_{Cu^{2+}/cu} = +0.34\ V, E^\circ_{Ag^+/Ag} = 0.80\ V$,

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) $2.5\Omega^{-1}\ cm^2\ equiv^{-1}$ b) $5.0\Omega^{-1}\ cm^2\ equiv^{-1}$

c) $2.5\Omega^{-1}\ cm^{-2}\ equiv^{-1}$ d) $5.0\Omega^{-1}\ cm^{-2}\ equiv^{-1}$

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

$5Ce^{4+}(aq) + Mn^{2+}(aq) + 4H_2O(l) \rightarrow 5Ce^{3+}(aq) + MNO_4^-(aq) + 8H^+(aq)$

$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)\ E^\circ = +1.51$ V

$Ce^{4+}(aq) + e^- \rightarrow Ce^{3+}(aq)\ E^\circ = +1.61$ 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 + Salt of Y (or Z) = Y (or Z) + Salt of X. Which observation is probably incorrect?

a) Y + Salt of X = No action observed

b) Y + Salt of Z = Nothing can be decided
c) Z + Salt of X = X + Salt of Z
d) Z + Salt of X = No action observed

364. During the charging of lead storage battery, the reaction at anode is represented by :
a) $Pb^{2+} + SO_4^{2-} \longrightarrow PbSO_4$
b) $PbSO_4 + H_2O \longrightarrow PbO_2 + SO_4^{2-} + 2H^+$
c) $Pb \longrightarrow Pb^{2+} + 2e$
d) $Pb^{2+} + 2e \longrightarrow Pb$

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) $\Lambda_{eq} = \kappa \times V_{\text{in mL}}$ 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. $Pt(H_2)| H^+(1\ M)||\ Ag^+|Ag$
The E_{cell} is 0.503 V at 298 K. The percent of lead in the ore is $\left(E^{\circ}_{Ag^+|Ag} = 0.80\ V\right)$
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 mho cm^2 eq $^{-1}$ .If equivalent conductivity of the acid at infinite dilution is 350 mho cm^2 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,
$Pt; \frac{1}{2} H_2(g)|H^+(10^{-8}\ M)||H^+(0.001\ M)|1/2H_2(g)Pt$?
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, $Zn + Cu^{2+} \rightarrow Cu + Zn^{2+}$ 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) $Pb(NO_3)_2$ b) $Zn(NO_3)_2$ c) $AgNO_3$ d) $Cd(NO_3)_2$

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 S cm^2mol^{-1} respectively. What is the molar conductivity of $NaNO_3$?
a) 101 S cm^2mol^{-1} b) 87 S cm^2mol^{-1} c) -101 S cm^2mol^{-1} d) -391 S cm^2mol^{-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.
$Mg\ (s)\ |\ Mg^{2+}(aq)\ x\ M\ ||\ Fe^{2+}(aq)\ 0.01\ M\ |\ Fe\ (s)$
$E^{\circ}\ Mg/Mg^{2+}\ (aq) = 2.37\ V\ \ E^{\circ}\ Fe/Fe^{2+}(aq) = 0.45\ V$

a) $x = 0.01\ M$ b) $x < 0.01\ M$
c) $x > 0.01\ M$ 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
$2MnO_2(s) + Zn^2 + 2e^- \longrightarrow ZnMn_2O_4(s)$
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,
Cd (s) | $CdCl_2$ (aq) (0.1 M) || AgCl (s) | Ag (s)
In which the cell reaction is
Cd (s) + 2AgCl (s)
$\rightarrow$ 2 Ag (s) + Cd^{2+} (aq) + $2Cl^-$ (aq)
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) -167.26kJ

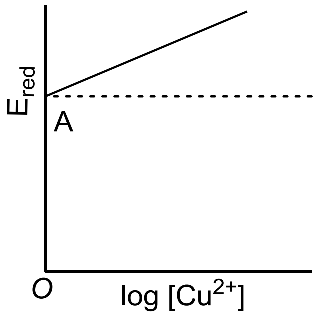
380. The factor which is not affecting the conductivity of any solution is
a) Temperature b) Dilution c) Nature of electrolyte d) None of these

381. The standard reduction potential for the half-cell having reaction
$NO_3^-(aq) + 2H^+(aq) + e^- \rightarrow NO_2(g) + H_2O$
Is 0.78 V. What will be the reduction potential of the half-cell is 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 $Cu(NO_3)_2$ and molten $Al(NO_3)_3$ 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 10^7 erg c) An unit of energy d) All of these

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



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

385. If ϕ denotes standard reduction potential, which is true:
a) $E^\circ_{cell} = \phi_R - \phi_L$ b) $E^\circ_{cell} = \phi_L + \phi_R$ c) $E^\circ_{cell} = \phi_L - \phi_R$ d) $E^\circ_{cell} = (\phi_L + \phi_R)$

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) $K^+ > Rb^+ > Na^+ > Li^+$

b) $Rb^+ > K^+ > Na^+ > Li^+$
c) $Li^+ > Na^+ > K^+ > Rb^+$
d) $Na^+ > K^+ > Rb^+ > Li^+$

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 100Ω. 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 520Ω. The molar conductivity of 0.02 M solution of the electrolyte will be
a) 124×10^{-4} $S\ m^2\ mol^{-1}$ b) 1240×10^{-4} $S\ m^2\ mol^{-1}$
c) 1.24×10^{-4} $S\ m^2\ mol^{-1}$ d) 12.4×10^{-4} $S\ m^2\ mol^{-1}$

390. Ionic mobility of Ag^+ at infinite dilution is :
$(\lambda^0_{Ag^+} = 5 \times 10^{-3}\ S\ m^2\ eq^{-1})$
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 voltameters connected in series. The first voltameter contains $XSO_4\ (aq)$ while the second voltameter contains $Y_2SO_4\ (aq)$. The relative atomic masses ox 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
$Fe^{2+} + 2e^- \rightarrow Fe, E^\circ = -0.440$ 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) + 0.771 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
[E° for $Zn^{2+}/Zn = -0.76$; E° for $Cu^{2+}/Cu = +0.34$]
a) $-0.76 + 0.34 = -0.42$ V
b) $-0.34 - (-0.76) = +0.42$ V
c) $0.34 - (-0.76) = +1.10$ V
d) $-0.76 - (+0.34) = -1.10$ V

398. The speed of migration of Ag^+ ion and NO_3^- ion are 0.00057 cm sec^{-1} and 0.00063 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 Agalvanic 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?

$\mathrm{Pt} \mid \mathrm{Br_2(g)} \mid \mathrm{Br^-(g)} \parallel \mathrm{Cl^-}(aq) \mid \mathrm{Cl_2(g)} \mid \mathrm{Pt}$

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

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., $\mathring{\Lambda}_m(NH_4OH)$ is equal to :

a) $\mathring{\Lambda}_m(NH_4OH) + \mathring{\Lambda}_m(NH_4Cl) - \mathring{\Lambda}_m(HCl)$
b) $\mathring{\Lambda}_m(NH_4Cl) + \mathring{\Lambda}_m(NaOH) - \mathring{\Lambda}_m(NaCl)$
c) $\mathring{\Lambda}_m(NH_4Cl) + \mathring{\Lambda}_m(NaCl) - \mathring{\Lambda}_m(NaOH)$
d) $\mathring{\Lambda}_m(NaOH) + \mathring{\Lambda}_m(NaCl) - \mathring{\Lambda}_m(NH_4Cl)$

407. Given :

i) $Cu^{2+} + 2e^- \longrightarrow Cu, \quad E° = 0.337$ V
ii) $Cu^{2+} + e^- \longrightarrow Cu^+, \quad E° = 0.153$ V
Electrode potential, $E°$ for the reaction,
$Cu^+ + e^- \longrightarrow Cu$, 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) $2SO_4^{2-} \longrightarrow S_2O_3^{2-} + 2e$
b) $Cu^{2+} + 2e \longrightarrow Cu$
c) $2H_2O \longrightarrow O_2 + 4H^+ + 4e$
d) $2H^+ + 2e \longrightarrow H_2$

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) $A^{3-} \longrightarrow A^{2-} + e; \quad E° = 1.5$ V
ii) $B^{2+} + e \longrightarrow B^+; \quad E° = -2.1$ V
iii) $C^{2+} + e \longrightarrow C^+; \quad E° = +0.5$ V
iv) $D \longrightarrow D^{2+} + 2e; \quad E° = -1.5$ 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. $Zn^{2+} + 2e^- \rightarrow Zn(s), E^\circ = -0.76$
$Fe^{3+} + e^- \rightarrow Fe^{2+}, E^\circ = -0.77$
$Cr^{3+} + 3e^- \rightarrow Cr, E^\circ = -0.79$
$H^+ + e^- \rightarrow \frac{1}{2} H_2, E^\circ = 0.00$
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
$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s); \quad E^\circ = -0.762\ V$
$Cr^{3+}(aq) + 3e^- \longrightarrow Cr(s); \quad E^\circ = -0.740\ V$
$2H^+(aq) + 2e^- \longrightarrow H_2(g); \quad E^\circ = 0.00\ V$
$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq); \quad E^\circ = +0.762\ V$
The strongest reducing agent is
a) Zn (s) b) Cr (s) c) H_2(g) d) Fe^{2+}(aq)

415. Strong electrolytes are those which
a) Conduct electricity b) Dissolve readily in water
c) Dissociate into ions at high dilution d) Completely dissociate into ions at all dilutions

416. The cell reaction of Acell is
$Mg(s) + Cu^{2+}(aq) \longrightarrow Cu(s) + Mg^{2+}(aq)$.
If the standard reduction potentials of Mg and Cu are – 2.37 and + 0.34 V respectively. The emf of the cell is
a) 2.03 V b) -2.03 V c) +2.71 V d) -2.71 V

417. Consider the following reaction :
$2Ag^+C_6H_{12}O_6 + H_2O \rightarrow 2Ag(s) + C_6H_{12}O_7 + 2H^+$
When ammonia is added to the solution, pH is raised to 11. Which half-cell reaction is affected by pH and by how much?
$Ag^+ + e^- \rightarrow Ag; \; E^\circ_{\text{red}} = 0.8\ V$
$C_6H_{12}O_6 + H_2O \rightarrow C_6H_{12}O_7 + 2H^+ + 2e^-; \; E^\circ_{\text{red}} = -0.05\ V$
$Ag(NH_3)_2^+ + e^- \rightarrow Ag(s) + NH_3; \; E^\circ_{\text{red}} = 0.337\ V$
a) $E^\circ_{\text{oxid.}}$ will increase by a factor of 0.65 from $E^\circ_{\text{oxid.}}$ b) $E^\circ_{\text{oxid.}}$ will decrease by a factor of 0.65 from $E^\circ_{\text{oxid.}}$
c) $E^\circ_{\text{red.}}$ will increase by a factor of 0.65 from $E^\circ_{\text{red.}}$ d) $E^\circ_{\text{red.}}$ will decrease by a factor of 0.65 from $E^\circ_{\text{red.}}$

418. Which process occurs in the electrolysis of aqueous solution of nickel chloride at nickel anode?
a) $Ni^{2+} + 2e \longrightarrow Ni$ b) $2H^+ + 2e \longrightarrow H_2$ c) $2Cl^- \longrightarrow Cl_2 + 2e$ d) $Ni \longrightarrow Ni^{2+} + 2e$

419. A solution containing one mole per litre of each $Cu(NO_3)_2$, $AgNO_3$, $Hg_2(NO_3)_2$ and $Mg(NO_3)_2$ is being electrolysed by using inert electrodes. The values of standard electrode potentials in volts (reducing potentials) are $Ag/Ag^+ = +0.80, 2Hg/Hg^{2+} = +0.79, Cu/Cu^{2+} = +0.34, Mg/Mg^{2+} = -2$ With increasing voltage, the sequence of deposition of metals on the cathode will be
a) Ag, Hg, Cu b) Cu, Hg, Ag c) Ag, Hg, Cu, Mg d) Mg, Cu, Hg, Ag

420. The metal that does not displace hydrogen from an acid is
a) Ca b) Al c) Zn d) Hg

421. For an electrolyte $AxBy$, the molar conductivity at infinite dilution can be given by :
a) $\Lambda^\circ_M = x\lambda^\circ A^{Y+} + y\lambda^\circ B^{X-}$

b) $\Lambda^\circ_{\mathrm{M}} = \frac{1}{x}\lambda^\circ A^{Y+} + \frac{1}{y}\lambda^\circ B^{X-}$
c) $\Lambda^\circ_{\mathrm{M}} = \frac{1}{y}\lambda^\circ A^{Y+} + \frac{1}{x}\lambda^\circ B^{X-}$
d) $\Lambda^\circ_{\mathrm{M}} = \lambda^\circ A^{Y+} + \lambda^\circ B^{X-}$

422. ` for $\mathrm{Sn^{4+}/Sn^{2+}}$ couple is + 0.15 V and that for the $\mathrm{Cr^{3+}}$/ Cr couple is − 0.74 V. These two couples in their standard state are connected to make a cell. The cell potential will be :
a) + 1.83 V b) + 1.19 V c) + 0.89 V d) + 0.18 V

423. The standard reduction potentials for $\mathrm{Zn^{2+}/Zn, Ni^{2+}/Ni}$ and $\mathrm{Fe^{2+}/Fe}$ are − 0.76, −0.23 and −0.44 V respectively. The reaction $X + Y^{2+} \longrightarrow X^{2+} + Y$ will be spontaneous when :
a) $X = \mathrm{Ni}, Y = \mathrm{Zn}$ b) $X = \mathrm{Fe}, Y = \mathrm{Zn}$ c) $X = \mathrm{Zn}, Y = \mathrm{Ni}$ d) $X = \mathrm{Ni}, Y = \mathrm{Fe}$

424. Given that $E^\circ_{\mathrm{Fe^{3+}|Fe}}$ and $E^\circ_{\mathrm{Fe^{3+}|Fe}}$ are −0.36 V and − 0.439 V, respectively. The value of $E^\circ_{\mathrm{Fe^{3+},Fe^{2+}|Pt}}$ would be :
a) (−36 − 0.439)V
b) [3(−0.36) + 2(−0.439)]V
c) (−0.36 + 0.439)V
d) [3(−0.36) − 2(−0.439)]V

425. The standard emf for the cell reaction,
$2\mathrm{Cu}^+\,(aq) \rightarrow 2\mathrm{Cu}\,(s) + \mathrm{Cu}^+\,(aq)$
is +0.36 V at 298 K. The equilibrium constant of the reaction is
a) 5×10^6 b) 1.4×10^{12} c) 7.4×10^{12} d) 1.2×10^6

426. Electrolytic conduction is due to migration of :
a) Protons b) Electrons c) Ions d) All of these

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) $2\mathrm{H_2(g)} + \mathrm{O_2(g)} \rightarrow 2\mathrm{H_2O}(l)$
b) $2\mathrm{Fe}(s) + \mathrm{O_2(g)} + 4\mathrm{H^+}(aq) \rightarrow 2\mathrm{Fe^{2+}}(aq) + 2\mathrm{H_2O}(l)$
c) $\mathrm{Pb}(s) + \mathrm{PbO_2}(s) + 2\mathrm{H_2SO_4}(aq) \rightarrow 2\mathrm{PbSO_4}(s) + \mathrm{H_2O}(l)$
d) $\mathrm{Cd}(s) + 2\mathrm{Ni(OH)_3}(s) \rightarrow \mathrm{CdO}(s) + 2\mathrm{Ni(OH)_2} + 2\mathrm{H_2O}(l)$

429. When 1 F of electricity is passed through acidulated water, $\mathrm{O_2}$ evolved is
a) 11.2 $\mathrm{dm^3}$ b) 5.6 $\mathrm{dm^3}$ c) 22.4 $\mathrm{dm^3}$ d) 1.0 $\mathrm{dm^3}$

430. The Gibbs energy for the decomposition of $\mathrm{Al_2O_3}$ at 500°C is as follows :
$\frac{2}{3}\mathrm{Al_2O_3} \longrightarrow \frac{4}{3}\,\mathrm{Al} + \mathrm{O_2};$
$\Delta_r G = +960\ \mathrm{kJ\ mol^{-1}}.$
The potential difference needed for the electrolytic reduction of aluminium oxide ($\mathrm{Al_2O_3}$) 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 $\mathrm{Br_2}$ from KBr?
a) HI b) $\mathrm{I_2}$ c) $\mathrm{Cl_2}$ d) $\mathrm{SO_2}$

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) $u_{\mathrm{H^+}} > u_{\mathrm{Na^+}}$ (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 $Al = 27 \text{ g mol}^{-1}$)

a) 1.3×10^4 g b) 9.0×10^3 g c) 8.05×10^4 g d) 2.4×10^5 g

437. The reaction,

$$\frac{1}{2} H_2(g) + AgCl(s) \rightarrow H^+ aq) + Cl^-(aq) + Ag(s)$$

Occurs in the galvanic cell

a) $Pt/H_2(g)$ KCl (sol)|| $AgCl(s)$|Ag
b) $Pt/H_2(g)$ HCl(sol)|| $AgNO_3$ (sol)| Ag
c) $Pt/H_2(g)$ HCl(sol)|| $AgCl(s)$| Ag
d) Ag/$AgCl(s)$KCl(sol)||$AgNO_3$|Ag

438. Which of the following ions can be replaced by H^+ inos 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, $Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$ 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) $\lambda_\infty = u_\infty \times$ Faraday b) $u_\infty = \lambda_\infty \times$ Faraday c) Faraday $= u_\infty \times \lambda_\infty$ d) None of these

441. Coulomb is the quantity of current defined as :

a) One ampere of current passing for 1 sec
b) One which deposits 0.001118 g of Ag on cathode
c) One which deposits electrochemical equivalence of metal
d) All of the above

442. The standard electrode potential is measured by

a) Electrometer b) Voltmeter c) Pyrometer d) 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 :

a) Violet due to formation of I_2
b) Orange due to formation of Br_2
c) Colourless due to formation of F_2
d) 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

a) 9.3 b) 19.3 c) 38.3 d) 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?

a) 4.68×10^{18} b) 2.34×10^{18} c) 1.24×10^{18} d) 0.46×10^{18}

446. If 'F' is faraday and 'N' is Avogadro number, then charge of electron can be expressed as

a) $F \times N$ b) $\frac{F}{N}$ c) $\frac{N}{F}$ d) F^2N

447. By how much is the oxidizing power of $Cr_2O_7^{2-}/Cr^{3+}$ couple decreased if the H^+ concentration is decreased from 1 M to 10^{-3} M at 25°C?

a) 0.207 V b) 0.414 V c) 0.001 V d) 0.287 V

448. Which process involves corrosion?

a) Brown deposits on iron articles
b) Green deposits on battery terminals

c) Black deposits on silver coin
d) All of the above

449. The electric conduction of a salt solution in water depends on the
a) Size of its molecules
b) Shape of its molecules
c) Size of solvent molecules
d) Extent of its ionization

450. The electrode potentials for
$$\mathrm{Cu^{2+}}(aq) + e^- \longrightarrow \mathrm{Cu^+}(aq)$$
and $\mathrm{Cu^+}(aq) + e^- \longrightarrow \mathrm{Cu}(s)$
are + 0.15 V and + 0.50 V respectively. The value of $E^\circ_{\mathrm{Cu^{2+}/Cu}}$ will be :
a) 0.150 V
b) 0.500 V
c) 0.325 V
d) 0.650 V

451. By diluting a weak electrolyte , specific conductivity (K_c) and equivalent conductivity (λ_c) change as
a) Both increase
b) K_c increases , λ_c decreases
c) K_c decreases , λ_c increases
d) Both decrease

452. The cell reaction for the given cell is :
$$\underset{P_1=1\text{ atm}}{\mathrm{Pt(H_2)}}\,\big|\,\mathrm{pH}=2\,\big|\big|\,\underset{P_2=1\text{ atm}}{\mathrm{pH}=3}\,\big|\,\mathrm{Pt(H_2)}$$
a) Spontaneous
b) Non-spontaneous
c) In equilibrium
d) Either of these

453. If the molar conductance value of $\mathrm{Ca^{2+}}$ and $\mathrm{Cl^-}$ at infinite dilution are respectively 118.88×10^{-4} $\mathrm{m^2\ mho\ mol^{-1}}$ and 77.33×10^{-4} $\mathrm{m^2\ mho\ mol^{-1}}$ then that of $\mathrm{CaCl_2}$ is (in $\mathrm{m^2\ mho\ mol^{-1}}$)
a) 118.88×10^{-4}
b) 154.66×10^{-4}
c) 273.54×10^{-4}
d) 196.21×10^{-4}

454. During electrolysis, the species discharged at cathode are
a) Anion
b) Cation
c) Ions
d) All of these

455. In the electrolysis of which solution, $\mathrm{OH^-}$ ions are discharged in preference to $\mathrm{Cl^-}$ 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) $\Lambda_\mathrm{M} = b\sqrt{c} - \Lambda^\infty$
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 expanded 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
$$\underset{p_1}{\mathrm{Pt\ (H_2)}}\ |\ \underset{1\text{ M}}{\mathrm{H^+}\ (aq)}\ |\ \underset{p_2}{Pt\ \mathrm{(H_2)}}$$
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 C)

a) 10 b) 50 c) 1000 d) 100

465. For the redox reaction,

$\mathrm{Zn}\,(s) + \mathrm{Cu}^{2+}\,(0.1\ \mathrm{M}) \rightarrow$
$\mathrm{Zn}^{2+}\,(1\ \mathrm{M}) + \mathrm{Cu}\,(s)$

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 $\Lambda°$ for NaCl, KBr and KCl are 126, 152 and 150 S cm^2 mol^{-1} respectively. The $\Lambda°$ for NaBr is

a) 128 S cm^2 mol^{-1} b) 248 S cm^2 mol^{-1} c) 328 S cm^2 mol^{-1} d) 348 S cm^2 mol^{-1}

467. The emf of the cell,

Ag | Ag^+ (0.1 M) || Ag^+ (1 M) | 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 ;

$\mathrm{Sn}(s) + 2\mathrm{Fe}^{3+}(aq) \longrightarrow 2\mathrm{Fe}^{2+}(aq) + \mathrm{Sn}^{2+}(aq)$ is :

(Given $E^{\circ}_{\mathrm{Fe}^{3+}/\mathrm{Fe}^{2+}} = 0.77$ V and $E^{\circ}_{\mathrm{Sn}^{2+}/\mathrm{Sn}} = -0.14$ 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) $\alpha \longrightarrow 1$, 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) $Zn^{2+} + 2e \longrightarrow Zn(s)$; $E° = -0.762\ V$
(ii) $Cr^{3+} + 3e \longrightarrow Cr(s)$; $E° = -0.740\ V$
(iii) $2H^+ + 2e \longrightarrow H_2(g)$; $E° = -0.000\ V$
(iv) $Fe^{3+} + e \longrightarrow Fe^{2+}$; $E° = +0.77\ V$
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) $Zn \mid Zn^{2+}$ and $Cu^{2+} \mid Cu$ b) $Cu \mid Cu^{2+}$ and $Zn^{2+} \mid Zn$ c) $Fe \mid Fe^{2+}$ and $Cu^{2+} \mid Cu$ d) $Cu \mid Cu^{2+}$ and $Fe^{2+} \mid Fe$

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 $\Delta G°$ will be :
a) – 89.0 kJ b) – 89.0 J c) – 44.5 kJ d) – 98.0 kJ

483. E^0 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

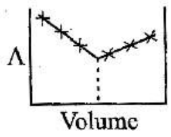
485. Which of the following metal can replace zinc from $ZnSO_4$ solution?
a) Cu b) Hg c) Fe d) Al

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- seArocks

487. $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$; $E° = 1.51$ V
$MnO_4 + 4H^+ + 2e^- \rightarrow Mn^{2++}\ 2H_2O$; $E° = 1.23$ V
$E°_{MnO_4^- \mid MnO_2}$ 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



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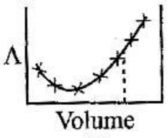
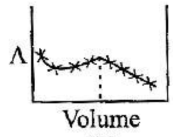
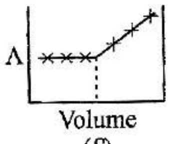


(Q)



(R)



(S)

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

490. For the electrochemical cell, $M \mid M^+ \mid\mid X^- \mid X, E^\circ(M^+ \mid M) = 0.44$ V and $E^\circ(X \mid X^-) = 0.33$ V. From this data one can deduce that
a) $E_{\text{cell}} = 0.77$ 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,
$E^0(K^+/K) = -2.9$ V,
$E^0(Zn^{2+}/Zn) = -0.76$ V, $E^0(H^+/H_2) = -0.00$V,
$E^0(Cu^{2+}/Cu) = +0.34$ 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. $2Fe^{3+} + 3I^- \rightleftharpoons 2Fe^{2+} + I_3^-$
The standard reduction potentials in acidic conditions are 0.77 and 0.54 V respectively for Fe^{3+} / 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
$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
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 an 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, Zn | Zn^{2+} (1 M) || Cu^{2+} (1 M) | Cu
($E^\circ_{\text{cell}} = 1.10$ V), was allowed to be completely discharged at 298 K. The relative concentration of
Zn^{2+} to Cu^{2+} $\left(\frac{[Zn^{2+}]}{[Cu^{2+}]}\right)$ is

a) 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 ;
$Zn \longrightarrow Zn^{2+} + 2e; \quad E^\circ = +0.76$ V
$Ag \longrightarrow Ag^+ + e; \quad E^\circ = -0.77$ V
The standard e. m. f. of the cell,
$Ag^+ + Zn \longrightarrow Zn^{2+} + Ag$ 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 $Fe(OH)_2$ b) FeO and $Fe(OH)_3$ c) Fe_2O_3 and $Fe(OH)_3$ d) Fe_3O_4 and $Fe(OH)_3$

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 is .....if the electrodes in the cell are 2.2 cm apart and have an area of 3.8 cm^2.

a) 25.73 b) 30.75 c) 35.75 d) 15.75

507. The SI unit for ionic mobility is:

a) $m^2 volt^{-1} sec^{-1}$ b) $cm^2 volt^{-1} sec^{-1}$ c) $cm\ volt^{-1} sec^{-1}$ d) $cm^{-2} volt^{-1} sec^{-1}$

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) $CuSO_4$ c) $AuCl_3$ d) $AuCl_3 + NaCN$

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 Adevice 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. Astandard 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 $AgNO_3$, NaCl and $NaNO_3$ are 116.5, 110.3 and 105.2 mho $cm^2 mol^{-1}$ 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 × 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;
$\overset{\ominus}{A}|A^+(xM)||B^+(yM)|\overset{\oplus}{B}$
The e.m.f. measured is + 0.20 V. The cell reaction is :
a) The cell reaction cannot be predicted
b) $A + B^+ \longrightarrow A^+ + B$
c) $A^+ + B \longrightarrow A + B^+$
d) $A^+ + e^- \longrightarrow A; B^+ + e^- \longrightarrow B$

521. The laws of electrolysis ware 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 20⁰C is $0.0212\ \text{ohm}^{-1}\ \text{cm}^{-1}$ and the resistance of cell containing this solution at 20⁰ C is 55 ohm. The cell constant is:

a) $2.173\ \text{cm}^{-1}$ b) $1.166\ \text{cm}^{-1}$ c) $4.616\ \text{cm}^{-1}$ d) $3.324\ \text{cm}^{-1}$

528. What is the value of E_{cell}?

$\text{Cr} \mid \text{Cr}^{3+}\ (0.1\ \text{M}) \parallel \text{Fe}^{2+}\ (0.01\ \text{M}) \mid \text{Fe}$

Given, $E°_{\text{Cr}^{3+}/\text{Cr}} = -0.74\ V$

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

a) +0.2941 V b) +0.5212 V c) +0.1308 V d) -0.2606 V

529. The $E°$ for OCl^-/Cl^- and $\text{Cl}^-/\frac{1}{2}\text{Cl}_2$ are 0.94 V and $-1.36\ V$; $E°$ for $\text{OCl}^-/\frac{1}{2}\text{Cl}_2$ 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 :

$\underset{P_1}{\text{Pt}(\text{H}_2)}|\text{H}^+(1M)||\ \text{H}^+\ (1M)|\underset{P_2}{\text{Pt}(\text{H}_2)}$

a) $P_1 > P_2$ b) $P_1 < P_2$ c) $P_1 = P_2$ d) $P_1 = 1$ atm

531. When an acid cell is charged, then

a) Voltage of cell increases b) Resistance of cell increases
c) Electrolyte of cell dilutes d) None of the above

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

a) 1.6 g b) 0.8 g c) 21.6 g d) 107.88 g

533. For Acell given below,

$\text{Ag} \mid \text{Ag}^+ \parallel \text{Cu}^{2+} \mid \text{Cu}$

$-\qquad\qquad +$

$\text{Ag}^+ + \text{e}^- \longrightarrow \text{Ag}, \quad \overset{\circ}{E} = x$

$\text{Cu}^{2+} + 2\text{e}^- \longrightarrow \text{Cu}, \quad \overset{\circ}{E} = y$

$\overset{\circ}{E}_{\text{cell}}$ is

a) $x + 2y$ b) $2x + y$ c) $y - x$ d) $y - 2x$

534. EMF of a cell in terms of reduction potential of its left and right electrodes is

a) $E = E_{\text{left}} - E_{\text{right}}$ b) $E = E_{\text{right}} - E_{\text{left}}$ c) $E = E_{\text{left}} + E_{\text{right}}$ d) None of these

535. At 18⁰C the conductance of H^+ and CH_3COO^- at infinite dilution are 315 and 35 mho $\text{cm}^2\ \text{eq}^{-1}$ respectively. The equivalent conductivity of CH_3COOH at infinite dilution is ……mho $\text{cm}^2\ \text{eq.}^{-1}$:

a) 350 b) 280 c) 30 d) 315

536. An alloy of Pb-Ag weighing 1.08 g was dissolved in dilute HNO_3 and the volume made to 100 mL. A silver electrode was dipped in the solution and the emf of the cell set up

$\text{Pt}(s), \text{H}_2\ (g) \mid \text{H}^+\ (1\ \text{M}) \parallel \text{Ag}^+\ (aq) \mid \text{Ag}\ (s)$

Was 0.62 V. If $\overset{\circ}{E}_{\text{cell}} = 0.80$ V, what is the percentage of Ag in the alloy?

[At 25°C , RT /F = 0.06]

a) 25 b) 2.50 c) 10 d) 50

537. A lamp draws a current of 1.0 A. Find the charge in coulomb used by the lamp in 60 s.

a) 0.6 C b) 60 C c) 600 C d) 0.006 C

538. During electrolysis of water the volume of O_2 liberated is 2.24dm^3. The volume of hydrogen liberated, under same conditions will be

a) 2.24 dm^3 b) 1.12 dm^3 c) 4.48 dm^3 d) 0.56 dm^3

539. The amount of electricity required to liberate 1 g-equiv of Cu is

a) 96500 F b) 1 F c) 1 C d) 96500 A

540. Which of the following is correct?

a) Zinc acts as cathode in Daniell cell

b) In a Li − Zn couple, zinc acts as anode

c) Copper will displace iron in solution

d) Zinc displaces tin from its solution

541. The number of electrons involved in redox reactions when a faraday of electricity is passed through an electrolyte in solution is :

a) 6×10^{23} b) 8×10^{19} c) 69500 d) 6×10^{-23}

542. During electrolysis of fused sodium chloride, the reaction of the electrodes are:

	Anode	**Cathode**
a)	$\text{Na}^+ + e \longrightarrow \text{Na}$	$\text{Cl}^- \longrightarrow \frac{1}{2}\text{Cl} + e$
b)	$\text{Na} \longrightarrow \text{Na}^+ + e$	$\frac{1}{2}\text{Cl}_2 + e \longrightarrow \text{Cl}^-$
c)	$\text{Cl}^- \longrightarrow \frac{1}{2}\text{Cl}_2 + e$	$\text{Na}^+ + e \longrightarrow \text{Na}$
d)	$\frac{1}{2}\text{Cl}_2 + e \longrightarrow \text{Cl}^-$	$\text{Na} \longrightarrow \text{Na}^+ + e$

543. Which one is correct about conductivity water?

a) The water whose own conductance is very small

b) The water obtained after 7-8 times distillation

c) Kohlrausch prepared the conductivity water for the first time

d) All of the above

544. Blocks of magnesium metal are often strapped to the steel hulls of ocean going ships in order to:

a) Provide cathodic protection

b) Protect oxidation of steel

c) Both (a) and (b) are correct

d) Neither (a) nor (b) is correct

545. Given the limiting molar conductivity as

$\Lambda_m^0\text{ (HCl)} = 425.9\Omega^{-1}\text{ cm}^2\text{ mol}^{-1}$

$\Lambda_m^0\text{ (NaCl)} = 126.4\Omega^{-1}\text{ cm}^2\text{ mol}^{-1}$

$\Lambda_m^0\text{ (CH}_3\text{COONa)} = 91\Omega^{-1}\text{ cm}^2\text{ mol}^{-1}$

The molar conductivity , at infinite dilution, of acetic acid (in $\Omega^{-1}\text{cm}^2\text{ mol}^{-1}$) will be

a) 481.5 b) 390.5 c) 299.5 d) 516.9

546. If the standard electrode potential of Cu^{2+} / Cu electrode is 0.34 V, what is the electrode potential at 0.01 M concentration of Cu^{2+}?

$(T = 298\,K)$

a) 0.399 V b) 0.281 V c) 0.222 V d) 0.176 V

547. If the ΔG° of Acell reaction,

$\text{AgCl} + e^- \rightarrow \text{Ag}^+ + \text{Cl}^-\ is\ -21.20\text{ kJ};$

The standard emf of the cell is

a) 0.220 V b) -0.220 V c) 0.229 V d) -0.110 V

548. For the cell reaction, $\text{Cu}^{2+}_{c_2}(aq) + \text{Zn}(s) \longrightarrow \text{Zn}^{2+}_{c_1}(aq) + \text{Cu}(s)$, the change in free energy (ΔG) at a given temperature is a function of :

a) In c_1 b) In (c_2/c_1) c) In $(c_1 + c_2)$ d) In c_2

549. Consider the following cell reaction

$2Fe(s) + O_2(g) + 4H^+(aq) \longrightarrow$
$2Fe^{2+}(aq) + 2H_2O(l), E^\circ = 1.67\ V$
At$[Fe^{2+}] = 10^{-3}$ M , $P(O_2) = 0.1$ atm and pH = 3, the cell potential at 25°C is
a) 1.47 V b) 1.77 V c) 1.87 V d) 1.57 V

550. Aluminium oxide may be electrolysed at 1000°C to furnish aluminium metal (atomic mass = 27 u; 1 F = 96500 C). The cathode reaction is
$Al^{3+} + 3e^- \longrightarrow Al^0$
To prepare 5.12 kg of aluminium metal by this method would require
a) 5.49×10^1 C of electricity b) 5.49×10^4 C of electricity
c) 1.83×10^7 C of electricity d) 5.49×10^7 C of electricity

551. The standard potentials at 25°C for the following half reactions are given ahead,
$Zn^{2+} + 2e \longrightarrow Zn$; $E^\circ = -0.762$ V
$Mg^{2+} + 2e \longrightarrow Mg$; $E^\circ = -2.37$ V
When zinc dust is added to the solution of $MgCl_2$:
a) $ZnCl_2$ is formed
b) Zinc dissolves in the solution
c) No reaction takes place
d) Mg is precipitated

552. The cell reaction for the given cell is spontaneous if :
$\underset{P_1}{Pt_{Cl_2}} | Cl^-(1M) || Cl^-(1M) | \underset{P_2}{Pt_{Cl_2}}$
a) $P_1 > P_2$ b) $P_1 < P_2$ c) $P_1 = P_2$ d) $P_1 = 1$ atm

553. Passage of three faraday of charge through aqueous solution of $AgNO_3, CuSO_4, Al(NO_3)_3$ and NaCl will deposit metals at the cathode in the molar ratio of:
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°C
$Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$
at 25 °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) $2Cl^- \longrightarrow Cl_2 + 2e$

b) $2H^+ + 2e \longrightarrow H_2$

c) $2OH^- \longrightarrow H_2 + O_2 + 2e$

d) $Na^+ + e \longrightarrow Na$

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 Å 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) $Zn^{2+} + 2e \longrightarrow Zn(s)$ b) $Zn(s) \longrightarrow Zn^{2+} + 2e$ c) $Mn^{2+} + 2e \longrightarrow Mn(s)$ d) $Mn(s) \longrightarrow Mn^{2+} + 2e$

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 1atm 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, Zn| $ZnSO_4$ || $CuSO_4$ | 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

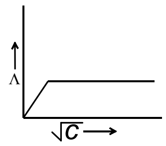
c) Equivalent conductivity

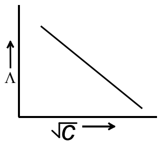
d) Molecular conductivity

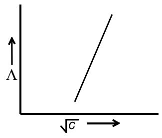
577. The limiting molar conductivities $\Lambda°$ for NaCl,KBr and KCl are 126,152 and 150 S $cm^2 mol^{-1}$ respectively. The $\Lambda°$ for NaBr is

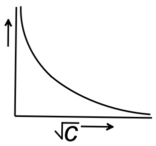
a) 128 S cm^2mol^{-1} b) 176 S cm^2mol^{-1} c) 278 S cm^2mol^{-1} d) 302 S cm^2mol^{-1}

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 oxugen 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 $Fe|Fe^{2+}$ is $+\,0.44$ V and standard $E°$ of half cell $Cu\,|\,Cu^{2+}$ 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) 3.14×10^{-3} mho cm^{-1}
   b) 3.14×10^{-3} mho^{-1} cm
   c) 3.14 mho cm^{-1}
   d) 3.14 mho^{-1} 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 mho cm^2 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) $K[Au(CN)_2]$ 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 Pt | $H_2(p_1)$ |$H^+(aq)$|| $H_2(p_2)$ | 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 5×10^{-3} 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 $H_2O(l), CO_2(g)$ 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 $l/a = 0.5 cm^{-1}, R = 50$ ohm, $N = 1.0$. The equivalent conductance of the electrolytic cell is

a) $10 \Omega^{-1}$ cm^2 g $equiv^{-1}$ b) $20 \Omega^{-1}$ cm^2 g $equiv^{-1}$

c) $300 \Omega^{-1}$ cm^2 g $equiv^{-1}$ d) $100 \Omega^{-1}$ cm^2 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 $Mg^{2+} + 2e \longrightarrow Mg(s)$; $E = -2.37$ V,

$Cu^{2+} + 2e \longrightarrow Cu(s)$; $E = +0.34$ V?

The e.m.f. of the cell Mg | Mg^{2+}||Cu^{2+}|Cu is :

a) 2.71 V b) 2.30 V c) 2.80 V d) 1.46 V

605. The standard reduction potentials of

Zn^{2+} | Zn , Cu^{2+} | Cu and Ag^+ | Ag are respectively – 0.76, 0.34 and 0.8 V. The following cells were constructed

I Zn | Zn^{2+} || Cu^{2+} | Cu

II Zn | Zn^{2+} || Ag^+ | Ag

III Cu | Cu^{2+} || Ag^+ | 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

$Zn \rightleftharpoons Zn^{2+} + 2e^-, E° =\oplus 0.76$ V
$Fe \rightleftharpoons Fe^{2+} + 2e^-, E° = +0.41$ 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, $H_2(g)$1 atm| H^+ (1 M)|| Cu^{2+}(1 M)| 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,
$Zn(s) + Ni^{2+}(1M) \longrightarrow Zn^{2+}(1M) + Ni(s)$ 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
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) $\Delta G° = -nFE^{\circ}_{cell}$ b) $\Delta G° = +nFE^{\circ}_{cell}$
c) $\Delta G° = -2.303\, RT\ nFE^{\circ}_{cell}$ d) $\Delta G° = -nF \log K_c$

616. For which cell e.m.f. is independent of the concentration of electrolytes used?
a) Fe | FeO(*s*)|KOH(*aq*)| N b) $Pt(H_2)$| HCl |$Pt(Cl_2)$ c) Zn|$Zn(NO_3)_2$|| $CuSO_4$ |C d) Hg, $HgCl_2$| KCl ||$AgNO_3$|

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) $Hg/Hg_2Cl_2 - KCl$ 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 $CuSO_4$ is
a) Fe b) He c) Na d) Ag

623. If the half-cell reaction $A + e \longrightarrow A^-$ 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 $AgNO_3$ 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 cm^3 b) 1.008 g c) 112 cm^3 d) 22400 cm^3

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 > Mg > Zn > Cu b) Mg > Zn > Cu > K c) K > Zn > Mg > Cu d) Cu > Zn > Mg > K

627. Calculate using appropriate molar conductance of the CH_3COOH from the molar conductances of electrolytes listed below at infinite dilution in H_2O at 25^0C :

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/Fe^{2+} and Cu/Cu^{2+} are −0.44 V and +0.32 V respectively, then

a) Cu^{2+} oxidises Fe b) Cu oxidises $FeFe^{2+}$ c) Cu reduces Fe^{2+} d) Cu^{2+} oxidises Fe^{2+}

629. The same amount of electricity was passed through two cells containing molten Al_2O_3 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

: ANSWER KEY :

1)	c	2)	b	3)	a	4)	d	189)	b	190)	c	191)	a	192)	d
5)	a	6)	d	7)	d	8)	a	193)	a	194)	a	195)	d	196)	a
9)	a	10)	c	11)	a	12)	a	197)	a	198)	d	199)	a	200)	a
13)	a	14)	d	15)	b	16)	c	201)	c	202)	b	203)	b	204)	c
17)	a	18)	c	19)	a	20)	a	205)	c	206)	d	207)	a	208)	a
21)	a	22)	b	23)	b	24)	a	209)	c	210)	d	211)	a	212)	a
25)	a	26)	b	27)	d	28)	d	213)	d	214)	d	215)	b	216)	b
29)	d	30)	c	31)	b	32)	d	217)	d	218)	b	219)	d	220)	a
33)	b	34)	d	35)	d	36)	a	221)	c	222)	c	223)	a	224)	c
37)	c	38)	a	39)	b	40)	a	225)	b	226)	a	227)	c	228)	a
41)	d	42)	c	43)	d	44)	c	229)	a	230)	a	231)	d	232)	a
45)	a	46)	b	47)	c	48)	a	233)	d	234)	a	235)	b	236)	a
49)	b	50)	d	51)	a	52)	a	237)	c	238)	b	239)	a	240)	a
53)	a	54)	d	55)	a	56)	a	241)	d	242)	c	243)	c	244)	a
57)	d	58)	d	59)	a	60)	c	245)	a	246)	d	247)	a	248)	c
61)	c	62)	b	63)	d	64)	b	249)	b	250)	b	251)	a	252)	c
65)	d	66)	d	67)	a	68)	a	253)	b	254)	d	255)	d	256)	c
69)	a	70)	d	71)	a	72)	c	257)	a	258)	d	259)	c	260)	d
73)	c	74)	d	75)	a	76)	d	261)	b	262)	c	263)	a	264)	c
77)	c	78)	d	79)	b	80)	b	265)	c	266)	a	267)	c	268)	a
81)	a	82)	a	83)	d	84)	c	269)	d	270)	c	271)	b	272)	a
85)	a	86)	b	87)	c	88)	b	273)	a	274)	c	275)	b	276)	a
89)	b	90)	a	91)	a	92)	b	277)	c	278)	b	279)	a	280)	b
93)	c	94)	b	95)	a	96)	c	281)	c	282)	b	283)	d	284)	c
97)	c	98)	c	99)	c	100)	c	285)	d	286)	d	287)	b	288)	a
101)	d	102)	b	103)	a	104)	a	289)	a	290)	c	291)	a	292)	a
105)	d	106)	c	107)	c	108)	b	293)	d	294)	c	295)	c	296)	c
109)	d	110)	a	111)	b	112)	c	297)	a	298)	b	299)	d	300)	a
113)	a	114)	d	115)	b	116)	b	301)	b	302)	c	303)	c	304)	d
117)	c	118)	b	119)	b	120)	d	305)	b	306)	c	307)	b	308)	d
121)	b	122)	c	123)	a	124)	c	309)	d	310)	b	311)	a	312)	b
125)	b	126)	d	127)	b	128)	a	313)	c	314)	d	315)	c	316)	d
129)	b	130)	c	131)	a	132)	d	317)	a	318)	a	319)	a	320)	a
133)	d	134)	d	135)	c	136)	b	321)	c	322)	d	323)	c	324)	a
137)	c	138)	a	139)	a	140)	d	325)	a	326)	d	327)	c	328)	a
141)	c	142)	b	143)	b	144)	b	329)	b	330)	a	331)	d	332)	a
145)	d	146)	c	147)	a	148)	a	333)	a	334)	a	335)	a	336)	a
149)	d	150)	d	151)	a	152)	b	337)	a	338)	d	339)	d	340)	d
153)	b	154)	c	155)	b	156)	a	341)	c	342)	b	343)	a	344)	b
157)	c	158)	d	159)	b	160)	b	345)	c	346)	a	347)	a	348)	a
161)	a	162)	b	163)	a	164)	b	349)	c	350)	a	351)	d	352)	b
165)	b	166)	b	167)	b	168)	d	353)	a	354)	a	355)	b	356)	a
169)	c	170)	b	171)	b	172)	b	357)	c	358)	a	359)	b	360)	d
173)	a	174)	b	175)	d	176)	c	361)	b	362)	a	363)	c	364)	b
177)	b	178)	b	179)	c	180)	a	365)	d	366)	b	367)	b	368)	a
181)	c	182)	a	183)	d	184)	b	369)	a	370)	c	371)	a	372)	c
185)	d	186)	b	187)	b	188)	b	373)	b	374)	b	375)	d	376)	a

377)	**d**	**378)**	**b**	**379)**	**d**	**380)**	**d**
381)	**c**	**382)**	**b**	**383)**	**d**	**384)**	**a**
385)	**a**	**386)**	**b**	**387)**	**b**	**388)**	**d**
389)	**d**	**390)**	**a**	**391)**	**b**	**392)**	**c**
393)	**c**	**394)**	**a**	**395)**	**a**	**396)**	**a**
397)	**c**	**398)**	**d**	**399)**	**a**	**400)**	**c**
401)	**c**	**402)**	**c**	**403)**	**a**	**404)**	**a**
405)	**d**	**406)**	**b**	**407)**	**b**	**408)**	**c**
409)	**c**	**410)**	**c**	**411)**	**d**	**412)**	**a**
413)	**d**	**414)**	**a**	**415)**	**d**	**416)**	**c**
417)	**a**	**418)**	**d**	**419)**	**a**	**420)**	**d**
421)	**a**	**422)**	**c**	**423)**	**c**	**424)**	**d**
425)	**d**	**426)**	**c**	**427)**	**a**	**428)**	**a**
429)	**b**	**430)**	**c**	**431)**	**c**	**432)**	**a**
433)	**b**	**434)**	**b**	**435)**	**c**	**436)**	**c**
437)	**c**	**438)**	**c**	**439)**	**b**	**440)**	**a**
441)	**d**	**442)**	**b**	**443)**	**a**	**444)**	**b**
445)	**a**	**446)**	**b**	**447)**	**b**	**448)**	**d**
449)	**d**	**450)**	**c**	**451)**	**c**	**452)**	**b**
453)	**c**	**454)**	**b**	**455)**	**b**	**456)**	**b**
457)	**a**	**458)**	**c**	**459)**	**b**	**460)**	**d**
461)	**c**	**462)**	**b**	**463)**	**b**	**464)**	**b**
465)	**c**	**466)**	**a**	**467)**	**b**	**468)**	**a**
469)	**c**	**470)**	**a**	**471)**	**a**	**472)**	**c**
473)	**a**	**474)**	**c**	**475)**	**c**	**476)**	**d**
477)	**b**	**478)**	**a**	**479)**	**a**	**480)**	**d**
481)	**b**	**482)**	**a**	**483)**	**d**	**484)**	**c**
485)	**d**	**486)**	**c**	**487)**	**a**	**488)**	**c**
489)	**d**	**490)**	**c**	**491)**	**d**	**492)**	**d**
493)	**c**	**494)**	**d**	**495)**	**c**	**496)**	**a**
497)	**c**	**498)**	**c**	**499)**	**a**	**500)**	**d**
501)	**b**	**502)**	**c**	**503)**	**a**	**504)**	**c**
505)	**a**	**506)**	**a**	**507)**	**a**	**508)**	**d**
509)	**a**	**510)**	**c**	**511)**	**c**	**512)**	**d**
513)	**b**	**514)**	**a**	**515)**	**b**	**516)**	**a**
517)	**d**	**518)**	**a**	**519)**	**a**	**520)**	**b**
521)	**b**	**522)**	**a**	**523)**	**a**	**524)**	**a**
525)	**a**	**526)**	**a**	**527)**	**b**	**528)**	**a**
529)	**c**	**530)**	**a**	**531)**	**a**	**532)**	**c**
533)	**c**	**534)**	**b**	**535)**	**a**	**536)**	**d**
537)	**b**	**538)**	**c**	**539)**	**b**	**540)**	**d**
541)	**a**	**542)**	**c**	**543)**	**d**	**544)**	**c**
545)	**b**	**546)**	**b**	**547)**	**a**	**548)**	**b**
549)	**d**	**550)**	**d**	**551)**	**c**	**552)**	**b**
553)	**c**	**554)**	**a**	**555)**	**c**	**556)**	**b**
557)	**a**	**558)**	**d**	**559)**	**b**	**560)**	**d**
561)	**a**	**562)**	**b**	**563)**	**a**	**564)**	**d**
565)	**d**	**566)**	**b**	**567)**	**a**	**568)**	**d**
569)	**c**	**570)**	**a**	**571)**	**c**	**572)**	**a**
573)	**d**	**574)**	**b**	**575)**	**a**	**576)**	**b**
577)	**a**	**578)**	**d**	**579)**	**a**	**580)**	**a**

581)	**c**	**582)**	**b**	**583)**	**a**	**584)**	**b**
585)	**b**	**586)**	**a**	**587)**	**b**	**588)**	**a**
589)	**d**	**590)**	**d**	**591)**	**c**	**592)**	**c**
593)	**d**	**594)**	**a**	**595)**	**b**	**596)**	**c**
597)	**b**	**598)**	**c**	**599)**	**c**	**600)**	**b**
601)	**c**	**602)**	**a**	**603)**	**d**	**604)**	**a**
605)	**b**	**606)**	**c**	**607)**	**b**	**608)**	**c**
609)	**a**	**610)**	**b**	**611)**	**b**	**612)**	**d**
613)	**a**	**614)**	**a**	**615)**	**a**	**616)**	**a**
617)	**b**	**618)**	**d**	**619)**	**c**	**620)**	**a**
621)	**d**	**622)**	**a**	**623)**	**d**	**624)**	**c**
625)	**d**	**626)**	**a**	**627)**	**c**	**628)**	**a**
629)	**a**	**630)**	**a**	**631)**	**d**		

: HINTS AND SOLUTIONS :

1 **(c)**

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

The charge required = 3 × 96500 C

2 **(b)**

Eq. of H_2 = Eq. of Cu

$\therefore \frac{0.504}{1} = \frac{W}{63.5/2}$

$\therefore W_{Cu} = 16$ 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° will be − ve and $K_{eq} < 1$; one should not write $\Delta G^\circ > 0$.

4 **(d)**

$E^\circ_{\text{cell}} = 0.87 + 0.40 = 1.27$ V

Cell reaction : $Fe + Ni_2O_3 \rightarrow FeO + 2NiO$

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)**

$$E^\circ_{\text{cell}} = E^\circ_{OP_{\text{Sn}}} + E^\circ_{RP_{\text{Fe}}} = 0.14 + (-0.44)$$
$$= -0.30 \text{ V}$$

8 **(a)**

Smallest ion possesses maximum mobility.

9 **(a)**

$\Lambda^0_M = \Lambda^0_a + \Lambda^0_c$

10 **(c)**

Molten NaCl possesses Na^+ and Cl^- ions.

11 **(a)**

Given, that

Zn / Zn^{2+} || Cu^{2+} / Cu

$\therefore$ Zn is anode and Cu is cathode.

Given,

Zn^{2+}/ Zn = − 0.76 V

Cu^{2+} / Cu = + 0.34 V

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

= 0.34 – (- 0.76)

= 0.34 + 0.76

= 1.10 V

12 **(a)**

Net redox change is zero.

13 **(a)**

Cathode $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$

Anode : $H_2O \rightarrow 2H^+ + \frac{1}{2}O_2 + 2e^-$

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)**

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)**

$\triangle G = \triangle H - T \triangle S$

For a spontaneous cell reaction, $\triangle H$ should be negative and $\triangle S$ should be positive. Hence, $\triangle G$ should be negative.

19 **(a)**

Cell reaction is $Mg + Sn^{2+} \rightarrow Mg^{2+} + Sn$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{2}\log\frac{[Mg^{2+}]}{[Sn^{2+}]}$$

$$= (2.34 - 0.14) - \frac{0.0591}{2}\log\frac{10^{-2}}{10^{-1}} = 2.23 \text{ V}$$

20 **(a)**

pH of solution increases due to formation of LiOH or due to increase in $[OH^-]$ 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)

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

Cell constant = $\kappa \times R$

= 0.0129 × 100 = 1.29

23 **(b)**

According to Nernst equation.

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

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

Or $y = c + (-m)x$

Thus, the slope is negative.

24 **(a)**

In MnO_4^- the oxidation number of Mn is + 7.

$$\therefore \overset{+7}{Mn} + 5e^- \longrightarrow \overset{+2}{Mn}$$

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 : $Cl^- \longrightarrow (1/2)Cl_2 + e$

At cathode : $H^+ \longrightarrow (1/2)H_2 + e$

31 **(b)**

Electrons flow from Zn to Cu in outside circuit and current from Cu to Zn.

32 **(d)**

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

Ni / Ni^{2+} [1.0 M] || Au^{3+} [1.0 M] | Au

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

E_{cell} (Ni^{2+} / Ni) = − 0.25 V

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$$
$$= 0.150 - (-0.25)$$
$$= 0.15 + 0.25$$
$$= +0.4 \text{ V}$$

33 **(b)**

50 % H_2SO_4aqueous solution can be electrolysed by using Pt electrodes as

$2H_2SO_4 \rightarrow 2HSO_4^- + 2H^+$

$2HSO_4^- \rightarrow H_2S_2O_8 + 2e^-$ (at anode)

34 **(d)**

It is fact.

35 **(d)**

For the given cell,

$$E_{cell} = E^\circ_{cell} - \frac{0.0591}{2}\log\frac{[Zn^{2+}]}{[Cu^{2+}]}$$

1. $$E_1 = E^\circ_{cell} - \frac{0.0591}{2}\log\frac{1}{0.1}$$
$$= E^\circ_{cell} - \frac{0.0591}{2}$$

2. $$E_2 = E^\circ_{cell} - \frac{0.0591}{2}\log\frac{1}{1}$$
$$= E^\circ_{cell} - \frac{0.0591}{2} \times 0$$
$$= E^\circ_{cell}$$

3. $$E_2 = E^\circ_{cell} - \frac{0.0591}{2}\log\frac{0.1}{1}$$
$$= E^\circ_{cell} + \frac{0.0591}{2}$$

$\therefore E_3 > E_2 > E_1$

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) $Fe(s) \rightarrow Fe^{2+} + 2e^-$, $E^\circ = +0.44$ V

and $\Delta G_1^\circ = -nE^\circ F = -2 \times 0.44 \times F$

(ii) $2H^+ + 2e^- + \frac{1}{2}O_2 \longrightarrow H_2O(l)$; $E^\circ = +1.23\ V$

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

Net reaction,

$Fe(s) + 2H^+ + \frac{1}{2}O_2 \rightarrow Fe^{2+} + H_2O(l)$

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

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)**

$E^\circ_{cell} = \frac{2.303RT}{nF} \log K_{eq}$

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

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

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

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)**

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

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

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

45 **(a)**

High value for E°_{red}. Shows more electronegativity $i.e.$, Zn is more electropositive than Fe.

$(E^\circ_{\mathrm{Zn^{2+}/Zn}} < E^\circ_{\mathrm{Fe^{2+}/Fe}})$

46 **(b)**

Eq. of Cu = Eq. of Ag

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

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

47 **(c)**

The cell reaction is

$$\mathrm{H_2(g) + I_2}(s) \rightleftharpoons 2\mathrm{H^+}(aq) + 2\mathrm{I^-}(aq)$$

$$0.7714 = 0.535 - \frac{0.0591}{2} \log \frac{[\mathrm{H^+}]^2[\mathrm{I^-}]^2}{\mathrm{p_{H_2}}}$$

$\therefore$ pH = 3

48 **(a)**

$E^\circ_{\mathrm{Cr^{3+}/Cr^{2+}}} = -0.41$ V

$E^\circ_{\mathrm{Mn^{3+}/Mn^{2+}}} = +1.57$ V

$E^\circ_{\mathrm{Fe^{3+}/Fe^{2+}}} = +0.77$ V

$E^\circ_{\mathrm{Co^{3+}/Co^{2+}}} = +1.97$ V

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

49 **(b)**

In other cells, two liquid are not present.

50 **(d)**

$\mathrm{AgNO_3} \overset{\Delta}{\rightarrow} \mathrm{Ag_2O} \overset{\Delta}{\rightarrow} \mathrm{Ag} + \mathrm{O_2}$

51 **(a)**

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

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

In the cell,

$$\underset{\text{anode}}{\mathrm{Cu \mid Cu^{2+}}} \; || \; \underset{\text{cathode}}{\mathrm{Zn^{2+} \mid Zn}}$$

In the cell ,

$E^\circ_{\mathrm{cell}} = E^\circ_{\mathrm{cathode}} - E^\circ_{\mathrm{anode}}$

$= 0.76 - (-0.34)$

$= 1.10$ V

52 **(a)**

Here Fe acts as anode while Sn act as cathode. We know that,

$E^\circ_{\mathrm{cell}} = E^\circ_{\mathrm{cathode}} - E^\circ_{\mathrm{anode}}$

$= (-0.14) - (-0.44)$

$= -0.14 + 0.44$

$= 0.30$ V

54 **(d)**

$\mathrm{Ag^+} + e^- \longrightarrow \mathrm{Ag}$

$\because$ 96500 C are required to deposite Ag = 108 g

$\therefore$ 965 C are required to deposite Ag

$= \frac{108}{96500} \times 965 = 1.08$ g

55 **(a)**

$\Lambda_{\mathrm{m}} = \Lambda_{eq} \times$ 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.

$\mathrm{Fe + Na_3PO_4} \rightarrow$ No reaction

57 **(d)**

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

58 **(d)**

Galvanic cell is

Cu (s) | $\mathrm{Cu^{2+}}$ (aq) || $\mathrm{Hg^{2+}}$ (aq) | Hg (l)

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

Cu (s) + $\mathrm{Hg^{2+}}$ (aq) $\rightarrow$ $\mathrm{Cu^{2+}}$ (aq) + Hg (l)

59 **(a)**

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

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

$= 0.0836$ litre

60 **(c)**

$$E^\circ_{\text{cell}} = \frac{2.303RT}{nF}\log K_{eq}$$

$$E^\circ_{\text{cell}} = \frac{0.0591}{n}\log K_{eq} \qquad \text{[At 298 K]}$$

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

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

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

61 **(c)**

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

$$\Delta G = -nFE^\circ_{\text{cell}}$$

or $= -nFE$

62 **(b)**

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

Given,

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

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

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

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

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

$$= -2.36 - 0.0591$$

$$= -2.419\ V$$

63 **(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)**

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

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

$$\therefore \ E^\circ_{\text{cell}} = 0.32$$

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

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

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

67 **(a)**

$\text{Ni}^- + 2e^- \rightarrow \text{Ni}$ (at cathode)

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

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

$$= 29.35$$

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

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

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

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

$= 7.883$ g

68 **(a)**

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

$\therefore \frac{W}{E} = 10^{-2}$ (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

$z > y > x$

Thus, $y + e^- \rightarrow y^-$

$x \rightarrow x^- + e^-$

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

($\kappa = C \times$ cell constant).

71 **(a)**

Given, $i = 2.5A$

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

Number of coulomb passed $= i \times t$

$= 2.5 \times 386$

$= 965$ C

$$\text{Cu}^{2+} + 2e^- \longrightarrow \text{Cu}$$

$\therefore$ 2×96500 C charge deposits Cu $= 63.5$ g

$\therefore$ 965 C charge deposits

Cu $= \frac{63.5}{2 \times 96500} \times 965$

$= 0.3175$ g

72 **(c)**

Metal placed above in electrochemical series replaces the other from its salt solutions.

73 **(c)**

$$E_{\text{cell}} = E^\circ_{OP_{\text{Zn}}} + E^\circ_{RP_{\text{Cu}}} + \frac{0.059}{2}\log\frac{[\text{Cu}^{2+}]}{[\text{Zn}^{2+}]}$$

$$\therefore \ 1.1 = 0.78 + E^\circ_{RP_{\text{Cu}}} + \frac{0.059}{2}1$$

$$\therefore \ E^\circ_{RP_{\text{Cu}^{2+}/\text{Cu}}} = 0.32$$

$$\therefore \ E^\circ_{RP_{\text{Cu}^{2+}/\text{Cu}}} = -0.32\text{V}$$

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

$$\mathrm{Co} + 2\mathrm{Ag}^+ \rightarrow \mathrm{Co}^{2+} + 2\mathrm{Ag}$$

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

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

77 **(c)**

Electrolysis of water takes place as follows

$$H_2O \rightleftharpoons \underset{\text{Cathode}}{H^+} + \underset{\text{anode}}{OH^-}$$

At anod

$$OH^- \xrightarrow{\text{oxidation}} OH + e^-$$

$$4OH \longrightarrow 2H_2O + O_2$$

At cathode

$$2H^+ + 2e^- \xrightarrow{\text{Reduction}} H_2$$

Given, time, $t = 1930s$

Number of moles of hydrogen collected

$$= \frac{1120 \times 10^{-3}}{22.4} \text{ moles}$$

$$= 0.05 \text{ moles}$$

∵ 1 mole of hydrogen is deposited by = 2 moles of electrons

∵ 0.05 moles of hydrogen will be deposited by

$$= 2 \times 0.05$$

$$= 0.10 \text{ mole of electrons}$$

Charge, $Q = nF$

$$= 0.1 \times 96500$$

Charge, $Q = it$

$$0.1 \times 96500 = i \times 1930$$

$$i = \frac{0.1 \times 96500}{1930}$$

$$= 5.0 \text{ A}$$

78 **(d)**

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

5. $E^\circ_{cell} = \frac{2.303\,RT}{nF}\log K_c$

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

79 **(b)**

$Cr_2O_7^{2-} + 6e^- \longrightarrow 2Cr^{3+}$

Reduction of 1 mol of $Cr_2O_7^{2-}$ to Cr^{3+} required 6 moles of electrons. Hence, charge required = 2 × 96500 C

80 **(b)**

Cell constant $= \frac{1}{a} = \frac{\text{length}}{\text{area}}$

∴ unit is cm^{-1}.

81 **(a)**

$$\lambda = k \times V = \frac{1}{R} \times \frac{1}{a} \times V$$

$$= \frac{1}{210} \times 0.66 \times 100000$$

$$= 314.28 \text{ mho cm}^2 \text{ eq.}^{-1}$$

82 **(a)**

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

$Cu + 2HCl \rightarrow CuCl_2 + H_2(g)$

$E^\circ_{\text{cell}} = E^\circ_{H^+/H_2} - E^\circ_{Cu^{2+}/Cu}$

$= 0.00 - (+0.34) = -0.34$ V

$Zn + 2HCl \rightarrow ZnCl_2 + H_2(g)$

$E^\circ_{\text{cell}} = E^\circ_{H^+/H_2} - E^\circ_{Zn^{2+}/Zn}$

$= 0.00 - (-0.76) = +0.76$ V

$Ag + 2HCl \rightarrow AgCl + \frac{1}{2}H_2(g)$

$E^\circ_{\text{cell}} = E^\circ_{H^+/H_2} - E^\circ_{Ag^{2+}/Ag}$

$= 0.00 - (0.80) = -0.80$ V

Hence, only reaction (ii) is feasible.

83 **(d)**

$Zn + MgCl_2 \rightarrow$ no reaction

This type of reaction does not occur because

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

$\Lambda^\infty_{ClCH_2COOH} = \Lambda^\infty_{ClCH_2COO^-} + \Lambda^\infty_{H^+}$

Given from Kohlrausch law,

$\Lambda^\infty_{ClCH_2COONa} = 224\Omega^{-1}\text{ cm}^2\text{ g eq}^{-1}$

$\Lambda^\infty_{ClCH_2COONa} = \Lambda^\infty_{ClCH_2COO^-} + \Lambda^\infty_{Na^+}$ ...(i)

$\Lambda^{\infty}_{\mathrm{HCl}} = 203\Omega^{-1}\ \mathrm{cm}^2\ \mathrm{g\ eq}^{-1}$

$\Lambda^{\infty}_{\mathrm{HCl}} = \Lambda^{\infty}_{\mathrm{H}} + \Lambda^{\infty}_{\mathrm{Cl}}$ ...(ii)

$\Lambda^{\infty}_{\mathrm{NaCl}} = 38.5\ \Omega^{-1}\ \mathrm{cm}^2\ \mathrm{g\ eq}^{-1}$

$\Lambda^{\infty}_{\mathrm{NaCl}} = \Lambda^{\infty}_{\mathrm{Na^+}} + \Lambda_{\mathrm{Cl}}$ ....(iii)

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

$\Lambda^{\infty}_{\mathrm{ClCH_2COO^-}} + \Lambda^{\infty}_{\mathrm{H^+}}$

$= \Lambda^{\infty}_{\mathrm{ClCH_2COONa}} + \Lambda^{\infty}_{\mathrm{HCl}} - \Lambda^{\infty}_{\mathrm{NaCl}}$

$= 224 + 203 - 38.5$

$= 427 - 38.5$

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

88 **(b)**

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

89 **(b)**

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

90 **(a)**

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

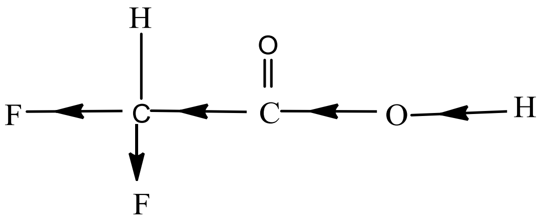
Also, $\alpha = \frac{\Lambda_v}{\Lambda^{\infty}}$

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

$= 172.8$

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.

$[\mathrm{Ag} \longrightarrow \mathrm{Ag}^+ + e$ (anode); $\mathrm{Ag}^+ + e \longrightarrow \mathrm{Ag}$(cathode)]

Thus, concentration of salt does not change.

94 **(b)**

Equivalent conductivity $(\Lambda_{\mathrm{eq}}) = \frac{\kappa \times 1000}{C}$

Conductivity $(\kappa) = \frac{\text{cell costant}}{\text{resistance}}$

$= \frac{1.15}{250}\ \mathrm{S\ cm}^{-1}$

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

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

95 **(a)**

Given, $E^{\circ}_{\mathrm{Zn/Zn^{2+}}} = 0.76\ \mathrm{V}$

$E^{\circ}_{\mathrm{Cu/Cu^{2+}}} = 0.34\ \mathrm{V}$

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

$\therefore\ E^{\circ}_{\mathrm{Zn^{2+}/Zn}} = -0.76\ V$

and $E^{\circ}_{\mathrm{Cu^{2+}/Cu}} = -0.34\ V$

$E^{\circ}_{\mathrm{cell}} = E^{\circ}_{\mathrm{cathode}} - E^{\circ}_{\mathrm{anode}}$

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

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

$= 0.42\ \mathrm{V}$

96 **(c)**

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

97 **(c)**

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

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

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

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, $\mathrm{Cl}^+ + 2e \longrightarrow \mathrm{Cl}^-$($i.e.$, reduction of OCl^-)

100 **(c)**

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

101 **(d)**

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

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

102 **(b)**

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

103 **(a)**

$2Al + dil.H_2SO_4 \rightarrow Al_2SO_4 + H_2 \uparrow$

104 **(a)**

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

Hence,

$E^\circ_{cell} = E^\circ_{cathode}$ - E°_{anode}

$= (2.23) - (-1.43)$

$= 2.23 + 1.43 = 3.66$ V

105 **(d)**

H_2SO_4 will furnish maximum H^+.

106 **(c)**

Hg is placed below H in electrochemical series.

107 **(c)**

$\text{Eq. of } H_2 = \text{Eq. of Al} = \frac{4.5}{27/3} = 0.5$

$\because$ 1 eq. $H_2 = 11.2$ L

$\therefore$ 0.5 eq. $H_2 = 5.6$ L

108 **(b)**

No. of moles of $H_2 = \frac{11.2}{22400}$

No .of equivalence of hydrogen

$= \frac{1.12 \times 2}{22400} = 10^{-4}$

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

$\therefore$ Current to be passed in one second

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

$= 9.65$ 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)**

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

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

113 **(a)**

E°_{op} of K $> E^\circ_{Op}$ 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)**

$Mn^{7+} + 5e \longrightarrow Mn^{2+}$;

Thus, 5 mole electron = 5 faraday.

116 **(b)**

$E_{cell} = E^\circ_{cell} + \frac{0.0592}{n}\log\frac{[Cu^{2+}]}{[T1^+]^2}$

$n = 2$

$\therefore$ According to above equation E_{cell} can be increased by increasing [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

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

$2Cl^- \rightarrow Cl_2 + 2e^-$

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)**

$E_{cell} = E_{OP_{anode}} + E_{RP_{cathode}}$

$= E_{OP_{anode}} - E_{OP_{cathode}}$

121 **(b)**

Electrode potential of cell must be positive for spontaneous reaction.

$Zn^{2+} \longrightarrow Zn$; $E^\circ = -0.76\ V$

$Cu^{2+} \longrightarrow Cu$; $E^\circ = -0.34\ V$

Redox reaction is

$$\begin{array}{ll} Zn \rightarrow Zn^{2+} + 2e^- & \text{(oxidation)} \\ Cu^{2+} + 2e^- \rightarrow Cu & \text{(reduction)} \\ \hline Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu & \\ \hline \end{array}$$

$E_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$= -0.34 - (-0.76)$

$= +0.42$ 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)**

Cr/Cr^{3+} (0.1 M) || Fe^{2+} (0.01 M) | Fe

Oxidation half-cell; $Cr \rightarrow Cr^{3+} + 3e^- \times 2$

Reduction half-cell; $Fe^{2+} + 2e^- \rightarrow Fe \times 3$

Net cell reaction;

$2Cr + 2Fe^{2+} \rightarrow 2Cr^{3+} + 3Fe \quad (n = 6)$

$$E^\circ_{\text{cell}} = E^\circ_{\text{oxidation}} - E^\circ_{\text{reduction}}$$
$$= 0.72 - 0.42$$
$$= 0.30 \text{ V}$$

$$E^\circ_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{n}\log\frac{[Cr^{3+}]^2}{[Fe^{2+}]^3}$$
$$= 0.30 - \frac{0.0591}{6}\log\frac{(0.1)^2}{(0.01)^3}$$
$$= 0.30 - \frac{0.0591}{6}\log\frac{10^{-2}}{10^{-6}}$$
$$= 0.30 - \frac{0.0591}{6}\log 10^4$$
$$E_{\text{cell}} = 0.2606\, V$$

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.

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

126 **(d)**

The elements which are below H_2 in electrochemical series, cannot displaceH_2 .

$\because$Out of Li^+, Sr^{2+}, Al^{3+} and Ag^+, Ag^+ is below H_2 in electrochemical series,

so Ag^+ cannot displaceH_2.

127 **(b)**

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

Hence , the correct reaction will be

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$$

(oxidation)

$$\begin{array}{l} \quad 2Ag^+(aq) + 2e^- \rightarrow 2Ag(s) \\ \hline Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s) \end{array}$$

128 **(a)**

$W \propto i \times t$ 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. Reducing character $\propto \frac{1}{\text{reduction potentials}}$

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

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

132 **(d)**

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

133 **(d)**

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$
$$\therefore \quad 2.46 = (+0.80) - E^\circ_{Al^{3+}/Al}$$

Or $E^\circ_{Al^{3+}/Al} = 0.80 - 2.46 = -1.66$ V

134 **(d)**

E° for reaction in (d) $= E^\circ_{OP_{\text{Br}}} + E^\circ_{RP_1} = -1.09 + (-0.54)$

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

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.

$$K_4[Fe(CN)_6] \longrightarrow 4K^+ + [Fe(CN)_6]^{4-}$$

139 **(a)**

The metals having higher negative value of standard reduction potential are placed above hydrogen in electrochemical series. The metals places 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)**

$$w = 60 \text{ g}$$
$$i = 5\text{A}$$

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

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

According to first law of Faraday electrolysis

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

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

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

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

= 16.08 h

142 **(b)**

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

At anode $\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$ (oxidation)

At cathode $\text{Cu}^{2+} + 2e^- \longrightarrow \text{Cu}$ (reduction)

Cell reaction $\text{Zn} + \text{Cu}^{2+} \longrightarrow \text{Zn}^{2+} + \text{Cu}$

Or $\text{Zn(s)} + \text{CuSO}_4\text{(aq)} \rightarrow \text{Cu(s)} + \text{ZnSO}_4\text{(aq)}$

143 **(b)**

$\Lambda^{\infty}_{\text{AcOH}} = \Lambda^{\infty}_{\text{AcONa}} + \Lambda^{\infty}_{\text{HCl}} - \Lambda^{\infty}_{\text{NaCl}}$

= 91.0 + 426.2 – 126.5

= 390.7

144 **(b)**

The metal with more E°_{OP} is oxidised.

145 **(d)**

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

For standard hydrogen electrode,

$E^{\circ}_{\text{cell}} = 0.00V$

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

Given, pH = 1.0

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

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

[$\because$ The reaction occurring is $2\text{H}^+ + 2e^- \rightarrow \text{H}_2$]

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

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

= - 0.05915 V

= - 59.15 mV

146 **(c)**

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

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

147 **(a)**

E°_{OP} of Mg $>$ E°_{OP} of Al.

148 **(a)**

For the given cell, reaction is

$\text{Zn} + \text{Fe}^{2+} \rightarrow \text{Zn}^{2+} + \text{Fe}$

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

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

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

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

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

$K_c = 10^{\frac{0.32}{0.295}}$

149 **(d)**

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

At anode

$\text{Pb} + \text{H}_2\text{SO}_4 \longrightarrow \text{PbSO}_4 + 2\text{H}^+ + 2e^-$

At cathode

$\text{PbO}_2 + 4\text{H}^+ + \text{SO}_4^{2-} + 2e^- \longrightarrow \text{PbSO}_4 + 2\text{H}_2\text{O}$

150 **(d)**

$2\text{H}^+ + 2e^- \rightarrow \text{H}_2$

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 pH

151 **(a)**

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

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

$E^{\circ}_{\text{cell}} = E^{\circ}_{OP_{\text{Fe/Fe}^{2+}}} + E^{\circ}_{RP_{\text{Fe}^{3+}/\text{Fe}^{2+}}}$ (See redox change)

= + 0.441 + 0.771 = 1.212 V

152 **(b)**

$E^{\circ}_{OP_{\text{Zn}}} > E^{\circ}_{OP_{\text{Cu}}}$ or $E^{\circ}_{RP_{\text{Zn}}} < E^{\circ}_{RP_{\text{Cu}}}$

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)**

$$\frac{1}{2}H_2| H^+ || Ag^+ | Ag$$

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$$

$$= E^\circ_{Ag^+/Ag} - E^\circ_{H^+/\frac{1}{2}H_2}$$

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

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)**
250mL of 1 M $AgNO_3$ contain $= \frac{250}{1000}$
$= 0.25$ mole $AgNO_3$
∵ Electricity required to liberate 1 g equivalent of metal
$= 96500$ C
∴ Electricity required to liberate 0.25 g equivalent of metal
$= \frac{96500 \times 0.25}{1}$
$= 24125$ C

160 **(b)**
1 faraday = 1 eq. of Cu = 1/2 mole Cu
$= N/2$ 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 withdil HCl.
Ag + dil. HCl → No reaction

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.
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
$= 0.5 \times 23$
$= 11.50$ g of sodium (E = 23)

165 **(b)**
Current $(i) = 1.5\,A$
Time (t) 10 min $= 10 \times 60 = 600$ s
Quantity of electricity passed $Q = i \times t$
$= (1.5\,A) \times (600\,s)$
$= 900\,C$
Copper is deposited as
$Cu^{2+} + 2e^- \longrightarrow Cu(s)$
2 moles of electrons or $2 \times 96500\,C$ of current deposit copper = 63.56 g

900 C of current will deposit copper

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

$= 0.296$ g

166 **(b)**
Ionic mobility depends upon the charge to size ratio of ion. The ionic size in case of hydrated cation is

$$K^+(aq) < Na^+(aq) < Li^+(aq)$$

167 **(b)**
Eq. of A = Eq. of B = Eq. of C
or $\frac{2.1}{7/n_1} = \frac{2.7}{27/n_2} = \frac{7.2}{48/n_3}$
$0.3\,n_1 = 0.1n_2 = 0.15\,n_3$
∴ $n_1 = \frac{n_2}{3} = \frac{n_3}{2}$
If $n_1 = 1$ then $n_2 = 3$, $n_3 = 2$

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})
∴ Molar conductance $(\Lambda_m) \propto (\frac{1}{C})$

170 **(b)**
Rusting of iron is catalyzed by moist air.

171 **(b)**
$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$
For 0.01 mole H_2, 0.02 mole of electrons are consumed charge required
$= 0.02 \times 96500\,C = i \times t$

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

172 **(b)**
Metal having higher E°_{OP} replaces the other from its solution.

173 **(a)**
Eq. of Ag = Eq. of H_2;
$$\frac{W}{108} = \frac{5600 \times 2}{22400 \times 1}$$
$\therefore \quad W_{Ag} = 54$ g

174 **(b)**
$Ag^+ + e^- \rightarrow Ag$
9650 C = 0.1 F = 0.1 equivalent Ag
= 0.1 mol Ag
= 10.8 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 $Cu^{2+} + 2e \longrightarrow Cu$;
In CuCN, change is $Cu^+ + e \longrightarrow Cu$;
Thus, $W \propto E_{Cu}$, which is more in CuCN.

179 **(c)**

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

180 **(a)**
$2H^- \longrightarrow H_2 + 2e$; 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".

$i.e., \quad w \propto Q$

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."

$i.e., \quad \frac{w_1}{w_2} = \frac{E_1}{E_2}$ 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)**
$$\frac{\text{Weight of Cu}}{\text{Weight of } H_2} = \frac{\text{Eq. wt. of Cu}}{\text{Eq. wt. of } H_2}$$
$$\frac{\text{Weight of Cu}}{0.504} = \frac{63.6/2}{1}$$
$\therefore$ 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)**
$E_{cell} = E_{OP_L} + E_{RP_R} = -E_{RP_L} + E_{RP_R}$.

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

192 **(d)**
$$\begin{array}{ll} AgI\,(s) + e^- \rightleftharpoons Ag\,(s) + I^- ; & E^\circ = 0.152 \text{ V} \\ Ag\,(s) \longrightarrow Ag^+ + e^- & E^\circ = -\,0.8 \text{ V} \\ \hline AgI\,(s) \longrightarrow Ag^+ + I^- & E^\circ = -\,0.952 \end{array}$$
$$E^\circ_{cell} = \frac{0.059}{n} \log K_{sp}$$
$$-0.952 = \frac{0.059}{1} \log K_{sp}$$
$$\log K_{sp} = \frac{-0.952}{0.059} = -16.135$$

193 **(a)**
$$\Lambda^\infty_{CH_3COOH} = \Lambda^\infty_{CH_3COONa} + \Lambda^\infty_{HCl} - \Lambda^\infty_{NaCl}$$

194 **(a)**

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

At anode : $Cu \longrightarrow Cu^{2+} + 2e$

195 **(d)**

$$\frac{\text{Wt. of Cu deposited}}{\text{Wt. of } H_2 \text{ produced}} = \frac{\text{eq. wt. of Cu}}{\text{eq. wt. of H}}$$

$$\frac{0.16}{\text{wt. of } H_2} = \frac{64/2}{1} = \frac{32}{1}$$

Wt. of $H_2 = \frac{0.16}{32} = 5 \times 10^{-3}$ g

Volume of H_2 liberated at STP

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

= 56 cc

196 **(a)**

Faraday's laws are independent of external factors.

197 **(a)**

E°_{OP} for Li is more,

$$Li \longrightarrow Li^{+} + e$$

Thus, Li is strong oxidant.

198 **(d)**

All are electrolytic cells.

199 **(a)**

96500 C or 1F will liberate 1 eq. of O_2 or 1/4 mole O_2 or 5.6 litre O_2 at NTP.

200 **(a)**

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

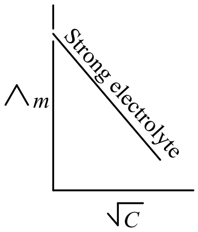
201 **(c)**

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

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

202 **(b)**

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



Variation of molar conductance $(\Lambda_m)$with$\sqrt{C}$ for strong electrolyte.

203 **(b)**

$\Lambda^{\infty}_{eq}(NH_4\ OH) = \Lambda^{\infty}_{eq}\ (NH_4Cl) + \Lambda^{\infty}_{eq}\ (NaOH) - \Lambda^{\infty}_{eq}\ (NaCl)$

$= (149.74 + 248.1 - 126.4)$

$= 271.44\ \Omega^{-1}\ cm^2 eq^{-1}$

204 **(c)**

$E^{\circ}_{cell} = E^{\circ}_{cathode} - E^{\circ}_{anode}$

$E^{\circ}_{Ag^{+}/Ag} - E^{\circ}_{Cu^{+}/Cu}$

= - 0.80 – 0.34

= + 0.46

205 **(c)**

$E_{cell} = E_{OP_{Ni/Ni_2^{2+}}} + E_{RP_{Au^{3+}/Au}}$

=

$E^{\circ}_{OP_{Ni}} - \frac{0.059}{2}\log[Ni^{2+}] + E^{\circ}_{RP_{Au}} + \frac{0.059}{3}\log[Au^{3+}]$

$= 0.25 - \frac{0.059}{2}\log(1.0) + 1.50 + \frac{0.059}{3}\log 1.0 = 1.75$V

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,

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

Zn is anode and Cu is cathode.

∴ Cell representation is

$Zn \mid Zn^{2+} \mid\mid Cu^{2+} \mid Cu$

209 **(c)**

2 × 96500 C electricity is used to liberate

= 22400 mL O_2 at STP

∴ 9.65 × 1000 C electricity will liberate

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

= 1120mL

210 **(d)**

Oxidation half-cell

$ClO_3^- \rightarrow ClO_2^- + 2e^-;\quad E^\circ_{cell} = -0.36\text{ V}$

Reduction half-cell

$ClO_3^- + 2e^- \rightarrow ClO_2^-;\ E^\circ_{cell} = 0.33\text{ V}$

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

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

$2ClO_3^- \rightleftharpoons ClO_4^- + ClO_2^-$

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

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

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 occur as follows

$$\overset{0}{Cu} + AgNO_3 \xrightarrow{\text{Oxidation}} \overset{+1}{CuNO_3} + Ag$$

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.

214 **(d)**

$$\begin{array}{ll} Fe^{2+} + 2e \longrightarrow Fe; & -\Delta G_1 = 2 \times (-0.44) \times F \\ \underset{-}{Fe^{3+}} + \underset{-}{3e} \longrightarrow \underset{-}{Fe}; & -\underset{+}{\Delta G_2} = 3 \underset{-}{\times} (-0.036) \times F \\ \hline \end{array}$$

$Fe^{2+} \longrightarrow Fe^{3+} + e;\ \Delta G_3 = -0.88 + 0.108 = 0.772$ or $0.772 = 1 \times E^\circ \times F$

$\therefore Fe^{3+} + e \longrightarrow Fe^{2+}; E^\circ = +0.772\text{ V}$

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)**

$Cu^{2+} + 2KI \longrightarrow CuI_2 + 2K^+;$

$2CuI_2 \longrightarrow Cu_2I_2 + I_2$

217 **(d)**

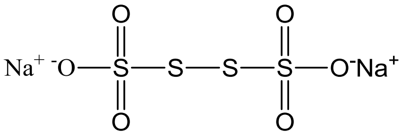
Salt bridge is used to remove or eliminate liquid junction potential arised due to different relative speed of ions of electrolytes at the junction of two electrolytes in an electrochemical cell. Thus, a salt bridge such as KCl is placed in between two electrolytes. A salt used for this purpose should have almost same speeds of its cation and anion.

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)**

$Na_2S_4O_6$ is



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 ofE°, more is the reducing power.

The order of E° values (negative value) is

$$\underset{(\text{Mg})}{-2.37} > \underset{(\text{Zn})}{-0.76} > \underset{(\text{Fe})}{-0.44}$$

$\therefore$ Mg can reduce both Zn^{2+} andFe^{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)**

$Sn^{2+} \rightarrow Sn^{4+} + 2e^-$

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

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

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

224 **(c)**

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

225 **(b)**

Reduction always occurs at cathode during electrolysis.

226 **(a)**

$E^\circ_{\text{cell}} = \frac{0.0591}{2} \log K_c$

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

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

$\therefore \log K_c = 20$

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

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 g of silver.

228 **(a)**

$CaF_2 = k_{\text{solution}} - k_{\text{water}}$

229 **(a)**

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

230 **(a)**

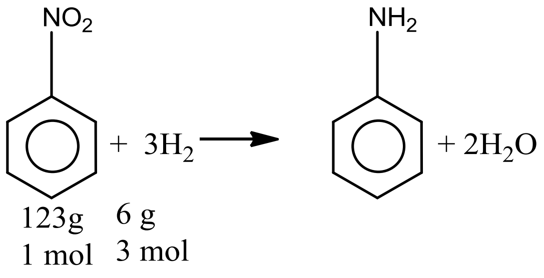
Follow Debye Hückel theory of strong electrolytes.

231 **(d)**

Given, mass of nitrobenzene = 12.3 g

Molecular mass of nitrobenzene = 123 g

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



$\therefore$ H_2 required for reduction of 0.1 mole nitrobenzene

$= 0.1 \times 3 = 0.3$ mol of hydrogen

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

$= 2 \times 96500 \times 0.3$ ($\because H_2 \longrightarrow 2H^+$)

= 57900 C

232 **(a)**

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)**

$E^\circ_{Zn/Zn^{2+}} > E^\circ_{H/H^+}$

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.

$CaH_2 + 2H_2O \longrightarrow Ca(OH)_2 + 2H_2 \uparrow$

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 ∝ 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 ∝ 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

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

At cathode

$$2Cu^{2+}(aq) + 4e^- \rightarrow 2Cu(s)$$

Net electrolysis reaction is

$2Cu^{2+}(aq) + 2H_2O(l) \rightarrow 2Cu(s) + 4H^+(aq) + O_2(g)$

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).

$Mg + ZnO \rightarrow MgO + Zn$

248 **(c)**

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

249 **(b)**

Cell reaction will be

$Br_2 + Sn^{2+} \rightarrow 2Br^- + Sn^{4+}$, $E^\circ = 0.95$ V

$$E^\circ_{cell} = \frac{0.059}{2}\log K_{eq}$$

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

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

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

250 **(b)**

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

251 **(a)**

Ionic mobility

$$\mu_{Ag^+} = \frac{\text{ionic conductance}}{96500}$$

$$= \frac{5 \times 10^{-4}}{96500}$$

$$= 5.2 \times 10^{-9} \text{cm/s}$$

252 **(c)**

Given,

$Fe^{3+} + 3e^- \longrightarrow Fe$; $E^\circ_1 = -0.036$ V ....(i)

$Fe^{2+} + 2e^- \longrightarrow Fe$; $E^\circ_2 = -0.439$ V ....(ii)

We need to calculate

$Fe^{3+} + e^- \longrightarrow Fe^{2+}$ $E^\circ_3 = ?$ .....(iii)

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 extenwise property)

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

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

$$-1 \times F \times E^\circ_3 = -0.770\,F$$

$$E^\circ_3 = 0.770\,F$$

253 **(b)**

$E^\circ_{Cell} = E^\circ_{OP_{Sn}} + E^\circ_{RP_{Pb}}$

or $E^\circ_{Rp_{Sn}} < E^\circ_{RP_{Pb}}$, 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

$$\underline{Cl^- < Fe^{2+} < H_2 < Zn}$$

strength of reducing agent increases

255 **(d)**

$\mathring{\Lambda}_m$ (CH_3COOH)

$$= \mathring{\Lambda}_m(CH_3COONa) + \mathring{\Lambda}_m(HCl) - \mathring{\Lambda}_m(NaCl)$$

$$= 91.0 + 425.9 - 126.4$$

$$= 516.9 - 126.4$$

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

256 **(c)**

$AgNO_3 + KCl \longrightarrow AgCl \downarrow + KNO_3$.

257 **(a)**

Λ^∞_m for $BaCl_2 = \Lambda^\infty_m\, Ba^{2+} + 2\Lambda^\infty_m\, Cl^-$

$\therefore \Lambda^\infty_{eq}$ for $BaCl_2 = 1/2\Lambda^\infty_m\, Ba^{2+} + \Lambda^\infty_m\, Cl^-$

$= 127 / 2 + 76$

$= 139.5\ \Omega^{-1}$ cm^2

258 **(d)**

$-\Delta G = nEF$

$-966 \times 10^3 = 4 \times E \times 96500$

$\therefore E = -2.5$ V

Thus, an e.m.f. of + 2.5 is needed to carryout 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)**

$$\Lambda = k \times \frac{1000}{M}$$

$$= C \times \frac{1}{a} \times \frac{1000}{M}$$

$$\therefore \quad C = \frac{\Lambda \times a \times M}{l \times 1000}$$

$i.e.,\ \ C \propto$ conc.

$$\propto a$$

$$\propto \frac{1}{l}$$

$\therefore$ unit of constant $(\Lambda) = \dfrac{C \times l \times 1000}{a \times M}$

$$= \frac{S \times \text{m}}{\text{m}^2 \times \text{mole m}^{-3}}$$

$$= s\ \text{m}^2\ \text{mol}^{-1}$$

262 **(c)**

Cu can displace Ag from $AgNO_3$ but it cannot liberate H_2 from

HCl because $E^{\circ}_{OP_{\text{Cu}}} > E^{\circ}_{OP_{\text{Ag}}}$ and $E^{\circ}_{OP_{\text{Cu}}} = -$ve.

263 **(a)**

Cell I : $\quad Fe \longrightarrow Fe^{2+} + 2e$

$$\frac{2Fe^{3+} + 2e \longrightarrow 2Fe^{2+}}{Fe + 2Fe^{3+} \longrightarrow n=2\ ;} \qquad \Delta G^{\circ} =$$

$-nE^{\circ}F$

Cell II : $\quad Fe \longrightarrow Fe^{3+} + 3e$

$$\frac{3Fe^{3+} + 3e \longrightarrow 3Fe^{2+}}{Fe + 2Fe^{3+} \longrightarrow Fe^{2+}\ ; \quad n=3;} \qquad \Delta G^{\circ} =$$

$-nE^{\circ}F$

Cell III : $\quad 2Fe^{3+} \longrightarrow 6e + 2Fe$

$$\frac{3Fe \longrightarrow 3Fe^{2+} + 6e}{Fe + 2Fe^{3+} \longrightarrow 3Fe^{2+}\ ; n=3;} \qquad \Delta G^{\circ} =$$

$-nE^{\circ}F$

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 places below it in electrochemical series.

$\therefore$ $CuSO_4$ reacts with Zn and Fe placed above it.

$CuSO_4 + Fe \rightarrow FeSO_4 + Cu$

$CuSO_4 + Zn \rightarrow ZnSO_4 + Cu$

$\therefore$ $CuSO_4$ does not react with Ag, placed below it.

$CuSO_4 + Ag \rightarrow$ no reaction.

265 **(c)**

Rust is Fe_2O_3 and $Fe(OH)_3$.

267 **(c)**

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

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

268 **(a)**

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

$Cu^{2+} + e^{-} \rightarrow Cu^{+}, E^{\circ}_{Cu^{2+}/Cu^{+}} = 0.15$ V $\quad$ ...(i)

$Cu^{2+} + 2e^{-} \rightarrow Cu, E^{\circ}_{Cu^{2+}/Cu} = 0.34$ V $\quad$ ...(ii)

Multiplying Eq. (i) by 2, we get

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

$\Delta G_1 = -nFE = -2 \times F \times 0.15$ $\quad$ ...(iii)

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

$\Delta G_2 = -nFE = -2 \times F \times 0.34$ $\quad$ ...(iv)

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

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

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

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

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

$E^{\circ} = -0.38$

This is the value for the reaction

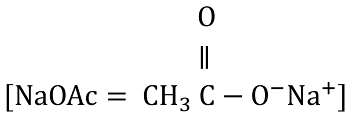
$Cu^{2+} + Cu \rightarrow 2Cu^{+}$

But the given reaction is just reverse of it

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

269 **(d)**

According to Kohlrausch's law



$$[\text{NaOAc} = CH_3\overset{\overset{\large O}{\|}}{C} - O^{-}Na^{+}]$$

$\Lambda^{\circ}_{CH_3COOH} = \Lambda^{\circ}_{CH_3COO^-} + \Lambda^{\circ}_{H^+}$ $\quad$ ...(i)

$\Lambda^{\circ}_{HCl} = \Lambda^{\circ}_{H^+} + \Lambda^{\circ}_{Cl^-}$ $\quad$ ..(ii)

$\Lambda^{\circ}_{CH_3COONa} = \Lambda^{\circ}_{CH_3COO^-} + \Lambda^{\circ}_{Na^+}$ $\quad$ ...(iii)

Thus, on adding (ii) and (iii), if $\Lambda^{\circ}_{Na^+}$ and $\Lambda^{\circ}_{Cl^-}$ are subtracted we can obtained the value of Λ°_{HOAc}. Thus, additional value required is Λ°_{NaCl}.

270 **(c)**

$\Lambda^{0}_{CH_3COOH} = \Lambda^{0}_{CH_3COONa} + \Lambda^{0}_{HCl} - \Lambda^{0}_{NaCl}$

271 **(b)**

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

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

272 **(a)**

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

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

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, $[OH^-]$ increases in solution.

275 **(b)**

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

$Fe^{3+} + e \longrightarrow Fe^{2+}$; $E^\circ_{RP} = 0.77$ V

$[Fe(CN)_6]^{3-} + e \longrightarrow [Fe(CN)_6]^{4-}$; $E^\circ_{RP} = 0.35$ V

276 **(a)**

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

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

277 **(c)**

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

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

$= 2.45$ ampere

278 **(b)**

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

$= E^\circ_{\text{cathode}} + E^\circ_{\text{anode}}$

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

$= + 0.11$ V

Since, E°_{cell} is positive for the above reaction . Hence, this reaction is spontaneous.

279 **(a)**

At anode; $\frac{1}{2} H_2 + OH^- \rightarrow H_2O + e^-$

At cathode; $H^+ + e^- \rightleftharpoons \frac{1}{2}H_2$

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

$$K = \frac{[H_2O]}{[H^+][OH^-]}$$

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

280 **(b)**

$E^\circ_{M^{2+}/M}$ values follow the order with negative sign

Mn $< Cr < Fe < Co$

E°_{RP} -1.18 -0.74 -0.44 -0.27 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 × cell constant

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

$\therefore$ Cell constant

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

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 $\Delta G^\circ = -nE^\circ F$)

12. Cu $(s) \rightarrow Cu^{2+} (aq) + 2e^-$

$$\Delta G^\circ_1 = -2 \times (-0.34) \times F$$

13. $Cu^{2+} (aq) + e^- \rightarrow Cu^{2+} (aq)$

$$\Delta G^\circ_2 = -1 \times (0.15) \; F$$

On addition,

Cu $(s) \rightarrow Cu^{2+} (aq) + e^-. \Delta G^\circ_1 = -1 \times E^\circ \times F$

$\Delta G^\circ_3 = \Delta G^\circ_1 + \Delta G^\circ_2$

$= (-2 \times -0.34 \times F) + (-1 \times 0.15 \times F)$

$=+ 0.68 \text{ F} - 0.15 \text{ F} = 0.53 \text{ F}$

0r $E^\circ = -0.53\,V$

Reaction,

$2Cu^+(aq) \rightleftharpoons Cu^{2+} (aq) + Cu\,(s),\; E^\circ = ?$

So, $Cu^+ (aq) + e^- \rightleftharpoons Cu\,(s)$, $E^\circ = +0.53\,V$

$Cu^+ (aq) \rightleftharpoons Cu^{2+} (aq) + e^-$; $E^\circ = -0.15$ V

$2Cu^+ \text{ (aq)} \rightleftharpoons Cu^{2+} \text{ (aq)} + Cu \text{ (s)};\; E^0 = +0.38V$

286 **(d)**

The cell reaction is,

$\frac{1}{2} H_2(g) + AgCl(s) \longrightarrow H^+(aq) + Cl^-(aq) + Ag \text{ (s)}$

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

Pt | H_2(g) , HCl solution || AgCl(s) | Ag

287 **(b)**

For all Zn | Zn^{2+} (A = 0.1 M) || Fe^{2+} (A = 0.01 M) | Fe.

The cell reaction,

$\mathrm{Zn}\,(s) \rightarrow \mathrm{Zn}^{2+}\,(aq) + 2e^-$

$\mathrm{Fe}^{2+}\,(aq) + 2e^- \rightarrow \mathrm{Fe}\,(s)$

$\mathrm{Zn}\,(s) + \mathrm{Fe}^{2+}\,(aq) \rightarrow \mathrm{Zn}^{2+}\,(aq) + \mathrm{Fe}\,(s)$

On applying Nernst equation,

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

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

$$0.2905 = E^\circ_{\text{cell}} - 0.0295 \times \log_{10} 10$$

$$0.2905 = E^\circ_{\text{cell}} - 0.0295 \times 1$$

$$\therefore E^\circ_{cell} = 0.2905 + 0.0295 = 0.32 \text{ V}$$

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

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

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

$$\text{or } E^\circ_{\text{cell}} = \frac{0.0591}{2} \log_{10} K_c$$

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

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

288 **(a)**

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

289 **(a)**

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

290 **(c)**

Eq. of Ag = Eq. of Hg^{2+} = Eq. of Hg_2^{2+}

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

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

293 **(d)**

For the cell, Ni | Ni^{2+} || Au^{3+}| Au

Given, $E^\circ_{\mathrm{Ni}^{2+}/\mathrm{Ni}} = -0.25\text{V}$

$E^\circ_{\mathrm{Au}^{3+}/\mathrm{Au}} = +1.5\ V$

Here, Ni is anode and Au is cathode.

$\therefore \quad E_{\text{cell}} = E_C - E_A$

$= 1.5 - (-0.25)$

$= 1.5 + 0.25$

$= 1.75$ V

294 **(c)**

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

Hence,

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$= -0.44 - (-0.14) = -0.30 \text{ V}$$

The negative emf suggests that the reaction goes spontaneously in reversed direction

295 **(c)**

$\mathrm{Zn}\,(s) + 2\mathrm{H}^+\,(aq) \rightarrow \mathrm{Zn}^{2+}(aq) + \mathrm{H}_2\,(\mathrm{g})$

Reaction quotient $Q = \frac{[\mathrm{Zn}^{2+}]}{[\mathrm{H}^+]^2}$

Corresponding cell is

Zn | Zn^{2+} (C_1) || H^+ (aq) | Pt (H_2)

-anode + cathode

$$\text{and } E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{2} \log K$$

$$= E^\circ_{\text{cell}} - \frac{0.0591}{2} \log \frac{[\mathrm{Zn}^{2+}]}{[\mathrm{H}^+]^2}$$

$$= E^\circ_{\text{cell}} + \frac{0.0591}{2} \log \frac{[\mathrm{H}^+]^2}{[\mathrm{Zn}^{2+}]}$$

If H_2SO_4 is added to cathodic compartment, (towards reactant side), then Q decreases (due to increase in H^+).

Hence, equilibrium is displaced towards right and E_{cell} increases.

296 **(c)**

$$E_{\text{cell}} = E^\circ_{\text{cell}} + \frac{0.059}{2} \log \frac{[\mathrm{Ag}^+]^2}{[\mathrm{Sn}^{2+}]}$$

297 **(a)**

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

$\mathrm{Fe}^{2+} + 2e^- \longrightarrow \mathrm{Fe}$

$\Delta G^\circ = -2 \times \mathrm{F} \times (-0.440 \text{ V})$

$= 0.880$ F ...(i)

$\mathrm{Fe}^{3+} + 3e^- \longrightarrow \mathrm{Fe}$

$$\Delta G^\circ = -3 \times F \times (-0.036\ \text{V})$$
$$= 0.108\ \text{F} \qquad \text{...(ii)}$$

On subtracting Eqs. (i) from (ii)

$$Fe^{3+} + e^- \rightarrow Fe^{2+}$$
$$\Delta G^\circ = 0.108\ \text{F} - 0.880\ \text{F} = -0.772\ \text{F}$$
$$E^\circ = -\frac{\Delta G^\circ}{nF} = -\frac{-0.772\ F}{1 \times F} = +0.772\ \text{V}$$

298 **(b)**

$$E^\circ_{\text{cell}} = E_{OP_{\text{Al}}} + E^\circ_{RP_{\text{H}}}$$
$$\therefore\ 1.66 = E^\circ_{OP_{\text{Al}}} + 0$$

300 **(a)**

$$\Lambda^\infty_{NH_4OH} = \Lambda^\infty_{NH_4Cl} + \lambda^\infty_{OH^-} - \lambda^\infty_{Cl^-}$$

302 **(c)**

Number of g-equivalent = number of faraday pass

4 g = 4 F

303 **(c)**

Eq. of Al deposited = 5

$\therefore$ wt. of Al $= 5 \times \frac{27}{3} = 45$ g

304 **(d)**

At equilibrium $E_{\text{cell}} = 0$.

305 **(b)**

For, $Sn(s) + 2Fe^{3+}(aq) \rightarrow 2Fe^{2+}(aq) + Sn^{2+}(aq)$

$$E^\circ_{\text{cell}} = E^\circ_{Sn/Sn^{2+}} + E^\circ_{Fe^{3+}/Fe^{2+}}$$

$$= (0.14) + (0.77) = 0.91\ \text{V}$$

306 **(c)**

This is Kohlrausch law for a_2b type of electrolytes.

307 **(b)**

$E^\circ_{OP} = -E^\circ_{RP}$ for any element.

308 **(d)**

Rusting of iron is catalysed by $[H^+]$

309 **(d)**

$$\text{Resistivity} = \frac{1}{k}$$

Or $k = \frac{1}{15.8}\ \text{ohm}^{-1}\ \text{cm}^{-1}$

Also, $\Lambda_v = \frac{k \times 1000}{M} = \frac{1 \times 49 \times 1000}{18.5 \times 15}$

Now, $\alpha = \frac{\Lambda_v}{\Lambda_M} = \frac{1 \times 49 \times 1000}{18.5 \times 15 \times 348} = 0.507$

Or $= 50.7\%$

310 **(b)**

$$E_{\text{cell}} = E^\circ_{\text{anode(op)}} - E^\circ_{\text{cathode(op)}}$$
$$= 0.76 - 0.41$$
$$= +0.35\ \text{V}$$

312 **(b)**

Standard oxidation potentials of Zn, Cu, Ag and Ni electrode are + 0.76, - 0.34, - 0.80 and + 0.25 V respectively. (Given)

$Zn^{2+} + 2e^- \rightarrow Zn$; E_{cell} = - 0.76 V

$Cu^{2+} + 2e^- \rightarrow Cu$; E_{cell} = + 0.34 V

$Ag^+ + e^- \rightarrow Ag$; E_{cell} = + 0.80 V

$Ni^{2+} + 2e^- \rightarrow Ni$; E_{cell} = - 0.25 V

14. $Cu + 2Ag\ (aq) \rightarrow Cu^{2+}\ (aq) + 2\ Ag$

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$E_{\text{cell}} = 0.80 - 0.34 = 0.46$ V

15. $Zn + 2\ Ag^+\ (aq) \rightarrow Zn^{2+}\ (aq) + 2\ Ag$

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

(maximum voltage)

16. $H_2 + Ni^{2+}\ (aq) \rightarrow 2H^+\ (aq) + Ni$

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

17. $Zn + Cu^{2+}\ (aq) \rightarrow Zn^{2+}\ (aq) + Cu$

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

18. $Zn + 2\ H^+\ (aq) \rightarrow Zn^{2+}\ (aq) + H_2$

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

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)**

$E^\circ_{RP_{Cr_2O_7^{2-}/Cr^{3+}}} = 1.33$ V and $E^\circ_{RP_{Fe^{3+}/Fe^{2+}}} = 0.77$ V

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

$$E^\circ_{\text{cell}} = E^\circ_{OP_{\text{Fe}}} + E^\circ_{RP_{\text{Cr}}} = -0.77 + 1.33$$

315 **(c)**

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

$$Zn + 2NH_4Cl \rightarrow Zn\ (NH_3)_2Cl_2 + \underset{\text{Colourless gas}}{H_2 \uparrow}$$

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 ofE°_{OP} .

318 **(a)**

Oxidizing tensdncy $\propto \dfrac{1}{\text{electrode potential}}$

$TX \rightarrow$ No reaction

$TY \rightarrow X, Z$

$TZ \rightarrow X$

⇒ Order of electrode potential is

$TY < TZ < TX$

⇒ Order of oxidation of the anion is

$Y^- > Z^- > X^-$

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)**

$\text{Sn}\,(aq) + 2\text{Fe}^{3+}(aq) \rightarrow 2\text{Fe}^{2+}(aq) + \text{Sn}^{2+}(aq)$

$E^\circ_{\text{cell}} = E^\circ_{\text{oxi}} + E^\circ_{\text{red}}$

$= E^\circ_{\text{Sn}\,/\text{Sn}^{2+}} + E^\circ_{\text{Fe}^{3}/\,\text{Fe}^{2+}}$

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

$\therefore \quad E^\circ_{\text{Sn}\,/\text{Sn}^{2+}} = +0.14\ \text{V}$

$E^\circ_{\text{Fe}^{3+}/\,\text{Fe}^{2+}} = 0.77\ \text{V}$

$E^\circ_{cell} = 0.14 + 0.77 = 0.91\ \text{V}$

322 **(d)**

$\text{Zn}^{2+} + 2\text{e}^- \longrightarrow \text{Zn};\ E^\circ = -0.76\,V$

$\text{Fe}^{2+} + 2\text{e}^- \longrightarrow \text{Fe};\ E^\circ = -0.44\,V$

Cell reaction is

$\text{Fe}^{2+} + \text{Zn} \longrightarrow \text{Zn}^{2+} + \text{Fe}$

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

= - 0.44 – (0.76)

= - 0.44 + 0.76

=0.32 V

323 **(c)**

Reduction hydrogen half-cell is

$\text{H}^+ \mid (xM) \mid \text{Pt}\,(\text{H}_2)$

\+

Pressure $p\text{H}_2$

Half – cell reaction is

$2\text{H}^+\,(aq) + 2\text{e}^- \longrightarrow \text{H}_2\,(\text{g})$

Reaction quotient

$= Q = P_{H_2}/[\text{H}^+]^2, n = 2$

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

	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

$E^\circ_{\text{red}} = 0.00$ V for standard hydrogen electrode

If $Q > 1$, then E_{red} = -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 H_2(g) and O_2(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 CH_3COOH 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)**

$\text{Ag}^+ + e^- \rightarrow \text{Ag}, E_{\text{Ag}} = \dfrac{\text{atomic mass}}{1} = 108$

Number of faraday $= \dfrac{w_{\text{Ag}}}{E_{\text{Ag}}} = \dfrac{108}{108} = 1$

335 **(a)**

$[\text{H}^+] = c.\alpha = 0.0133 \times 0.1$

$= 0.00133$ M.

336 **(a)**

$E_1 = E^\circ - \dfrac{0.059}{1}\log[\text{H}^+]_1$

$E_2 = E^\circ - \dfrac{0.059}{1}\log[\text{H}^+]_2$

on adding (also $E^\circ_{\text{H}} = 0$)

$E_1 + E_2 = -\dfrac{0.059}{1}[\log(\text{H}^+)_1 + \log(\text{H}^+)_2]$

Now for $\text{CH}_3\text{COOH} = \text{CH3COO}^- + \text{H}^+$

$[\text{H}^+] = \dfrac{K_a[\text{CH}_3\text{COOH}]}{[\text{CH}_3\text{COO}^-]}$

$\therefore [\text{H}^+]_1 = K_a.\frac{y}{x}$

$[H^+]_2 = K_a \cdot \frac{x}{y}$

$\therefore E_1 + E_2 = -\frac{0.059}{1}\left[\log\frac{K_a \cdot y}{x} + \log\frac{K_a \cdot x}{y}\right]$

$= -0.059[2\log K_a]$

$\log K_a = \frac{E_1+E_2}{2\times(-0.059)}$

$\log K_a = \frac{E_1+E_2}{0.118}$

or $pK_a = \frac{E_1+E_2}{0.118}$

337 **(a)**

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

$\therefore \frac{112\times 2}{22400} = \frac{1\times i \times 965}{96500}$

$\therefore \quad i = 1$ ampere

338 **(d)**

pH = 3, $[H^+] = 10^{-3}$

$E = E^\circ_{red} + 0.059 \log(\text{ion})$

$E = 0 + 0.059 \log(10^{-3})$

$E = +0.059(-3) = -0.177\,V$

339 **(d)**

Given, $E^\circ_{cell} = 0.34\,V$

$E^\circ_{cell} = E^\circ_{(Cu^{2+}/Cu)} - E^\circ_{(H_2/H^+)}$

$0.34 = E^\circ_{(Cu^{2+}/Cu)} - 0.00$

$\therefore \quad E^\circ_{(Cu^{2+}/Cu)} = +0.34$

340 **(d)**

Reduction potential of hydrogen electrode,

$$E_H = -\frac{2.303RT}{F}\log\frac{1}{[H^+]}$$

$= -0.059$ pH $= -0.059 \times 3 = -0.177$ 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

$2H_2O \rightleftharpoons 4H^+ + 2O^{2-}$

$2O^{2-} \rightarrow O_2 + 4e^-$

$4e^- + 4H^+ \rightarrow 2H_2$

$\therefore n = 4$,

So, 4 F charge liberates = 1 mol = 22.4 dm^3 oxygen

$\therefore$ 1 F charge will liberate $= \frac{22.4}{4} = 5.6\ dm^3$ oxygen

343 **(a)**

$Q = 2.5 \times 386 = 965$ C

$2F(2\times 96500\ C)$ deposited, Cu = 63.5 g

Hence, 965 C will deposit, Cu = 0.3175 g

344 **(b)**

Given, product = 0.1 M and reactant = 1 M

$E^\circ_{cell} = E^\circ_{cell} - \frac{0.0591}{2}\log\frac{[\text{products}]}{[\text{reactants}]}$

$1.50 = E^\circ_{cell} - 0.02955 \log\left[\frac{0.1}{1}\right]$

$E^\circ_{cell} = 1.470\,V$

$E^\circ_{cell} = E^\circ_{H^+/H_2} - E^\circ_{M/M^{2+}}$

$E^\circ_{M/M^{2+}} = -1.470$

So, $E^\circ_{M^{2+}/M} = 1.470\,V$

345 **(c)**

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

Thus, on doubling concentration of both Cu^{2+} and Zn^{2+}, there will be no effect no E_{cell}.

346 **(a)**

$\Lambda^\infty_{H_2O} = \lambda^\infty_{H^+} + \Lambda^\infty_{OH^-}$

347 **(a)**

$\kappa = \Lambda_{eq} \cdot C$

$$= (91\ \Omega^{-1}\ cm^2\ eq^{-1})\left(\frac{2.54}{159/2\times 1000}\ eq\cdot cm^{-3}\right)$$

$= 2.9\times 10^{-3}\ \Omega^{-1}\ cm^{-1}$

348 **(a)**

MnO_4^- will oxidise Cl^- ion according to equation

$Mn^{7+} + 5e \longrightarrow Mn^{2+}$

$2Cl^- \longrightarrow Cl_2 + 2e$

Thus, $E^\circ_{cell} = E^\circ_{OP_{Cl^-/Cl_2}} + E^\circ_{RP_{MnO_4^-/Mn^{2+}}}$

$= -1.40 + 1.51 = 0.11$ V or reaction is feasible MnO_4^- will oxidise Fe^{2+} to Fe^{3+}

$Mn^{7+} + 5e \longrightarrow Mn^{2+}$

$Fe^{2+} \longrightarrow Fe^{3+} + e$

$\therefore \quad E^\circ_{cell} = E^\circ_{OP_{Fe^{2+}/Fe^{3+}}} + E_{MnO_4^-/Mn^{2+}}$

$= -0.77 + 1.51$

$= 0.74$ 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)**

$E_{cell} = E_C - E_A$

Given, $\quad E^\circ_{Ag^+/Ag} = 0.80$ V

$H_2 \mid H^+ \mid\mid Ag^+ \mid Ag$

$\therefore$ Hydrogen is anode and silver is cathode.

$E_{\text{cell}} = E_C - E_A$

$= 0.80 - 0 \quad (\because E^{\circ}_{\text{H}^+/\text{H}} = 0)$

$= 0.80 \text{ V}$

351 **(d)**

$$2\text{H}^+ + 2e \longrightarrow \text{H}_2; 2\text{OH}^- \longrightarrow \text{H}_2\text{O} + \frac{1}{2}\text{O}_2 + 2e$$

352 **(b)**

$\text{Fe}^{2+} + 2e^- \rightarrow \text{Fe}$

$E_{\text{Fe}} = \frac{56}{2} = 28$

$w_{\text{Fe}} = \text{E}_{\text{Fe}} \times$ number of faraday

$= 28 \times 3 = 84$ 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)**

$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{red (cathode)}} - E^{\circ}_{\text{oxi (anode)}}$

$= E^{\circ}_{\text{Ag}^+/\text{Ag}} - E^{\circ}_{\text{Cu+ /Cu}^{2+}}$

$= 0.80 - (+0.34) = +0.46 \text{ V}$

357 **(c)**

Because conductance increases when the dissociation is more

358 **(a)**

At cathode $2\text{H}^+(aq) + 2e^- \rightarrow 2\text{H}$

$$2\text{H} + \frac{1}{2}\text{O}_2 \rightarrow \text{H}_2\text{O}$$

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

---

359 **(b)**

Resistance of $\frac{\text{N}}{10}$ 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 $\text{H}_2 \left| \underset{c_1}{\text{HCl}} \right| \left| \underset{c_2}{\text{HCl}} \right| \text{Pt H}_2$) are used. Moreover electrolyte of one strength but electrodes of two different concentrations are used.

361 **(b)**

$\text{E}^{\circ}_{\text{cell}} = \text{E}^{\circ}_{\text{(RHS)}} - \text{E}^{\circ}_{\text{(LHS)}} = 1.61 - 1.51 = 0.10 \text{ V}$

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

$= -5 \times 96500 \times 0.10 \text{ J}$

$\Delta G^{\circ} = -48.25 \text{ kJ}$

362 **(a)**

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

or $0.635 = \frac{63.5 \times i \times t}{2 \times 96500}$

$\therefore \quad i \times t = 1930 \text{ C}$

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)**

$\because R \propto l$ and, $R \propto \frac{1}{a} \ \therefore R \propto \frac{1}{a}$ or $R = \rho . \frac{1}{a}$

Where R is resistance, ρ is specific resistance and l, a are length and area of cross-section of wire.

$\because \quad \kappa = \frac{1}{\rho}$ where κ is conductivity

Then $\kappa = \frac{1}{R} \times \frac{1}{a} = C \times \frac{1}{a}$;

Where C is conductance $= \frac{1}{R}$

Also, Equivalent conductivity = Conductivity × $V_{\text{in mL}}$ where V is solution containing 1 equivalent in it or

Eq. conductivity $= \kappa \times \frac{1000}{N}$

366 **(b)**

1 g atom of Mg = 2g eq. Mg = 2 faraday.

367 **(b)**

$\text{NaCl} \longrightarrow \text{Na}^+ + \text{Cl}^-$

molten

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

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

When one mole of NaCl is electrolysed, 1 mole Na and $\frac{1}{2}$ mole Cl_2 is obtained.

Thus, when 230 g$\left(\frac{230}{23} = 10\text{mol}\right)$ Na is formed, the moles of

Cl_2obtained will be = $\frac{\text{moles of Na}}{2}$

$= \frac{10}{2}$

= 5.0 mol

368 **(a)**

Cell reaction : $\frac{1}{2}H_2(g) + Ag^+(x) \rightleftharpoons Ag(x) + H^+(1\ M)$

$$E = E^\circ = -\frac{0.0591}{n}\log\frac{[H^+]}{[Ag^+]}$$

$$0.503 = 0.80 - \frac{0.0591}{1}\log\frac{1}{x}$$

$$x = 9.43 \times 10^{-6}\ M$$

Number of moles of Ag^+ in 350 mL

$$= \frac{9.43 \times 10^{-6} \times 350}{1000} = 3.3 \times 10^{-6}$$

Mass of Ag = $3.3 \times 10^{-6} \times 108 = 3.56 \times 10^{-4}$ g

% 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)**

$$E_{cell} = \frac{0.059}{1}\log\frac{[H^+]_{RHE}}{[H^+]_{LHE}}$$

$= 0.059\log\frac{10^{-3}}{10^{-8}} = 0.059 \times 5 = 0.295$ V

371 **(a)**

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

$$= 1.10 - \frac{0.059}{2}\log\frac{0.1}{0.1}$$

= 1.10 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

$\therefore \quad \frac{0.3}{59/2} = \frac{W}{52/3}$

$\therefore \quad W_{Cr} = 0.176$ g

374 **(b)**

$NaCl + KNO_3 \longrightarrow NaNO_3 + KCl$

KCl – NaCl

Molar conductivities 152 128

Difference = 24 S $cm^2\ mol^{-1}$

Molar conductivities of KNO_3 = 111

Molar conductivities of $NaNO_3$ = 111 – 24

= 87 S $cm^2\ mol^{-1}$

375 **(d)**

The degree of ionisation of an electrolyte increases with dilution and temperature as well as in presence of polar solvent.

376 **(a)**

$E^\circ_{cell} = E_C - E_A$

= - 0.45 – (-2.37)

= -1.92 V

Mg (s) | Mg^{2+}(aq) x M || Fe^{2+}(aq) 0.01 M | Fe (s)

The cell reaction is

$Mg + Fe^{2+} \rightarrow Mg^{2+} + Fe$

$$E_{cell} = E^\circ_{cell} - \frac{0.059}{n}\log\frac{[Mg^{2+}]}{[Fe^{2+}]}$$

$$1.92 = 1.92 - \frac{0.059}{n}\ \log\frac{x}{0.01}$$

$$0 = \frac{-0.059}{2}\ \log\frac{x}{0.01}$$

$\therefore \quad x = 0.01$ M

377 **(d)**

Follow theory of corrosion.

378 **(b)**

$2MnO_2\ (s) + Zn^{2+} + 2e^- \longrightarrow ZnMn_2\ O_4\ (s)$

2 × 87 g MnO_2 required 2 × 96500 F charge

8 g MnO_2 = $\frac{2 \times 96500 \times 8}{2 \times 87}$

= 8873.56 F

$Q = it$

$8873.56 = 2 \times 10^{-3} \times t$

$$t = \frac{8873.56 \times 10^3}{2}$$

= 4436781.5 s

$$= \frac{4436781.5}{24 \times 60 \times 60}$$

= 51.35 days

379 **(d)**
The cell reaction is
Cd (s) + 2AgCl (s) → 2 Ag (s) + Cd^{2+} (aq) + $2Cl^-$ (aq)
$E_1 = 0.6915$ V at 0°C
$E_2 = 0.6753$ V at 25°C
Now, $\frac{\partial E_{cell}}{\partial T} = \frac{E_2 - E_1}{T_2 - T_1}$
$= \frac{0.6753 - 0.6915}{298 - 273}$
$= -6.48 \times 10^{-4}$
$\Delta S = nF\left[\frac{\partial E_{cell}}{\partial T}\right]$
Now, we put the value
$\Delta S = 2 \times 96500\ (-6.48 \times 10^{-4}) = -125.064$
We know that,
$\Delta G = -nFE_{cell}$
$= -2 \times 96500 \times 0.6753$
$= -1.303 \times 10^5$
As, $\Delta G = \Delta H - T\Delta S$
For calculating $\Delta H = \Delta G + T\Delta S$
$= -1.303 \times 10^5 + 298\ (-125.064\ kJ)$
$\Delta H = -1.6726 \times 10^5$ J
$= -167.26$ kJ

381 **(c)**
$$E_{red} = 0.78 + \frac{0.0591}{1}\log(10^{-7})^2$$
$= 0.78 - 0.059 \times 7 = 0.367$ V

382 **(b)**
$$\frac{\text{wt. of Al deposited}}{\text{wt. of Cu deposited}} = \frac{\text{eq. wt. of Al}}{\text{eq. wt. of Cu}}$$
$$\frac{2.7}{\text{wt.of Cu}} = \frac{27/3}{63.5/2}$$
wt. of Cu = 9.525 g

383 **(d)**
$E = Q.V = 1C \times 1V = 1J$. It is unit of energy. Also $1J = 10^7$ erg.

384 **(a)**
$$E_{Cu/Cu^{2+}} = E^\circ_{Cu/Cu^{2+}} - \frac{0.059}{2}\log[Cu^{2+}]$$
if $\log[Cu^{2+}] = 0$, i.e., $[Cu^{2+}] = 1$, then $E_{Cu/Cu^{2+}} = E^\circ_{Cu/Cu^{2+}}$
or $OA = E^\circ_{Cu/Cu^{2+}} = -E^\circ_{Cu/Cu^{2+}} = -0.34$
Now, $E_{Cu/Cu^{2+}} = -0.34 - \frac{0.059}{2}\log 0.1$
$$= -0.34 + \frac{0.059}{2}\text{V}$$

385 **(a)**
$E^\circ_{cell} = E^\circ_{OP_L} + E^\circ_{RP_R} = -\phi_L + \phi_R$.

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)**
$$H_2O \longrightarrow H_2 + \frac{1}{2}O_2.$$
$2H^+ + 2e^- \longrightarrow H_2$
∴ 1 mole of H_2 is liberated from 2 moles of electrons.
2 moles electrons = 2 F charge
= 2 × 96500 C charge
Given, $i = 4\,A, t = 30$ min
∴ 2 × 96500 C charge liberates
= 22400mL H_2 at NTP
∴ 4 × 30 × 60 C charge liberates
$= \frac{22400}{2 \times 96500} \times 4 \times 30 \times 60$ mL H_2
= 835.6mL
= 0.836 L of H_2

389 **(d)**
$R = 100\Omega$
$$\kappa = \frac{1}{R}\left(\frac{l}{a}\right)$$
$\frac{l}{a}$ (cell constant) = $1.29 \times 100\text{m}^{-1}$
Given, $R = 520\Omega$; $C = 0.2$ 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)
Hence, $\mu = \frac{1}{520} \times 129 \times \frac{1000}{0.2} \times 10^{-6}\ \text{m}^3$
$= 12.4 \times 10^{-4}\ \text{S m}^2\ \text{mol}^{-1}$

390 **(a)**
$$u^0 = \frac{\lambda^0}{96500} = \frac{5 \times 10^{-3}}{96500} = 5.18 \times 10^{-8}$$

391 **(b)**
$Q = 10^{-6} \times 1C$;
Also, 96500 C = 6.023×10^{23} 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

$= \frac{12 \times 9.65}{96500} = 1.2 \times 10^{-3}\text{g Mg}$

$= \frac{1.2 \times 10^{-3}}{24} = 5 \times 10^{-5}\text{ mol Mg}$

$R - X + Mg \longrightarrow RMg\,X$ Grignard reagent

5×10^{-5} Mol $\quad 5 \times 10^{-5}$ mol

Hence, number of moles of $R\text{Mg}X$ produced

$= 5 \times 10^{-5}$.

394 **(a)**

$\Delta G = \Delta H - T\Delta S, \Delta G = -nEF$ and

$$\Delta G = \Delta H + T\left[\frac{\partial(\Delta G)}{\partial T}\right]_P$$

$$\therefore \quad \Delta H = -nF\left[E - T\left(\frac{\partial E}{\partial T}\right)_P\right]$$

$$\therefore \quad \Delta S = nF\left(\frac{\partial E}{\partial T}\right)_P$$

395 **(a)**

Equal equivalent of each are liberated.

Eq. of X = Eq. of Y

$\frac{W_1}{2M/2} = \frac{W_2}{M/1} \qquad \therefore W_1 = W_2$

396 **(a)**

$\Delta G° = -nFE°$

$Fe^{2+} + 2e^- \rightarrow Fe$ ...(i)

$\Delta G° = -2 \times F \times (-0.440\text{ V}) = 0.880\text{ F}$

$Fe^{3+} + 3e^- \rightarrow Fe$ ...(ii)

$\Delta G° = -3 \times F \times (-0.036) = 0.108\text{ F}$

On subtracting Eqs. (i) from (ii), we get

$Fe^{3+} + e^- \rightarrow Fe^{2+}$

$\Delta G° = 0.108\text{ F} - 0.880\text{ F} = -0.772\text{ F}$

$E°$ for the reaction $= -\frac{\Delta G°}{nF} = -\frac{(-0.772\text{ F})}{1 \times \text{F}}$

$= +0.772\text{ V}$

397 **(c)**

$E^{\circ}_{\text{cell}} = E^{\circ}_{OP_{\text{ZN}}} + E^{\circ}_{RP_{\text{Cu}}} E°_{\text{RP}_{\text{Cu}}}$

$= 0.76 + 0.34 = +1.10\text{ V}.$

398 **(d)**

$\lambda_c^\infty = \mu_c^\infty \times$ Faraday;

$\lambda_a^\infty = \mu_a^\infty \times$ Faraday;

Also, $\lambda^\infty_{\text{electrolyte}} = \lambda_a^\infty \times \lambda_c^\infty$.

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

or $\frac{1}{3}$ mole Al = $\frac{1}{2}$ mole Cu = 1 mole Na

or 2 : 3 : 6 or 1 : 1.5 : 3 mole ratio.

402 **(c)**

In Agalvanic cell, oxidation (*i.e.*, removal ofe^-) 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)**

$2Br^-(aq) + Cl_2 \longrightarrow 2Cl^-(aq) + Br_2$

Bromine is relased 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 Fe^{2+} = Eq. of Fe^{3+}

or $\frac{W_1}{A/2} = \frac{W_2}{A/3}$

or $\frac{W_{Fe^{2+}}}{W_{Fe^{3+}}} = \frac{3}{2}$

406 **(b)**

$\mathring{\Lambda}_m(NH_4Cl) = \mathring{\Lambda}_m(NH_4^+) + \mathring{\Lambda}_m(Cl^-)$

.....(i)

$\mathring{\Lambda}_m(NaOH) = \mathring{\Lambda}_m(Na^+) + \mathring{\Lambda}_m(OH^-)$

.....(ii)

$\mathring{\Lambda}_m(NaCl) = \mathring{\Lambda}_m(NH^+) + \mathring{\Lambda}_m(Cl^-)$

.....(iii)

By (i) + (ii) + (iii)

$4\mathring{\Lambda}_m\,(NH_4OH) = \mathring{\Lambda}_m\,(NH_4^+) + \mathring{\Lambda}_m(OH^-)$

$= \mathring{\Lambda}_m(NH_4Cl) + \mathring{\Lambda}_m(NaOH)$

$- \mathring{\Lambda}_m(NaCl)$

407 **(b)**

$$\begin{array}{l} Cu^{2+} + 2e \longrightarrow Cu; \quad \Delta G_1^\circ = -2 \times 0.337 \times F \\ \underset{-}{Cu}^{2+} + \underset{-}{e} \longrightarrow \underset{-}{Cu}^{+}; \quad \underset{-}{\Delta G_2^\circ} \underset{+}{=} 1 \times 0.513 \times F \\ \hline Cu^+ + e \longrightarrow Cu; \ \Delta G_3^\circ = \Delta G_1^\circ - \Delta G_2^\circ = -0.52\,F \end{array}$$

$\therefore -1 \times E° \times F = 0.52\,F \Longrightarrow E° = 0.52\text{ V}$

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) $E^{\circ}_{\text{cell}} = 3.6$ V.

410 **(c)**

6×10^{23} electron = 1 eq.

411 **(d)**

$E^{\circ}_{\text{cell}} = E^{\circ}_{OP_{\text{I}}} + E^{\circ}_{RP_{\text{II}}}$; I is oxidised, II is reduced.

412 **(a)**

Charge = 2 × charge of electron = $2 \times 1.602 \times 10^{-19}$ 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)**

$Mg \rightarrow Mg^{2+} + 2e^-$ (at anode)

$Cu^{2+} + 2e^- \rightarrow Cu$ (at cathode)

$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$

$= +0.34 - (-2.37)$

$= +2.71$ 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 $\left(E^\circ_{Mg^{2+}/Mg} < E_{H_2O/\frac{1}{2}H_2+OH^-}\right)$. 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 $AxBy$.

422 **(c)**

$Sn^{4+} + 2e \longrightarrow Sn^{2+}$; $E^\circ_{Rp} = 0.15$ V

$Cr^{3+} + 3e \longrightarrow Cr$; $E^\circ_{RP} = -0.74$ (Higher E_{OP})

Redox change $2Cr + 2Sn^{4+} \longrightarrow 2Cr^{3+} + 3Sn^{2+}$

$\therefore E_{\text{cell}} = E^\circ_{OP_{\text{Cr}}} + E^\circ_{RP_{\text{Sn}}} = 0.74 + 0.15 = +0.89$ V

423 **(c)**

E°_{cell} for $Zn|Zn^{2+}||Ni^{2+}|Ni$ is positive

$E^\circ = E^\circ_{OP_{Zn}} + E^\circ_{RP_{\text{Ni}}} = 0.76 - 0.23 = 0.53$ V

424 **(d)**

$$\begin{array}{l} \quad Fe^{3+} + 3e \longrightarrow Fe;\ -\Delta G_1^\circ = -0.36 \times F \times 3 \\ \quad \underset{-}{Fe^{2+}} + \underset{-}{2e} \longrightarrow \underset{-}{Fe};\ -\Delta G_2^\circ = -\underset{+}{0}.439 \times F \times 2 \\ \hline Fe^{3+} + e \longrightarrow Fe^{2+};\ -\Delta G^\circ = [3 \times (-0.36F) + 2 \times (0.439F)] \end{array}$$

$\therefore \quad +1 \times F \times E^\circ = [3 \times (-0.36F) + 2 \times (0.439\,F)]$

or $\quad E^\circ = [3 \times (-0.36) + 2 \times 0.439]$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

$0 = E^\circ - \frac{0.059}{n} \log K_{eq}$

Or $\log K_{eq} = \frac{nE^\circ}{0.059}$

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)**

$W_{\text{Na}} = \frac{E \times i \times t}{96500} = \frac{23 \times 5 \times 10 \times 69}{96500}$

$= 0.715$ g

429 **(b)**

$2H_2O \rightleftharpoons 4H^+ + 2O^{2-}$

At cathode, $4H^+ + 4e^- \rightarrow 2H_2$

At anode, $2O^{2-} - 4e^- \rightarrow O_2$

$\therefore$ 4 Faraday of charge liberates = 1 mol

$= 22.4\ dm^3 O_2$

$\therefore$ 1 Faraday of charge liberates

$= \frac{22.4}{4} \times 1$

$= 5.6\ dm^3 O_2$

430 **(c)**

$\frac{2}{3}(Al^{3+})_2 + 4e^- \longrightarrow \frac{2}{3}Al$

or $\frac{2}{3}(O^{2-})_3 \longrightarrow 4e^- + O_2$

$-\Delta G = nFE$ (E is potential required)

$E = -\frac{960 \times 1000}{4 \times 96500} = -2.5$ 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)**
$H_2O \longrightarrow H_2 + 1/2O_2$
$\therefore$ Mole ratio of $H_2O : O_2 :: 1 : 1/2$

435 **(c)**
$E^\circ_{OP} = +\text{Ve}$ thus, $E^\circ_{RP} = -\text{ve}$.

436 **(c)**
Eq. of Al formed $\frac{W}{E} = \frac{it}{96500}$
$\therefore W_{\text{Al}} = \frac{4 \times 10^4 \times 6 \times 60 \times 60 \times 27}{96500} \left(E = \frac{27}{3}\right)$
$\therefore W_{\text{Al}} = 8.05 \times 10^4$ g

437 **(c)**
H_2 undergoes oxidation and AgCl(Ag^+) undergoes reduction

438 **(c)**
The ions which lies below H^+ in electrochemical series, displace H^+ from solution.
$\because$ Cu^{2+} lies below H^+ in electrochemical series.
$\therefore$ 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 × Faraday.

441 **(d)**
$1F = N \times e = E \text{ g} = 96500 \text{ C}$

442 **(b)**
Standard electrode potential is measured by voltmeter.

443 **(a)**
$Cl_2 + 2KI \longrightarrow 2KCl + I_2$; 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 $(i) = 0.25 \text{ mA} = 0.025 \text{ A}$
Time $(t) = 60$ s
$Q = i \times t = 60 \times 0.025 = 1.5 \text{ C}$
No. of electrons $= \frac{1.5 \times 6.023 \times 10^{23}}{96500}$
$e^- = 9.36 \times 10^{18}$
$Ca \rightarrow Ca^{2+} + 2e^-$
$2e^-$ are required to deposite one Ca atom
$9.36 \times 10^{18}\ e^-$ will be used to deposite
$= \frac{9.36 \times 10^{18}}{2} = 4.68 \times 10^{18}$

446 **(b)**
If F is Faraday and N is Avogadro number, charge of
electron $= \frac{F}{N}$

447 **(b)**
Half-cell reaction is

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$$

$$E_{Cr_2O_7^{2-}/Cr^{3+}} - E^\circ_{Cr_2O_7^{2-}/Cr^{3+}}$$

$$= -\frac{0.0591}{n} \log \frac{[Cr^{3+}]^2}{[Cr_2O_7^{2-}][H^+]^{14}}$$

$$= \frac{0.0591}{6} \log[10^{-3}]^{14} = -0.414 \text{ V}$$

448 **(d)**
$2Fe + \frac{3}{2}O_2 \longrightarrow \underset{\text{Brown}}{Fe_2O_3}$
$Cu + CO_2 + H_2O \longrightarrow CuCO_3 . \underset{\text{Green}}{Cu} (OH)_2$
$Ag + H_2S \longrightarrow \underset{\text{Black}}{Ag_2S}$

450 **(c)**
$Cu^{2+}_{aq} + e \longrightarrow Cu^+_{aq}$; $\Delta G^\circ_1 = -[1 \times 0.15 \times F]$
$\underline{Cu^{2+}_{aq} + e \longrightarrow Cu}$; $\Delta G^\circ_2 = -[1 \times 0.50 \times F]$
On adding
$\therefore Cu^{2+}_{aq} + 2e \longrightarrow Cu$; $\Delta G^\circ_3 = -[2 \times E^\circ_3 \times F]$
$\therefore 2E^\circ_3 = 0.65 \text{ V}$ or $E^\circ_3 = 0.325 \text{ V}$

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)**

$$E_{\text{cell}} = E^{\circ}_{OP_{\text{H}}} + E^{\circ}_{RP_{\text{H}}} + \frac{0.059}{2} \log \frac{[\text{H}^+]^2_{\text{RHE}} P_1}{[\text{H}^+]^2_{\text{LHE}} P_2}$$

$$= \frac{0.059}{2} \log \frac{(10^{-3})^2}{(10^{-2})^2} \left(E^{\circ}_{OP} = E^{\circ}_{RP}\right)$$

$$= -\,0.059 \text{ V}$$

Thus, cell reaction is non-spontaneous.

453 **(c)**

From Kohlrausch's law

$\Lambda^{\infty}_{m} = \text{v}_{+}\,\lambda^{\infty}_{+} + \text{v}_{-}\,\lambda^{\infty}_{-}$

For CaCl_2

$\Lambda^{\infty}_{\text{m}}(\text{CaCl}_2) = \Lambda^{\infty}_{\text{Ca}^{2+}} + 2\Lambda^{\infty}_{\text{Cl}^-}$

$= 118.88 \times 10^{-4} + 2 \times 77.33 \times 10^{-4}$

$= 118.88 \times 10^{-4} + 154.66 \times 10^{-4}$

$= 273.54 \times 10^{-4}$ m^2 mho mol^{-1}

455 **(b)**

In case of very dilute solution of NaCl, electrolysis brings in the following changes :

Anode : $2\text{OH}^- \longrightarrow \text{H}_2\text{O} + \frac{1}{2}\text{O}_2 + 2e$

Cathode : $2\text{H}^+ + 2e \longrightarrow \text{H}_2$

456 **(b)**

Mg lie above Cu in electrochemical series and hence, Cu electrode acts as cathode

$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{Cu}^{2+}/\,\text{Cu}} = E^{\circ}_{\text{Mg}^{2+}/\,\text{Mg}}$

$\therefore$ 2.70 V $= 0.34 - E^{\circ}_{\text{Mg}^{2+}/\text{Mg}}$

$\therefore E^{\circ}_{\text{Mg}^{2+}/\,\text{Mg}} = -2.36$ 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

Equivalent weight of Ag $= \frac{108}{1} = 108$

Equivalent weight of Cu $= \frac{63.6}{2} = 31.8$

$$\frac{w_1}{w_2} = \frac{E_1}{E_2}$$

$$\therefore \frac{10.79}{w_{\text{Cu}}} = \frac{108}{31.8}$$

Or $w_{\text{Cu}} = \frac{10.79 \times 31.8}{108} = 3.2$ g

459 **(b)**

Energy = charge × potential

= 1 × 100 × 115 J

= 11.5 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

$$\underset{P1}{\text{H}_2(\text{g})} \rightarrow 2\text{H}^+ \,(1\text{ M}) + 2e^-$$

RHS half cell

$$2\text{H}^+ \,(1\text{ M}) + 2e^- \rightarrow \underset{P2}{\text{H}_2\,(\text{g})}$$

$$\underset{\text{P1}}{\text{H}_2(\text{g})} \rightarrow \underset{\text{p2}}{\text{H}_2\,(\text{g})}$$

$$E^{\circ}_{\text{cell}} = 0.00 \text{ V}, K = \frac{p_2}{p_1}, n = 2$$

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{RT}{nF} \log_{\text{e}} K$$

$$= 0 - \frac{RT}{2F} \log_{\text{e}} \frac{P_2}{P_1}$$

$$E_{cell} = \frac{RT}{2F} \log_{e} \frac{p_1}{p_2}$$

463 **(b)**

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

96500 C charge liberates silver = 108 g

$\therefore$ 96500 C will liberate silver = 10.8 g

464 **(b)**

Given 125mL of 1 M AgNO_3solution. It means that

$\because$ 1000mL of AgNO_3 solution contains

= 108 g Ag

$\therefore$ 125mL of AgNO_3 solution contains

$$= \frac{108 \times 125}{1000} \text{ gAg}$$

= 13.5 g Ag

$\because$ 108 g of Ag is deposited by 96500 C

$\therefore$ 13.5 g of Ag is deposited by

$$= \frac{96500}{108} \times 13.5$$

= 12062.5 C

$Q = it$

Or $t = \frac{Q}{i} = \frac{12062.5}{241.25} = 50$

465 **(c)**

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0591}{n} log\, Q$$

$$\underset{0.1\text{ M}}{\text{Cu}^{2+}} + \text{Zn} \rightarrow \underset{1\text{ M}}{\text{Zn}^{2+}} + \text{Cu}$$

$$Q = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} = \frac{1}{0.1} = 10$$

$$E_{\text{cell}} = 1.10 - \frac{0.0591}{2} \log 10$$

= 1.10 – 0.0295

= 1.0705 V

466 **(a)**

$(126 \text{ Scm}^2 \text{ mol}^{-1})\Lambda^{\infty}_{\text{NaCl}} = \Lambda^{\infty}_{\text{Na}^+} + \Lambda^{\infty}_{\text{Cl}^-}$ ...(i)

$(152 \text{ Scm}^2 \text{ mol}^{-1})\Lambda^{\infty}_{\text{KBr}} = \Lambda^{\infty}_{\text{K}^+} + \Lambda^{\infty}_{\text{Br}^-}$ ...(ii)

$(150\ \text{Scm}^2\ \text{mol}^{-1})\Lambda^{\infty}_{\text{KCl}} = \Lambda^{\infty}_{\text{K}^+} + \Lambda^{\infty}_{\text{Cl}^-}$ ...(iii)

By Eqs. (i) + (ii) – (iii)

$\because\ \Lambda^{\infty}_{\text{NaBr}} = \Lambda^{\infty}_{\text{Na}^+} + \Lambda^{\infty}_{\text{Br}^-}$

$= 126 + 152 - 150$

$= 128\ \text{Scm}^2\ \text{mol}^{-1}$

467 **(b)**

$\underset{\text{anode}}{\text{Ag / Ag}^+\ (0.1\ \text{M})}\ ||\ \underset{\text{cathode}}{\text{Ag}^+\ (1\ \text{M}) / \text{Ag}}$

$E_{cell} = \frac{2.303RT}{nF}\log\frac{C_1}{C_2}$

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

$\therefore\quad E_{cell} = \frac{0.059}{1}\log\frac{1}{0.1}$

$E_{cell} = 0.059\ \text{V}$

468 **(a)**

At cathode $2\text{H}^+ + 2e^- \rightarrow \text{H}_2$

At anode $2\text{OH}^- \rightarrow \text{H}_2\text{O} + \frac{1}{2}\text{O}_2 + 2e^-$

469 **(c)**

$E^{\circ}_{\text{cell}} = E^{\circ}_{OP_{\text{Sn}}} + E^{\circ}_{RP_{\text{Fe}}} = 0.14 + 0.77 = 0.91\ \text{V}$

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)**

Conductance $= \frac{1}{R} = \frac{1}{210}$

$= 4.76 \times 10^{-3}$ mho.

472 **(c)**

According to Faraday law's

$m = Z \times Q = \frac{E \times Q}{96500}$

$m = \frac{108}{96500} \times 2 \times 96500 = 216\ g$

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)**

$E^{\circ}_{Cu^{2+}/Cu} = +0.34\ \text{V} = E^{\circ}_{RP}$

(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)**

$\lambda^{\infty}_{\text{K}^+} = \mu^{\infty}_{\text{K}^+} \times \text{F}$

or $\mu^{\infty}_{\text{K}^+} = \frac{64.35}{96500}$

$= 6.67 \times 10^{-4}\ \text{cm}^2\text{sec}^{-1}\text{volt}^{-1}$

478 **(a)**

Zn has more +ve E°_{OP} and thus possesses more tendency to get oxidise and act as reducing agent.

479 **(a)**

In Danial 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.

$\because$ Reduction potential of Zn is more negative than Cu.

$\therefore$ Zn | Zn^{2+} is anode and Cu^{2+} | Cu is cathode.

480 **(d)**

To prevent rusting or oxidation of Fe by the action of O_2 in presence of H^+, galvanization of iron object is made.

481 **(b)**

It is definition of equivalent conductivity.

482 **(a)**

$\Delta G^{\circ} = -nE^{\circ}F$ (for $\text{Cu} + 2\text{Ag}^+ \longrightarrow \text{Cu}^{2+} + 2\text{Ag}$) $n = 2$

$= -2 \times 0.46 \times 96500$

$= -88788\ \text{J} = -89.0\ \text{kJ}$

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

$\Delta G_1{}^\circ = -5(1.51)F = -7.55F$

$MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O \;\; E^\circ = 1.23\ V$

$\Delta G_2{}^\circ = -2(1.23)F = -2.46\ V$

On substrating

$MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O\ \Delta G_3{}^\circ = -5.09\ F$

$$E^\circ_{MnO_4^- | MnO_2} = \frac{\Delta G_3{}^\circ}{-nF} = \frac{-5.09\ F}{-3F} = 1.70\ V$$

488 **(c)**
Salt bridge is used to remove or eliminate liquid junction potential arised 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 $AgnO_3$ is added to solution , KCl will be displaced according to following reaction
$AgNO_3\ (aq) + KCl\ (aq) \longrightarrow AgCl_{(s)} + KNO_3\ (aq)$
For every mole of KCl displaced from solution, one mole of KNO_3 comes in solution resulting in almost constant conductivity. As the end point is reached, added $AgNO_3$ 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 :
RHS :Reduction: $X + e^- \rightarrow X^-$ ...(i)
LHS :Oxidation: $M \rightarrow M^+ + e^-$ ...(ii)
Add Eqs. (i) an (ii)
$M + X \rightarrow M^+ + X^-$
The cell potential = $-\ 0.11$ V
Since, $E_{cell} = -$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) × ϕ
where, ϕ = volume in cm^3 containing 1 g equivalent of electrolyte
1. g equivalent is dissolve in
= 1000 cm^3

1 g equivalent is dissolve in = $\frac{1000}{0.1}$ cm^3

= 10000cm^3

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
or $\frac{W_{Mg}}{12} = \frac{W_{Al}}{9}$

497 **(c)**
For the change , $2Fe^{3+} + 3I^- \rightleftharpoons 2Fe^{2+} + I_3^-$
$E^\circ_{cell} = 0.77 - 0.54 = 0.23\ V$
$E^\circ = \frac{0.059}{2}\log K_c$
$0.23 = \frac{0.0591}{2}\log K_c$
$K_c = 6.26 \times 10^7$

498 **(c)**
$Zn(s) + 2H^+(aq) \rightleftharpoons Zn^{2+}(aq) + H_2(g)$

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

When H_2SO_4 is added then $[H^+]$ will increase therefore E_{cell} will also increase and equilibrium will shift towards right

499 **(a)**
$2H + 2e \longrightarrow H_2$;
$2OH^- \longrightarrow H_2O + 1/2O_2 + 2e$
$\therefore$ Wt. ratio of $H_2 : O_2 = 2 : 16$ or $1 : 8$

500 **(d)**
Ionic mobilities increase on dilution.

501 **(b)**

Cell is completely discharged, it means equilibrium gets established,

$E_{\text{cell}} = 0$

Zn | Zn^{2+} (1 M) || Cu^{2+} (1 M) | Cu

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,

$E^{\circ}_{\text{cel}l} = E^{\circ}_{\text{cell}} - \frac{0.0591}{n} \log K_{\text{eq}}$

$E^{\circ}_{\text{cell}} = \dfrac{0.0591}{n} \log K_{\text{eq}}$

Or $1.10 = \frac{0.0591}{2} \log K_{\text{eq}}$

$K_{eq} = \dfrac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} = \text{antilog}\ \dfrac{2.20}{0.0591}$

$= \text{antilog } 37.3$

502 **(c)**

$E_{\text{cell}} = E_{\text{OP}} + E_{\text{RP}}$

503 **(a)**

$E^{\circ}_{\text{cell}} = E^{\circ}_{OP_{\text{Zn}}} + E^{\circ}_{RP_{\text{Ag}}} = 0.76 + 0.77 = +1.53$ V.

504 **(c)**

Rust is $\text{Fe}_2\text{O}_3.x\text{H}_2\text{O}$ or mixture of $\text{Fe}_2\text{O}_3 + \text{Fe(OH)}_3$.

505 **(a)**

Metal with − ve. E°_{Rp} or + ve E°_{OP} possesses the tendency to get itself oxidised.

506 **(a)**

$$\Lambda = k \times \frac{1000}{M} = \frac{1}{R} \times \frac{1}{a} \times \frac{1000}{M}$$

$$= \frac{1}{45} \times \frac{2.2}{3.8} \times \frac{1000}{0.5} = 25.73$$

507 **(a)**

$\text{Ionic mobility} = \frac{\text{speed of ions}}{\text{potential gradient}} = \frac{\text{m sec}^{-1}}{\text{volt m}^{-1}}$

509 **(a)**

Charge = $10 \times 1 = 10$ C

Also, 96500 C = 6×10^{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 Aelectrochemical cell that converts the chemical energy of Aspontaneous 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)**

$\Lambda^{\infty}_{\text{AgCl}} = \lambda^{\infty}_{Ag^+} + \Lambda^{\infty}_{\text{Cl}^-}$

$= \lambda^{\infty}_{\text{AgNO}_3} + \lambda^{\infty}_{\text{NaCl}} - \lambda^{\infty}_{\text{NaNO}_3}$

517 **(d)**

These are characteristics of fuel cells.

518 **(a)**

$Q = N.e \text{ or } e = \dfrac{Q}{N} = \dfrac{\text{faraday}}{\text{Av. no.}}$

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

$A \longrightarrow A^+ + e$ LHS shows oxidation.

$B^+ + e \longrightarrow B$ RHS shows reduction.

$A + B^+ \longrightarrow A^+ + B$

521 **(b)**

The laws of electrolysis were given by Faraday.

522 **(a)**

Current i =?

Equivalent weight of Al $= \frac{27}{3} = 9$

$w = Z.i.t$ or $i = \frac{w}{Z.t} = \frac{0.09 \times 96500}{9 \times 96.5}$

$= 10$ 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)**

$$\text{Cell constant} = \frac{k}{\text{C}} = 0.0212 \times 55$$
$$= 1.166\ \text{cm}^{-1}$$

528 **(a)**

$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.059}{n} \log \left[\frac{\text{product}}{\text{reactant}}\right]$

Given , $E^{\circ}_{\text{Cr}^{3+}/\text{Cr}} = -0.74$ V

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

Cr | Cr^{2+} (0.1 M) |Fe^{2+}| (0.01 M) / Fe

∴ Cr^{3+} /Cr is anode and Fe^{2+} / Fe is cathode.

$$E^{\circ}_{\text{cell}} = E^{\circ}_{C} - E^{\circ}_{A}$$
$$= (-0.44) - (0.74)$$
$$= -0.44 + 0.74$$
$$= +0.30\ \text{V}$$

Cell reaction is

$2\text{Cr} + 3\text{Fe}^{2+} \rightarrow 2\text{Cr}^{3+} + 3\text{Fe}$

number of electrons in cell reaction = 6

$$E^{\circ}_{cell} = E^{\circ}_{cell} - \frac{0.059}{n} \log \left[\frac{\text{product}}{\text{reactant}}\right]$$
$$= +0.30\ \text{V} - \frac{0.059}{6} \log \left[\frac{(\text{Cr}^{3+})^2}{(\text{Fe}^{2+})^3}\right]$$
$$= 0.30 - \frac{0.059}{6} \log \left[\frac{(0.1)^2}{(0.01)^3}\right]$$
$$= 0.30 - \frac{0.059}{6} \log 10^4$$
$$= 0.30 - \frac{0.059}{6} \times 0.60$$
$$= 0.30 - 5.9 \times 10^{-3}$$
$$= 0.2941\ \text{V}$$

529 **(c)**

$2e + \text{Cl}^{+} \longrightarrow \text{Cl}^{-};\ \Delta G^{\circ}_{1}$ ....(i)

$2\text{Cl}^{-} \longrightarrow 2\text{Cl}_2 + 2e;\ \Delta G^{\circ}_{2}$ ....(ii)

$2e + 2\text{Cl}^{+} \longrightarrow \text{Cl}_2;\ \Delta G^{\circ}_{3}$

$\therefore\ \Delta G^{\circ}_{3} = 2 \times \Delta G^{\circ}_{1} + \Delta G^{\circ}_{2}$

$-2 \times E^{\circ}_{3} \times \text{F} = -2 \times 2 \times 0.94\, F - 2 \times (-1.36) \times F$

$\therefore E^{\circ}_{3} = 0.52$ V

530 **(a)**

$$E_{\text{cell}} = E^{\circ}_{OP_{\text{H}}} + E^{\circ}_{RP_{\text{H}}} + \frac{0.059}{2} \log \frac{[\text{H}^{+}]^2.P_1}{[\text{H}^{+}]^2.P_2}$$
$$= \frac{0.059}{+2} \log \frac{P_1}{P_2}$$

E_{cell} is +ve when $P_1 > P_2$.

532 **(c)**

Eq. of O_2 = Eq. of Ag or $\frac{1.6}{8} = \frac{W}{108}$;

$\therefore\ W_{\text{Ag}} = 21.6$ g

533 **(c)**

At LHS (oxidation) $2 \times (\text{Ag} \rightarrow \text{Ag}^{+} + e^{-})$,

$E^{\circ}_{\text{oxi}} = -x$

At RHS (reduction)

$\text{Cu}^{2+} + 2e^{-} \rightarrow \text{Cu}, \qquad E^{\circ}_{\text{red}} = +y$

$2\text{Ag} + \text{Cu}^{2+} \rightarrow \text{Cu} + 2\text{Ag}^{+}, E^{\circ}_{\text{cell}} = (y - x)$

Note E° values remain constant when half – cell equation is multiplied / divided.

535 **(a)**

$\Lambda^{\infty}_{eq.} = \lambda^{\infty}_{a} + \lambda^{\infty}_{c} = 315 + 35 = 350.$

536 **(d)**

$\text{Pt}(s), \text{H}_2\ (g)\ |\ \text{H}^{+}\ (1\ M)\ ||\ \text{Ag}^{+}\ (aq)\ |\ \text{Ag}\ (s)$

EMF of cell = 0.62 V, $E^{\circ}_{\text{cell}} = 0.80\ V$

$\text{H}_2 \rightarrow 2\text{H}^{+} + 2\text{e}^{-}$ (at anode)

$2\text{Ag}^{+} + 2\text{e}^{-} \rightarrow 2\text{Ag}$ (at cathode)

$\text{H}_2 + 2\text{Ag}^{+} \rightarrow 2\ \text{Ag} + 2\text{H}^{+}$ (overall reaction)

$$E_{\text{cell}} = E^{\circ} - \frac{2.303RT}{2F} \log \frac{[\text{H}^{+}]^2}{[\text{Ag}^{+}]^2\ [\text{H}_2]}$$
$$E_{\text{cell}} = E^{\circ} - \frac{2.303RT}{2F} \log \frac{1}{[\text{Ag}^{+}]^2}$$
$$0.62 = 0.80 + \frac{2.303 \times 0.06}{2} \log \frac{1}{[\text{Ag}^{+}]^2}$$
$$0.62 = 0.80 + \frac{2 \times 2.303 \times 0.06}{2} \log[\text{Ag}^{+}]$$
$$-0.18 = 0.1382 \log [\text{Ag}^{+}]$$
$$[\text{Ag}^{+}] = 0.05\ \text{m}$$

∴ Mole of Ag^{+} in 100 mL $= 0.05 \times \frac{100}{1000}$

Wt. of Ag^{+} in 100 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)**

$$Q = it$$
$$= 1 \times 60$$
$$= 60\ C$$

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 4.48dm^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

$CH_3COOH + NaOH \longrightarrow CH_3COONa + HCl$

$\Lambda_m^0\ CH_3COOH = \Lambda_m^0\ CH_3COONa + \Lambda_m^0\ HCl - \Lambda_m^0\ NaCl$

$= 91 + 425.9 - 126.4$

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

546 **(b)**

$$E_{cell} = E^{\circ}_{cell} + \frac{0.059}{n} \log[M^{n+}]$$

$$E_{cell} = 0.34 + \frac{0.059}{2} \log 10^{-2}$$

$$= 0.34 + \frac{0.059}{2} \times -2$$

$$= +\ 0.281\ V$$

547 **(a)**

We know that, $\Delta G^{\circ} = -nF . E^{\circ}$

Where, $n = 1$ (number of transferred electron in the cell reaction)

F =96500 C (Faraday's constant)

E° = ? (the standard electrode potential of cell)

ΔG° = - 21.20 kJ (standard free energy)

$\therefore$ -21.20 × 1000 J = - 1 × 96500 × E°

Or $E^{\circ} = \frac{21200}{96500} = 0.220$ V

548 **(b)**

$$-\Delta G = nEF = nE^{\circ}F + RT \ln \frac{[Cu^{2+}]}{[Zn^{2+}]}$$

$\therefore$ ΔG is function of $\ln(c_2/c_1)$.

549 **(d)**

The half reactions are

$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^- \times 2$

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

$2Fe(s) + O_2(g) + 4H^+$

$\rightarrow 2Fe^{2+}(aq) + 2H_2O\ (l)$

$$E = E^{\circ} - \frac{0.059}{4} \log \frac{(10^{-3})^2}{(10^{-3})^4(0.1)} = 1.57\ V$$

550 **(d)**

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

$w = ZQ$

Where, w = amount of metal

$w = 5.12$ kg

$= 5.12 \times 10^3$ g

Z = electrochemical equivalent

$$Z = \frac{\text{equivalent weight}}{96500} = \frac{\text{atomic mass}}{\text{electrons} \times 96500}$$

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

$$5.12 \times 10^3 = \frac{27}{3 \times 96500} \times Q$$

$$Q = \frac{5.12 \times 10^3 \times 3 \times 96500}{27} C$$

$$= 5.49 \times 10^7 C$$

551 **(c)**

E°_{OP} of Mg $> E^{\circ}_{OP}$ of Zn

552 **(b)**

$$E_{cell} = E^{\circ}_{OP_{Cl}} + E^{\circ}_{RP_{Cl}} + \frac{0.059}{2} \log \frac{[Cl^-]^2_{LHE} P_2}{[Cl^-]^2_{RHE} P_1}$$

$$= \frac{0.059}{2} \log \frac{P_2}{P_1}$$

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)**

$t^{\infty}_{K^+} = t^{\infty}_{KCl} \times t_{K^+}$

$130 \times 0.495 = 64.35$

555 **(c)**

$2Fe^{3+} + Zn \rightarrow Zn^{2+} + 2\ Fe^{2+}$

$2Fe^{3+} + 2e^- \rightarrow 2\ Fe^{2+}$

According to Nernst's equation,

$$E^{\circ}_{cell} = E^{\circ}_{cell} - \frac{RT}{nF} \ln \frac{[Fe^{2+}]^2}{[Fe^{3+}]^2}$$

E_{cell} is depend on concentration of Fe^{2+}.

Therefore on increasing the concentration of $[Fe^{2+}] E_{cell}$ decreases.

556 **(b)**

96500 C = 6×10^{23} electrons.

557 **(a)**

Because fluorine is most powerful oxidizing agent than other halogens

558 **(d)**

Equivalent conductivity, $\Lambda_{eq} = \frac{\kappa \times 1000}{\text{normality}}$

$= ohm^{-1}\ cm^2\ (g - equiv)^{-1}$

559 **(b)**

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

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

$$\log K_c = \frac{0.47 \times 2}{0.059}$$

$$K_c = 8.5 \times 10^{15}$$

560 **(d)**

Mg will not deposit on cathode during electrolysis.

561 **(a)**

$\Lambda^{\infty}_{\mathrm{NH_4OH}} = \Lambda^{\infty}_{\mathrm{NH_4Cl}} + \Lambda^{\infty}_{\mathrm{NaOH}} - \lambda^{\infty}_{\mathrm{NaCl}}$

562 **(b)**

At cathode, cations are reduced. Also discharge potential of $\mathrm{H^+}$ is less than $\mathrm{Na^+}$.

563 **(a)**

Conductance of a solution = conductance of ions present in solution = $\sum \mathrm{C}_i + \mathrm{C_{water}}$.

564 **(d)**

$\mathring{\Lambda}_{\mathrm{NaBr}} = \mathring{\Lambda}_{\mathrm{NaCl}} + \mathring{\Lambda}_{\mathrm{KBr}} - \mathring{\Lambda}_{\mathrm{KCl}}$

565 **(d)**

$\mathrm{Hg_2Cl_2}(s) + 2e \longrightarrow 2\mathrm{Hg}\ (l) + 2\mathrm{Cl^-}2(aq)$.

566 **(b)**

Zn acts as anode and gets oxidised. At graphite cathode the reaction is ,

$\mathrm{MnO_2 + NH_4^+} + e \longrightarrow \mathrm{Mn(OH)O + NH_3}$

567 **(a)**

Eq. of $\mathrm{H_2} = 8/1 = 8$

∴ 8 faraday are needed.

568 **(d)**

$\mathrm{C_6H_5NO_2\ + 6H^+ +\ 6e^- \rightarrow C_6H_5NH_2 +\ 2H_2O}$

1 mol = 123 g nitrobenzene requires 6 mol electrons

= 6 × 96500 C charge

∴12.3 g nitrobenzene will require $= \frac{6\times96500\times12.3}{123}$

= 6 × 9650 = 57900 C

569 **(c)**

At cathode,

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

$E_{\mathrm{Al}} = \frac{27}{3} = 9$

$w_{\mathrm{Al}} = E_{\mathrm{Al}} \times$ no.of faradays

= 9 × 0.1 = 0.9 g

570 **(a)**

$2\mathrm{H^+}\ + 2e \longrightarrow \mathrm{H_2}$;

$2\mathrm{OH^-} \longrightarrow \mathrm{H_2O} + 1/2\mathrm{O_2} + 2e$

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 ($\mathrm{H_2SO_4}$) used in lead storage battery.

576 **(b)**

It is the definition of conductivity.

577 **(a)**

By Kohlrausch's law

$\overset{\circ}{\Lambda}_{\mathrm{NaBr}} = \overset{\circ}{\Lambda}_{\mathrm{NaCl}} + \overset{\circ}{\Lambda}_{\mathrm{KBr}} - \overset{\circ}{\Lambda}_{\mathrm{KCl}}$

= 126 + 152 – 150

= 128 $\mathrm{S\ cm^2 mol^{-1}}$

578 **(d)**

$\Lambda\ vs\ \sqrt{c}$ 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)**

$\mathrm{Fe \rightarrow Fe^{2+} +\ 2e^-}$ (anode reaction)

$\mathrm{O_2 +\ 2H_2O} + 4e^- \longrightarrow\ 4\mathrm{OH^-}$ (cathode reaction)

The overall reaction is

$\mathrm{2Fe + O_2 +\ 2H_2O \rightarrow 2Fe(OH)_2}$

$\mathrm{Fe(OH)_2}$ may be dehydrated to iron oxide FeO, or further oxidized to $\mathrm{Fe(OH)_3}$ and then dehydrated to iron rust, $\mathrm{Fe_2O_3}$.

584 **(b)**

$\mathrm{Ag^+}\ + e \longrightarrow \mathrm{Ag}$

585 **(b)**

E°_{OP} of Fe > E°_{OP} of Cu;

Thus, Fe gets oxidised or Fe $\longrightarrow \mathrm{Fe^{2+}} + 2e$;

$\mathrm{Cu^{2+}} + 2e \longrightarrow \mathrm{Cu}$

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)**

$k = \frac{1}{R} \times \frac{1}{a} = \frac{1}{210} \times 0.66$

$= 3.14 \times 10^{-3}$ 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 K $[Au(CN)_2]$ is used.

592 **(c)**

$Q = i \times t = 100 \times 10^{-3} \times 30 \times 60 = 180$ C

594 **(a)**

Anode reaction $H_2(p_1) \rightarrow 2H^+$

Cathode reaction $2H^+ \rightarrow H_2(p_2)$

$$E_{\text{cathode}} = -\frac{RT}{2F} \ln \frac{p_2}{[H^+]^2}$$

$$E_{\text{anode}} = -\frac{RT}{2F} \ln \frac{[H^+]^2}{p_1}$$

$$E_{\text{cell}} = E_{\text{anode}} + E_{\text{cathode}}$$

$$= -\frac{RT}{2F} \ln \frac{[H^+]^2}{p_1} - \frac{RT}{2F} \ln \frac{p_2}{[H^+]^2}$$

$$= -\frac{RT}{2F} \ln \frac{p_2}{p_1} = \frac{RT}{2F} \ln \frac{p_1}{p_2}$$

595 **(b)**

Weight of Ag required $= 80 \times 5 \times 10^{-3} \times 1.05$(wt. $= v \times d$)

$= 0.42$ g

$\because \quad W = \frac{E\,i\,t}{96500}$

$\therefore \quad 0.42 = \frac{108 \times 3 \times t}{96500}$

$\therefore \quad t = 125$ sec

596 **(c)**

$2H^+ + 2e \longrightarrow H_2$ (cathode)

$2OH^- \longrightarrow H_2O + (1/2)O_2 + 2e$ (Anode)

597 **(b)**

$W = Z \times i \times t$

$\therefore W = 0.0011180 \times 0.5 \times 200 = 0.11180$ g

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)**

$C_5H_{12} + 8O_2 \longrightarrow 5CO_2 + 6H_2O$

$\Delta G^\circ = 5 \times G^\circ CO_2 + 6 \times G^\circ H_2O - G^\circ_{C_5H_{12}} - 8 \times G^\circ_{O_2}$

$= 5 \times (-394.4) + 6 \times (-237.2) + 8.2$

$= -3387$ kJ

$\Delta G^\circ = nE^\circ F$

$3387 \times 10^3 = 32 \times E^\circ \times 96500 \; [(C^{-2.4})_5 \longrightarrow$ $5C4{+}{+}32e$

$\therefore E^\circ = 1.0968$ V $\qquad n = 32$

601 **(c)**

Hydrolysis of water can be represented by the following equations

$2H_2O \rightleftharpoons O_2 + 4H^+ + 4e^-$

$\because$ 4 Faraday of charge liberate $O_2 = 32$ g

$\therefore$ 1 Faraday of charge liberate $O_2 = \frac{32}{4}$ g

$= 8$ g

602 **(a)**

Given, $\frac{l}{a} = 0.5$ cm^{-1}

$R = 50\Omega$

$N = 1.0$

Specific conductance $(\kappa) = \frac{1}{\rho} = \frac{l}{R.a} = \frac{0.5}{50}$

$\Lambda = \kappa \times \frac{1000}{N}$

$= \frac{0.5}{50} \times \frac{1000}{0.1}$

$= 10\ \Omega^{-1}$ cm^2 g eq^{-1}

603 **(d)**

The increasing order of deposition of cations at the cathode is

$Cu^{2+} < Ag^+ < Au^{3+}$

$E \propto Z$

$Ag^+ + e^- \rightarrow Ag$

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

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

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)**

$E^\circ_{\text{cell}} = E^\circ_{OP_{\text{Mg}}} + E^\circ_{RP_{\text{Cu}}}$

$= 2.37 + 0.34 = 2.71$ V.

605 **(b)**

Given,

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

$Cu^{2+} \rightarrow Cu, \quad E^\circ = 0.34\,V$

$Ag^+ \rightarrow Ag, \quad E^\circ = 0.8\,V$

Cell reaction of (I) is

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

$E^\circ_{\text{cell}} = E_{\text{oxidation}} + E_{\text{reduction}}$

$= +0.76 + 0.34$

$= +1.10$ V

Cell reaction of (II) is

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

$E^\circ_{cell} = 0.76 + 0.8$

$= +1.56$ V

Cell reaction of (III) is

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

$E^\circ_{cell} = -0.34 + 0.8$

$= +0.46$

So, the correct order of E°_{cell} of these cell is

II > I > III.

606 **(c)**

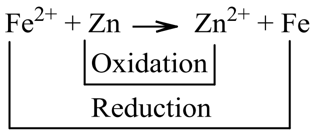
The equivalent conductance of strong electrolyte is increased by dilution, because its value is equal to the multiple of K_v and the volume of solution. By making dilution the volume of solution increases which also increase the value of equivalent conductivity.

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,



$\text{EMF} = E_{cathod} - E_{anode}$

$= -0.41 - (-0.76)$

$= +0.35$ V

609 **(a)**

H_2 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 $Li^+_{aq} > Na^+_{aq} > K^+_{aq} > Rb^+_{aq}$. Larger is ion, lesser is its mobility.

612 **(d)**

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

$\therefore\ 0.5105 = E^\circ_{cell} + \frac{0.059}{2} \log \frac{1}{1}$ or $E^\circ_{cell} = 0.5105$ 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

$\Delta G^\circ = nFE^\circ_{cell}$

and $\Delta G^\circ = -2.303\, RT \log K_{eq}$

616 **(a)**

Anode : $Fe(s) + 2OH^- \longrightarrow FeO(s) + H_2O(l) + 2e$

Cathode : $Ni_2O_3 + H_2O(l) + 2e \longrightarrow 2NiO(s) + 2OH^-$

$$E_{cell} = E^\circ_{OP_{Fe/FeO}} - \frac{0.059}{2} \log_{10} \frac{[H_2O]}{[OH^-]^2} + E^\circ_{RP_{Ni_2O_3/NiO}} + \frac{0.059}{2} \log_{10} \frac{[H_2O]}{[OH^-]^2}$$

$$E_{cell} = E^\circ_{OP_{Fe/FeO}} + E^\circ_{RP_{Ni_2O_3/NiO}}$$

617 **(b)**

$K_a = c\alpha^2$ for weak acid;

$HA \rightleftharpoons H^+ + A^-$

618 **(d)**

C is also calomel electrode.

619 **(c)**

$CuSO_4 \rightleftharpoons Cu^{2+} + SO_4^{2-}$

At cathode $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

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

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

$\therefore \frac{w_{Al}}{9} = \frac{w_{Ag}}{108}$ or $\frac{w_{Al}}{w_{Ag}} = \frac{9}{108}$

621 **(d)**

$W \propto E$; if Q = constant.

622 **(a)**

From Asolution 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

$$\frac{w_1}{w_2} = \frac{E_1}{E_2}$$

$$\therefore \quad \frac{1.08}{x} = \frac{108}{1}$$

$\therefore$ 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,

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{2.303\,RT}{nF}\log_{10}K$$

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

Given that,

$\therefore R = 8.315 \text{ JK}^{-1} \text{ mol}^{-1}$

$T = 25\,°\text{C} + 273 = 298K$

$F = 96500\,C$ and $n = 2$

$$\therefore \quad E^\circ_{\text{cell}} = \frac{2.303 \times 8.314 \times 298}{2 \times 96500}\log_{10}K$$

$$= \frac{0.0591}{2}\log_{10}K$$

$\because$ Given that $E^\circ_{cell} = 0.295$ V

$$\therefore \quad 0.295 = \frac{0.0591}{2}\log_{10}K$$

$$\log_{10}K = \frac{0.295 \times 2}{0.0591} = 10$$

$$\log_{10}K = \text{antilog } 10$$

$$K = 1 \times 10^{10}$$

626 **(a)**

Follow E°_{OP} values in electrochemical series.

627 **(c)**

$$\mathring{\Lambda}_{\text{CH}_3\text{COOH}} = \mathring{\Lambda}_{\text{CH}_3\text{COONa}} + \mathring{\Lambda}_{\text{HCl}} - \mathring{\Lambda}_{\text{NaCl}}$$

$$= 91.0 + 426.2 - 126.5$$

$$= 390.7\ S \text{ cm}^2 \text{ mol}^{-1}$$

628 **(a)**

Fe is more electropositive than copper

Hence, Cu^{2+} can oxidize Fe

629 **(a)**

Eq. of Al = Eq. of Na;

$$\therefore \frac{1.8}{27/3} = \frac{W}{23}$$

$\therefore W_{Na} = 4.6$ 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,

$$[\therefore \frac{W}{E} \text{ (for Ag)} = \frac{W}{E} \text{ (for Al)}]$$

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