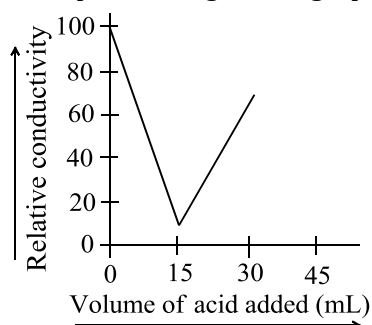


## 3.ELECTROCHEMISTRY

## Single Correct Answer Type

- The standard reduction potential values of three metallic cations, X, Y, and Z are 0.52,  $-3.03$ , and  $-1.18$  V, respectively. The order of reducing power of the corresponding metal is  
a)  $Y > Z > X$                       b)  $X > Y > Z$                       c)  $Z > Y > X$                       d)  $Z > X > Y$
- What is the value of  $pK_b(\text{CH}_3\text{COO}^-)$  if  $\Lambda_m^\circ = 390 \text{ S cm}^{-1} \text{ mol}^{-1}$  and  $\Lambda_m = 7.8 \text{ S cm}^2 \text{ mol}^{-1}$  for 0.04 M of  $\text{CH}_3\text{COOH}$  at  $25^\circ\text{C}$ ?  
a) 9.3                                      b) 9.2                                      c) 4.7                                      d) 4.8
- 20 mL of KOH solution was titrated with 0.20 M  $\text{H}_2\text{SO}_4$  solution in a conductivity cell. The data obtained were plotted to give the graph shown below



The concentration of the KOH solution was

- 0.30 mol  $\text{L}^{-1}$                       b) 0.15 mol  $\text{L}^{-1}$                       c) 0.12 mol  $\text{L}^{-1}$                       d) 0.075 mol  $\text{L}^{-1}$
- The resistance of 1 N solution of acetic acid 250 ohm, when measured in a cell of cell constant  $1.15 \text{ cm}^{-1}$ . The equivalent conductance (in  $\text{ohm}^{-1} \text{ cm}^2 \text{ eq}^{-1}$ ) of 1 N acetic acid is  
a) 4.6                                      b) 9.2                                      c) 18.4                                      d) 0.023
  - The two Pt electrodes fitted in a conductance cell are 1.5 cm apart while the crosssectional area of each electrode is 0.75 cm. What is the cell constant?  
a) 1.125                                      b) 0.5 cm                                      c)  $2.0 \text{ cm}^{-1}$                                       d)  $0.2 \text{ cm}^{-1}$
  - $$\text{I}_2 + 2e^- \longrightarrow 2\text{I}^\ominus; \quad E^\ominus = 0.54 \text{ V}$$

Given that:  $\text{Br}_2 + 2e^- \longrightarrow 2\text{Br}^\ominus; \quad E^\ominus = 1.69 \text{ V}$

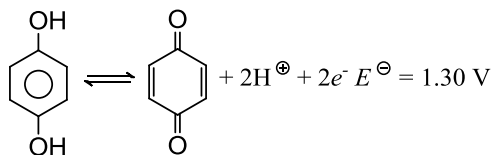
Predict which of the following is true

    - $\text{I}^\ominus$  ions will be able to reduce bromine
    - $\text{Br}^\ominus$  ions will be able to reduce iodine
    - Iodine will be able to reduce bromine
    - Bromine will be able to reduce iodide ions
  - What is the potential of the cell containing two hydrogen electrodes as represented below?  

$$\text{Pt}; \frac{1}{2} \text{H}_2(\text{g}) \left| \text{H}_2\text{O} \right| \left| \text{H}^\oplus(0.001 \text{ M}) \right| \frac{1}{2} \text{H}_2(\text{g}) \text{ Pt}$$
    - $-0.236 \text{ V}$                                       b)  $-0.0591 \text{ V}$                                       c)  $0.236 \text{ V}$                                       d)  $0.0591 \text{ V}$
  - The standard EMF of quinhydrone electrode is 0.699 V. The EMF of the quinhydrone electrode dipped in a solution with  $\text{pH} = 10$  is  
a) 0.109 V                                      b)  $-0.109 \text{ V}$                                       c) 1.289 V                                      d)  $-1.289 \text{ V}$
  - The potential of a hydrogen electrode in a solution with  $\text{pOH} = 4$  at  $25^\circ\text{C}$  is  
a)  $-0.059$                                       b) 0.059                                      c)  $-0.59$                                       d) 0.59
  - Faraday's laws of electrolysis are related to the  
a) The atomic number of the cation  
b) The atomic number of anion  
c) The equivalent weight of the electrolyte  
d) The speed of the cation
  - 500 mL of 1 N solution of  $\text{CuCl}_2$  was electrolyzed with a current of 2 amperes for 1 hours. What is the normality of the remaining  $\text{CuCl}_2$  solution?

- a) 0.85                                      b) 0.15                                      c) 0.30                                      d) 1.0
12. A solution of 0.65 A flowing for 10 min deposits 3.0 g of the metal which is monovalent. The atomic mass of the metal is  
 a) 10                                      b) 50                                      c) 30                                      d) 96.5
13. For the electrolytic production of  $\text{NaClO}_4$  from  $\text{NaClO}_3$  according to the reaction  $\text{NaClO}_3 + \text{H}_2\text{O} \rightarrow \text{NaClO}_4 + \text{H}_2$ . How many faradays of electricity would be required to produce 0.5 mole of  $\text{NaClO}_4$ ?  
 a) 1                                      b) 2                                      c) 3                                      d) 1.5

14. For the half cell



At pH = 2, the electrode potential is

- a) 1.36 V                                      b) 1.30 V                                      c) 1.42 V                                      d) 1.20 V
15. When electricity is passed through a solution of  $\text{AlCl}_3$  and 13.5 g of Al is deposited, the number of Faraday must be  
 a) 0.5                                      b) 1.0                                      c) 1.5                                      d) 2.0
16. Chromium plating can involve the electrolysis of an electrolyte of an acidified mixture of chromic acid and chromium sulphate. If during electrolysis the article being plated increases in mass by 2.6 g and  $0.6 \text{ dm}^3$  of oxygen are evolved at an inert anode, the oxidation state of chromium ions being discharged must be: (assuming atomic weight of Cr = 52 and 1 mole of gas at room temperature and pressure occupies a volume of  $24 \text{ dm}^3$ )  
 a) -1                                      b) Zero                                      c) +1                                      d) +2

17.  $E^\ominus_{\text{Ag}^\oplus | \text{Ag}} = +0.80 \text{ V}$ ;  $E^\ominus_{\text{Co}^{2+} | \text{Co}} = -0.28 \text{ V}$ ,

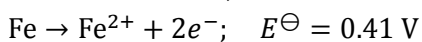
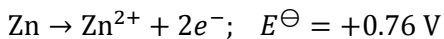
Given  $E^\ominus_{\text{Cu}^{2+} | \text{Cu}} = +0.34 \text{ V}$ ,  $E^\ominus_{\text{Zn}^{2+} | \text{Zn}} = -0.76 \text{ V}$

Which metal will corrode fastest?

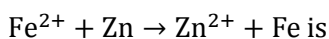
- a) Ag                                      b) Cu                                      c) Co                                      d) Zn
18. The volume of gases evolved at STP by passing 0.2 A of current for 965 s through an aqueous solution of sodium fumarate is  
 a) 22.4 mL                                      b) 11.2 mL                                      c) 89.6 mL                                      d) 44.8 mL
19. Two platinum electrodes were immersed in a solution of  $\text{CuSO}_4$  and electric current was passed through the solution. After some time, it was found that colour of  $\text{CuSO}_4$  disappeared with evolution of gas at the electrode. The colorless solution contains  
 a) Platinum sulphate                      b) Copper hydroxide                      c) Copper sulphate                      d) Sulphuric acid
20. Red hot carbon will remove oxygen from the oxides XO and YO but not from ZO. Y will remove oxygen from XO. Use this evidence to deduce the order of activity of the three metals X, Y, and Z, putting the most reactive first  
 a) X, Y, Z                                      b) Z, Y, X                                      c) Y, X, Z                                      d) Z, X, Y
21.  $\text{Zn} | \text{Zn}^{2+} (\text{A} = 0.1 \text{ M}) || \text{Fe}^{2+} (\text{A} = 0.01 \text{ M}) | \text{Fe}$ .

The emf of the above cell is 0.2905 V. Equilibrium constant for the cell reaction is

- a)  $10^{0.32/0.0591}$                                       b)  $10^{0.32/0.0295}$                                       c)  $10^{0.26/0.0295}$                                       d)  $e^{0.32/0.0295}$
22. Chromium plating is applied by electrolysis to objects suspended in a dichromate solution, according to the following (unbalanced) half reaction:  
 $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + e^- + \text{H}^\oplus(\text{aq}) \rightarrow \text{Cr}(\text{s}) + \text{H}_2\text{O}(\text{l})$   
 How many hours would it take to apply a chromium plating of thickness  $2.0 \times 10^{-2} \text{ mm}$  to a car bumper of surface area  $0.25 \text{ m}^2$  in an electrolysis cell carrying a current of 75.0 A? [Density of chromium is  $7.19 \text{ g cm}^{-3}$ ]  
 a) 2.2 h                                      b) 1.5 h                                      c) 3.0 h                                      d) 0.25 h
23. The standard reduction potential  $E^\ominus$  for the half reactions are follows:



The EMF for the cell reaction

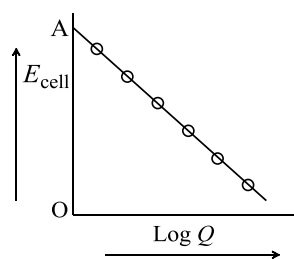


- a)  $-0.35 \text{ V}$                       b)  $+0.35 \text{ V}$                       c)  $+1.17 \text{ V}$                       d)  $-1.17 \text{ V}$

24. Which of the following statements is correct?  
 a) Specific conductance of a solution decreases with dilution, whereas molar conductance increases with dilution  
 b) Specific conductance of a solution increases with dilution, whereas molar conductance decreases with dilution  
 c) Both specific conductance and molar conductance decrease with dilution  
 d) Both specific conductance and molar conductance increase with dilution
25. The volume of gases evolved at STP by passing 0.1 A of current for 965 s, through an aqueous solution of potassium acetate  
 a) 22.4 mL                      b) 11.2 mL                      c) 89.6 mL                      d) 44.8 mL
26. The units of conductivity of the solution are  
 a)  $\text{ohm}^{-1}$                       b) ohms                      c)  $\text{ohm}^{-1}\text{cm}^{-1}$                       d)  $\text{ohm}^{-1} \text{eq}^{-1}$
27. Rust is a mixture of  
 a) FeO and Fe(OH)<sub>2</sub>                      b) FeO and Fe(OH)<sub>3</sub>                      c) Fe<sub>2</sub>O<sub>3</sub> and Fe(OH)<sub>3</sub>                      d) Fe<sub>3</sub>O<sub>4</sub> and Fe(OH)<sub>3</sub>
28. How many coulombs are required for the oxidation of 1 mol of H<sub>2</sub>O to O<sub>2</sub>?  
 a) 93000 C                      b)  $1.93 \times 10^5 \text{ C}$                       c)  $9.65 \times 10^4 \text{ C}$                       d)  $19.3 \times 10^2 \text{ C}$
29. The  $E_{\text{cell}}$  for a given a cell is 1.2346 and 1.2340 V at 300 K and 310 K, respectively. Calculate the change in entropy during the cell reaction if the redox change involves three electrons  
 a)  $-17.37 \text{ J K}^{-1}$                       b)  $+17.37 \text{ J K}^{-1}$                       c)  $173.7 \text{ J K}^{-1}$                       d)  $5.79 \text{ J K}^{-1}$
30. Which one is wrong if electrolysis of CH<sub>3</sub>COONa (aq) is made using Pt electrodes?  
 a) pH of solution increases  
 b) Molar ratio of gases at anode and cathode is 3:1  
 c) [CH<sub>3</sub>COO<sup>⊖</sup>] in solution decreases  
 d) The molar ratio of gases at anode and cathode is 2:1
31. For the given cell  $\text{Pt} | \text{D}_2 | \text{D}^{\oplus} || \text{H}^{\oplus} | \text{Pt} | \text{H}_2$ , if  $E^{\ominus}_{\text{D}_2 | \text{D}^{\oplus}} = 0.003 \text{ V}$ , what will be the ratio of  $\text{D}^{\oplus}$  and  $\text{H}^{\oplus}$  at 25°C when the reaction:  $\text{D}_2 + 2\text{H}^{\oplus} \longrightarrow 2\text{D}^{\oplus} + \text{H}_2$  attains equilibrium  
 a) 1.34                      b) 1.24                      c) 1.124                      d) 1.45
32. A cell constant is generally found by measuring the conductivity of aqueous solution of  
 a) BaCl<sub>2</sub>                      b) KCl                      c) NaCl                      d) MgCl<sub>2</sub>
33. A dilute aqueous solution of sodium fluoride is electrolyzed, the products at the anode and cathode are  
 a) O<sub>2</sub>, H<sub>2</sub>                      b) F<sub>2</sub>, Na                      c) O<sub>2</sub>, Na                      d) F<sub>2</sub>, H<sub>2</sub>
34. Given that  $E^{\ominus}_{\text{Fe}^{2+} | \text{Fe}} = -0.44 \text{ V}$ ;  $E^{\ominus}_{\text{Fe}^{2+} | \text{Fe}^{3+}} = -0.77 \text{ V}$   
 If Fe<sup>2+</sup>, Fe<sup>3+</sup> and Fe solid are kept together, then  
 a) Fe<sup>3+</sup> increases                      b) Fe<sup>3+</sup> decreases                      c) Fe<sup>2+</sup>, Fe<sup>3+</sup> remain unchanged                      d) Fe<sup>2+</sup> decreases
35. The reaction  $1/2\text{H}_2(\text{g}) + \text{AgCl}(\text{s}) \rightarrow \text{H}^{\oplus}(\text{aq}) + \text{Cl}^{\ominus}(\text{aq}) + \text{Ag}(\text{s})$  occurs in the galvanic cell  
 a) Ag|AgCl(s)|KCl(sol)|AgNO<sub>3</sub>|Ag                      b) Pt|H<sub>2</sub>(g)|HCl(sol)|AgNO<sub>3</sub>(sol)|Ag  
 c) Pt|H<sub>2</sub>(g)|HCl(sol)|AgCl(s)|Ag                      d) Pt|H<sub>2</sub>(g)|KCl(sol)|AgCl(s)|Ag
36. The cell EMF is independent of the concentration of the species of the cell in  
 a)  $\text{Fe} | \text{FeO}(\text{s}) | \text{KOH}(\text{aq}) | \text{NiO}(\text{s}) | \text{Ni}_2\text{O}_3(\text{s}) | \text{Ni}$                       b) Pt|(H<sub>2</sub>)|HCl|(Cl<sub>2</sub>)Pt  
 c) Zn|Zn<sup>2+</sup>|Cu<sup>2+</sup>|Cu                      d) Hg, Hg<sub>2</sub>Cl<sub>2</sub>|KCl||AgNO<sub>3</sub>|Ag
37. An electrochemical cell stops working after some time because

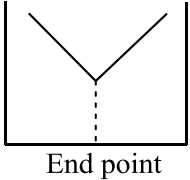
- a) Electrode potential of both the electrodes become zero  
 b) Electrode potential of both the electrodes become equal  
 c) One of the electrode is eaten away  
 d) The reaction starts proceeding in opposite direction
38. Which of the following reactions occurs at the anode during the recharging of lead storage battery?  
 a)  $\text{Pb} + \text{SO}_4^{2-} \rightarrow \text{PbSO}_4 + 2e^-$   
 b)  $\text{Pb} + \text{PbO}_2 + \text{H}_2\text{SO}_4 \rightarrow 2\text{PbSO}_4 + 2\text{H}_2\text{O}$   
 c)  $\text{PbSO}_4 + 2e^- \rightarrow \text{Pb} + \text{SO}_4^{2-}$   
 d)  $\text{PbSO}_4 + 2\text{H}_2\text{O} \rightarrow \text{Pb} + \text{PbO}_2 + 2\text{H}_2\text{O} + 2e^-$
39.  $\text{CH}_3\text{COOH}$  is titrated with  $\text{NaOH}$  solution. Which of the following statements is true?  
 a) Conductance decreases upto equivalence point, after which it increases  
 b) Conductance increases upto equivalence point, after which it decreases  
 c) Conductance first increases (but not rapidly) upto equivalence point and then increases rapidly after equivalent point  
 d) None of the above
40. The value of molar conductivity of  $\text{HCl}$  is greater than that of  $\text{NaCl}$  at a particular temperature because  
 a) Molecular mass of  $\text{HCl}$  is less than that of  $\text{NaCl}$   
 b) Velocity of  $\text{H}^\oplus$  ions is more than that of  $\text{Na}^\oplus$  ions  
 c)  $\text{HCl}$  is strongly acidic  
 d) Ionization of  $\text{HCl}$  is larger than that of  $\text{NaCl}$
41. The potential of the following cell at  $25^\circ\text{C}$  is  
 $\text{Pt} \left| \text{H}_2 \right| (1 \text{ atm}) \left| \text{HNO}_3 \right| (0.1 \text{ M}) \left| \text{NaOH} \right| (10^{-3} \text{ M}) \left| \text{H}_2 \right| (1 \text{ atm}) \left| \text{Pt} \right.$   
 a)  $-0.059 \text{ V}$                       b)  $0.050 \text{ V}$                       c)  $-0.59 \text{ V}$                       d)  $0.5 \text{ V}$
42. During the electrolysis of aqueous solution of  $\text{HCOOK}$ , the number of gases obtained at cathode, anode, and total number of gases are  
 a) 1, 2, 3                      b) 1, 2, 2                      c) 2, 1, 3                      d) 2, 1, 2
43. What volume of  $0.1 \text{ N}$   $\text{FeSO}_4$  can be oxidized by a current of 2 ampere hours?  
 a)  $0.746 \text{ L}$                       b)  $7.46 \text{ L}$                       c)  $1.482 \text{ L}$                       d)  $0.373 \text{ L}$
44. The potential of the cell at  $25^\circ\text{C}$  is  
 $\text{Pt} \left| \text{H}_2 \right| (1 \text{ atm}) \left| \text{NH}_4\text{OH} \right| (10^3 \text{ M}) \left| \text{NaOH} \right| (10^{-3} \text{ M}) \left| \text{H}_2 \right| (1 \text{ atm}) \left| \text{Pt} \right.$   
 Given  $\text{p}K_a$  of  $\text{NH}_4\text{OH} = 4.74$   
 a)  $0.05 \text{ V}$                       b)  $-0.05 \text{ V}$                       c)  $-0.28 \text{ V}$                       d)  $0.28$
45. The  $E^\ominus$  for  $\text{Cu}^{2+}/\text{Cu}^\oplus$ ,  $\text{Cu}^\oplus/\text{Cu}$ ,  $\text{Cu}^{2+}/\text{Cu}$ , are  $0.15 \text{ V}$ ,  $0.50 \text{ V}$ , and  $0.325 \text{ V}$ , respectively. The redox cell showing redox reaction  $2\text{Cu}^\oplus \rightarrow \text{Cu}^{2+} + \text{Cu}$  is made.  $E^\oplus$  of this cell reaction and  $\Delta G^\ominus = 0.350$   
 a)  $E^\ominus = 0.175 \text{ V}$  or  $E^\ominus = 0.350 \text{ V}$                       b)  $n = 1$  or  $2$   
 c)  $\Delta G^\ominus = -33.775 \text{ kJ}$                       d) All of these
46.  $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$ . For this, graph between  $E_{\text{red}}$  versus  $\ln [\text{Cu}^{2+}]$  is a straight line of intercept  $0.34 \text{ V}$ , then the electrode oxidation potential of the half cell  $\text{Cu}|\text{Cu}^{2+} (0.1 \text{ M})$  will be  
 a)  $0.34 + \frac{0.0591}{2}$                       b)  $-0.34 - \frac{0.0591}{2}$                       c)  $0.34$                       d)  $-0.34 + \frac{0.0591}{2}$
47. For an electrolyte solution of  $0.05 \text{ mol L}^{-1}$ , the conductivity has been found to be  $0.0110 \text{ S cm}^{-1}$ . The molar conductivity is  
 a)  $0.055 \text{ S cm}^2 \text{ mol}^{-1}$                       b)  $550 \text{ S cm}^2 \text{ mol}^{-1}$                       c)  $0.22 \text{ S cm}^2 \text{ mol}^{-1}$                       d)  $220 \text{ S cm}^2 \text{ mol}^{-1}$
48. On the basis of position in the electrochemical series, the metal which does not displace  $\text{H}_2$  from water and acid is:  
 a)  $\text{Hg}$                       b)  $\text{Al}$                       c)  $\text{Pb}$                       d)  $\text{Ba}$
49.  $\text{Zn} + \text{Cu}^{2+}(\text{aq}) \rightleftharpoons \text{Cu} + \text{Zn}^{2+}(\text{aq})$   
 Reaction quotient is  $Q = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$ . Variation of  $E_{\text{cell}}$  with  $\log Q$  is of the type with  $\text{OA} = 1.10 \text{ V}$ .  $E_{\text{cell}}$  will be

1.591 V when

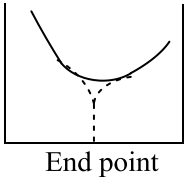


- a)  $[\text{Cu}^{2+}]/[\text{Zn}^{2+}] = 0.01$     b)  $[\text{Zn}^{2+}]/[\text{Cu}^{2+}] = 0.01$     c)  $[\text{Zn}^{2+}]/[\text{Cu}^{2+}] = 0.1$     d)  $[\text{Zn}^{2+}]/[\text{Cu}^{2+}] = 1$
50. A constant current was passed through a solution of  $\text{AuCl}_4^-$  ion between gold electrodes. After a period of 10.0 min, the increase in the weight of cathode was 1.314 g. The total charge passed through solution is (atomic weight of  $\text{AuCl}_4^- = 339$ )
- a)  $1.16 \times 10^{-2}$  F    b)  $3.5 \times 10^{-2}$  F    c)  $2 \times 10^{-2}$  F    d)  $4 \times 10^{-3}$  F
51. During the electrolysis of fused NaCl, which reaction occurs at anode?
- a) Chloride ions are oxidized    b) Chloride ions are reduced  
c) Sodium ions are oxidized    d) Sodium ions are reduced
52. During the electrolysis of  $\text{AgNO}_3$ , the volume of  $\text{O}_2$  formed at STP due to passage of 2 A of current for 965 s is
- a) 0.112 L    b) 0.224 L    c) 11.2 L    d) 22.4 L
53. The value of the reaction quotient  $Q$  for the cell  $\text{Z(s)}|\text{Zn}^{2+}(0.01 \text{ M})||\text{Ag}^+(1.25 \text{ M})|\text{Ag(s)}$  is
- a) 156    b) 125    c)  $1.25 \times 10^{-2}$     d)  $64 \times 10^{-3}$
54.  $\text{Pt}(\text{H}_2)(p_1)|\text{H}^+(1 \text{ M})|(\text{H}_2)(p_2)$ , Pt cell reaction will be exergonic if
- a)  $p_1 = p_2$     b)  $p_1 > p_2$     c)  $p_2 > p_1$     d)  $p_1 = 1 \text{ atm}$
55. 10800 C of electricity passed through the electrolyte deposited 2.977 g of metal with atomic mass 106.4 g  $\text{mol}^{-1}$ . The charge on the metal cation is
- a) +4    b) +3    c) +2    d) +1
56. In electrolysis of very dilute NaOH using platinum electrodes
- a)  $\text{H}_2$  is evolved at cathode and  $\text{O}_2$  at anode  
b)  $\text{NH}_3$  is produced at anode  
c)  $\text{Cl}_2$  is obtained at cathode  
d)  $\text{O}_2$  is produced at cathode and  $\text{H}_2$  at anode
57. The correct order of equivalent conductance at infinite dilution of LiCl, NaCl and KCl is
- a)  $\text{LiCl} > \text{NaCl} > \text{KCl}$   
b)  $\text{KCl} > \text{NaCl} > \text{LiCl}$   
c)  $\text{NaCl} > \text{KCl} > \text{LiCl}$   
d)  $\text{LiCl} > \text{KCl} > \text{NaCl}$
58. According to Kohlrausch law, the limiting value of molar conductivity of an electrolyte  $\text{A}_2\text{B}$  is
- a)  $\lambda_{(\text{A}^\oplus)}^\infty + \lambda_{(\text{B}^\ominus)}^\infty$     b)  $\lambda_{(\text{A}^\oplus)}^\infty - \lambda_{(\text{B}^\ominus)}^\infty$     c)  $2\lambda_{(\text{A}^\oplus)}^\infty + \frac{1}{2}\lambda_{(\text{B}^\ominus)}^\infty$     d)  $2\lambda_{(\text{A}^\oplus)}^\infty + \lambda_{(\text{B}^\ominus)}^\infty$
59. The potential of the cell at 25°C is
- $$\text{Pt} \left| \begin{array}{c} \text{H}_2 \\ (1 \text{ atm}) \end{array} \right| \left| \begin{array}{c} \text{CH}_3\text{COONa} \\ (0.1 \text{ M}) \\ + \\ \text{CH}_3\text{COOH} \\ (0.01 \text{ M}) \end{array} \right| \left| \begin{array}{c} \text{NH}_4\text{Cl} \\ (0.2 \text{ M}) \\ + \\ \text{NH}_4\text{OH} \\ (0.1 \text{ M}) \end{array} \right| \left| \begin{array}{c} \text{H}_2 \\ (1 \text{ atm}) \end{array} \right|$$
- Given  $\text{p}K_a$  of  $\text{CH}_3\text{COOH}$  and  $\text{p}K_b$  of  $\text{NH}_4\text{OH} = 4.74$
- a)  $-0.04 \text{ V}$     b)  $0.04 \text{ V}$     c)  $-0.189 \text{ V}$     d)  $0.189 \text{ V}$
60. Which of the following is a strong electrolyte?
- a)  $\text{Ca}(\text{NO}_3)_2$     b) HCN    c)  $\text{H}_2\text{SO}_3$     d)  $\text{NH}_4\text{OH}$
61. The EMF of a concentration cell consisting of two zinc electrodes, one dipping into M/4 solution of  $\text{ZnSO}_4$

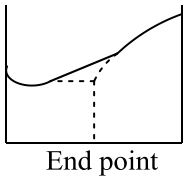
- and the other into M/16 solution of the same salt at 25°C is
- a) 0.0125 V                      b) 0.0250 V                      c) 0.0178 V                      d) 0.0356 V
62. The number of atoms of Ca that will be deposited from a solution of CaCl<sub>2</sub> by a current of 25 mA for 60 s will be  
a)  $4.68 \times 10^{18}$                       b)  $4.68 \times 10^{15}$                       c)  $4.68 \times 10^{10}$                       d)  $2.34 \times 10^{15}$
63. The electrode with reaction:  $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6e^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}$  can be represented as  
a) Pt|H<sup>⊕</sup>(aq), Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup>(aq)                      b) Pt|H<sup>⊕</sup>(aq), Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup>(aq), Cr<sup>3+</sup>(aq)  
c) Pt<sub>H<sub>2</sub></sub>|H<sup>⊕</sup>(aq), Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup>                      d) Pt<sub>H<sub>2</sub></sub>|H<sup>⊕</sup>(aq), Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup>(aq), Cr<sup>3+</sup>(aq)
64. In H<sub>2</sub> – O<sub>2</sub> fuel cell, the reaction occurring at cathode is  
a)  $2\text{H}_2\text{O} + \text{O}_2 \rightarrow 2\text{H}_2\text{O}(\text{l})$                       b)  $\text{H}^{\oplus} + \overset{\ominus}{\text{O}}\text{H} \rightarrow \text{H}_2\text{O}$   
c)  $\text{O}_2 + 2\text{H}_2\text{O} + 4e^- \rightarrow 4\overset{\ominus}{\text{O}}\text{H}$                       d)  $\text{H}^{\oplus} + e^- \rightarrow \frac{1}{2}\text{H}_2$
65. Which of the following statements is wrong?  
a) The conductance of one cm<sup>3</sup> of a material is called specific conductance  
b) Specific conductance increases while equivalent conductance decreases on progressive dilution  
c) The limiting equivalent conductances of weak electrolytes cannot be determined by the extrapolation of the plot of  $\Lambda$  against concentration  
d) The conductivity of metals is due to the movement of electrons
66. 0.5 F of electricity is passed through 500 mL of copper sulphate solution. The amount of copper which can be deposited will be  
a) 63.5 g                      b) 31.75 g                      c) 15.8 g                      d) Unpredictable
67. Assume that during the electrolysis of AgNO<sub>3</sub>, only H<sub>2</sub>O is electrolyzed and O<sub>2</sub> is formed as  
 $2\text{H}_2\text{O} \rightarrow 4\text{H} + \text{O}_2 + 4e^-$   
O<sub>2</sub> formed at NTP due to passage of 2 amperes of current for 96 seconds is  
a) 0.112 L                      b) 0.224 L                      c) 11.2 L                      d) 22.4 L
68. The rusting of iron takes place as follows:  
 $2\text{H}^{\oplus} + 2e^- + \frac{1}{2}\text{O}_2 \rightarrow \text{H}_2\text{O}(\text{l}); E^{\ominus} = +1.23 \text{ V}$   
 $\text{Fe}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Fe}(\text{s}); E^{\ominus} = -0.44 \text{ V}$   
Calculate  $\Delta G^{\ominus}$  for the net process  
a)  $-322 \text{ kJ mol}^{-1}$                       b)  $-152 \text{ kJ mol}^{-1}$                       c)  $-76 \text{ kJ mol}^{-1}$                       d)  $-161 \text{ kJ mol}^{-1}$
69. If  $x$  is the specific resistance of the solution and  $N$  is the normality of the solution, the equivalent conductivity of the solution is given by  
a)  $\frac{1000x}{N}$                       b)  $\frac{1000}{Nx}$                       c)  $\frac{1000N}{x}$                       d)  $\frac{Nx}{1000}$
70. The change required for the reduction of 1 mol of Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup> ions to Cr<sup>3+</sup> is  
a) 96500 C                      b)  $2 \times 96500 \text{ C}$                       c)  $3 \times 96500 \text{ C}$                       d)  $6 \times 96500 \text{ C}$
71. The increase in the value of molar conductivity of acetic acid with dilution is due to  
a) Decrease in interionic forces and increases in  $\alpha$   
b) Increase in the degree of ionization and interionic forces  
c) Increase in self ionization of water  
d) None of these
72. The standard reduction potentials of Cu<sup>2+</sup>|Cu and Cu<sup>2+</sup>|Cu<sup>⊕</sup> are 0.337 V and 0.153 V, respectively. The standard electrode potential of Cu<sup>⊕</sup>|Cu half cell is  
a) 0.184 V                      b) 0.827 V                      c) 0.521 V                      d) 0.490 V
73. An example of a simple fuel cell is  
a) Lead storage battery                      b) Leclanche cell                      c) H<sub>2</sub> – O<sub>2</sub> cell                      d) All of the above
74. By virtue of Faraday's second law of electrolysis, the electrochemical equivalent of the two metals liberated at the electrode has the same ratio as that of their  
a) Atomic masses                      b) Molecular masses                      c) Equivalent masses                      d) Any of three

75. Calculate  $E^\ominus$  for this reaction:  
 $\text{ZnY}^{2-} + 2e^- \rightleftharpoons \text{Zn(s)} + \text{Y}^{4-}$ , where  $\text{Y}^{4-}$  is the completely deprotonated anion of EDTA. The formation constant for  $\text{ZnY}^{2-}$  is  $3.2 \times 10^{16}$  and  $E^\ominus$  for  $\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$  is  $0.76 \text{ V}$   
 a)  $-1.25 \text{ V}$                       b)  $0.48 \text{ V}$                       c)  $+0.68 \text{ V}$                       d)  $-0.27 \text{ V}$
76. The combustion of butane in  $\text{O}_2$  at 1 bar and 298 K shows a decrease in free energy equal to  $2.75 \times 10^3 \text{ kJ mol}^{-1}$  in a fuel cell.  $K$  and  $E^\ominus$  of the fuel cell are  
 a)  $9.55 \times 10^{482}$ ,  $1.096 \text{ V}$     b)  $9.55$ ,  $1.096 \text{ V}$                       c)  $1.023 \times 10^{966}$ ,  $2.85 \text{ V}$     d)  $5.5 \times 10^{484}$ ,  $0.55 \text{ V}$
77. The reaction  $\text{Cu}^{2+}(\text{aq}) + 2\text{Cl}^\ominus(\text{aq}) \rightarrow \text{Cu(s)} + \text{Cl}_2(\text{g})$  has  $E_{\text{cell}}^\ominus = -1.03 \text{ V}$ . This reaction  
 a) Can be made to produce electricity in voltaic cell  
 b) Can be made to occur in an electrolytic cell  
 c) Can occur in acidic medium only  
 d) Can occur in basic medium only
78. If the specific conductance of 1 M  $\text{H}_2\text{SO}_4$  solution is  $26 \times 10^2 \text{ S cm}^2 \text{ mol}^{-1}$ , then the equivalent conductivity would be  
 a)  $1.3 \times 10^2 \text{ S cm}^{-1}$                       b)  $1.6 \times 10^2 \text{ S cm}^{-2}$   
 c)  $13 \text{ S cm}^2 \text{ mol}^{-1}$                       d)  $1.3 \times 10^3 \text{ S cm}^2 \text{ mol}^{-1}$
79. The electrical conductivity of a solution serves as a means of determining the end point in a chemical reaction, involved in the titration of acids, or precipitation. Which of the following conductometric titrations represents the curve of HCl vs  $\text{NH}_4\text{OH}$   
 $x$ -axis  $\Rightarrow$  Volume of alkali added  
 $y$ -axis  $\Rightarrow$  Conductivity
- 

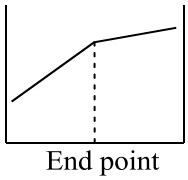
a)



b)



c)



d)
80. The standard reduction potential at  $25^\circ\text{C}$  for the reaction.  $2\text{H}_2\text{O} + 2e^- \rightleftharpoons \text{H}_2 + 2\text{OH}^\ominus$  is  $-0.8277 \text{ volt}$ . The equilibrium constant for the reaction:  
 $2\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^\oplus + \text{OH}^\ominus$  at  $25^\circ\text{C}$  is  
 a)  $10^{-12}$                       b)  $10^{-14}$                       c)  $10^{-11}$                       d)  $10^{-8}$
81. What is the cell entropy change (in  $\text{J K}^{-1}$ ) of the following cell:  
 $\text{Pt(s)} \mid \text{H}_2(\text{g}) \mid \text{CH}_3\text{COOH, HCl} \mid \text{KCl (aq)} \mid \text{Hg}_2\text{Cl}_2(\text{s}) \mid \text{Hg} \mid$   
 $p = 1 \text{ atm} \quad \quad \quad 0.1 \text{ M} \mid 0.1 \text{ M}$   
 The EMF of the cell is found to be  $0.045 \text{ V}$  at  $298 \text{ K}$  and temperature coefficient is  $3.4 \times 10^{-4} \text{ V K}^{-1}$   
 (Given:  $K_{\text{a}(\text{CH}_3\text{COOH})} = 10^{-5} \text{ M}$ )  
 a) 60                      b) 65.2                      c) 69.2                      d) 63.5
82. 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
83. Which metal can deposit copper from copper sulphate solution?  
 a) Mercury                      b) Iron                      c) Gold                      d) Platinum
84. Which of the following statement is correct for a galvanic cell?  
 a) Reduction occurs at cathode                      b) Oxidation occurs at anode  
 c) Electrons flow from anode to cathode                      d) All the statements are correct
85. Ionic strength of a solution made by mixing equal volumes of  $0.01 \text{ M NaCl}$  and  $0.02 \text{ M AlCl}_3$   
 a) 0.065                      b) 0.13                      c) 0.0325                      d) 0.0216
86.  $\text{COO}^\ominus$ ,  $\text{K}^\oplus$ , and  $\text{Na}^\oplus$  are  
 Given the ionic conductance of  $\text{COO}^\ominus$   $73 \text{ cm}^2 \text{ ohm}^{-1} \text{ eq}^{-1}$ , respectively. The equivalent





- b) The velocity of  $\text{NO}_3^\ominus$  is greater than that of  $\text{K}^\oplus$   
 c) The velocities of  $\text{K}^\oplus$  and  $\text{NO}_3^\ominus$  are nearly the same  
 d)  $\text{KNO}_3$  is highly soluble in water
97. The efficiency of a fuel cell is 80% and the standard the heat of reaction is  $-300$  kJ. The reaction involves two electrons in redox change.  $E^\ominus$  for the cell is  
 a) 1.24 V                      b) 2.48 V                      c) 0 V                      d) 0.62 V
98. When a lead storage battery is discharged  
 a)  $\text{SO}_2$  is evolved                      b) Lead is formed  
 c) Lead sulphate is consumed                      d) Sulphuric acid is consumed
99. If  $E^\ominus_{\text{Fe}^{3+}|\text{Fe}}$  and  $E^\ominus_{\text{Fe}^{2+}|\text{Fe}}$  are  $-0.36$  V  $-0.439$  V, respectively, then the value of  $E^\ominus_{\text{Fe}^{3+}|\text{Fe}^{2+}}$  is  
 a)  $(-36 - 0.439)$  V                      b)  $[(-0.36) + 2(-0.439)]$  V  
 c)  $(-0.36 - 0.439)$  V                      d)  $[3(0.36) - 2(-0.439)]$  V
100. The strongest oxidizing agent among the following is  
 a)  $\text{K}_2\text{Cr}_2\text{O}_7$                       b)  $\text{MnO}_2$                       c)  $\text{KMnO}_4$                       d)  $\text{Cl}_2$
101. Calculate the solubility product of  $\text{Co}_2[\text{Fe}(\text{CN})_6]$  in water at  $25^\circ\text{C}$   
 Given, conductivity of saturated solution of  $\text{Co}_2[\text{Fe}(\text{CN})_6]$  is  $2.06 \times 10^{-6} \Omega^{-1} \text{cm}^{-1}$  and that of water used is  $4.1 \times 10^{-7} \Omega^{-1} \text{cm}^{-1}$ . The ionic molar conductivities of  $\text{Co}^{2+}$  and  $\text{Fe}(\text{CN})_6^{4-}$  are  $86.0 \Omega \text{cm}^2 \text{mol}^{-1}$  and  $444.0 \Omega^{-1} \text{cm}^2 \text{mol}^{-1}$ , respectively  
 a)  $7.87 \times 10^{-7}$                       b)  $7.87 \times 10^{-6}$                       c)  $7.87 \times 10^{-8}$                       d)  $7.87 \times 10^{-9}$
102. A reaction:  

$$\frac{1}{2}\text{H}_2(\text{g}) + \text{AgCl}(\text{s}) \rightleftharpoons \text{H}^\oplus(\text{aq}) + \text{Cl}^\ominus(\text{aq}) + \text{Ag}(\text{s})$$
  
 Occurs in a galvanic cell. The structure of the cell will be  
 a)  $\text{Ag}, \text{AgCl}(\text{s})|\text{KCl}(\text{sol})|\text{AgNO}_3(\text{sol}), \text{Ag}$                       b)  $\text{Pt}, \text{H}_2(\text{g})|\text{HCl}(\text{sol})|\text{AgNO}_3(\text{sol}), \text{Ag}$   
 c)  $\text{Pt}, \text{H}_2(\text{g})|\text{HCl}(\text{sol})|\text{AgCl}(\text{s}), \text{Ag}$                       d)  $\text{Pt}, \text{H}_2(\text{g})|\text{KCl}(\text{sol})|\text{AgCl}(\text{s}), \text{Ag}$
103. A hydrogen electrode placed in a solution containing sodium acetate and acetic in the ratio of  $x:y$  and  $y:x$  has an electrode potential value  $E_1$  and  $E_2$  volts, respectively, at  $25^\circ\text{C}$ . The  $\text{pK}_a$  value of acetic acid is  
 a)  $\frac{-(E_1 + E_2)}{2 \times 0.059}$                       b)  $\frac{E_1 + E_2}{2 \times 0.059}$                       c)  $\frac{E_2 - E_1}{2 \times 0.059}$                       d)  $\frac{-(E_1 + E_2)}{0.059}$
104.  $\text{Zn}|\text{Zn}^{2+}(c_1)||\text{Zn}^{2+}(c_2)|\text{Zn}$ . For this cell  $\Delta G$  is negative if:  
 a)  $c_1 = c_2$                       b)  $c_1 > c_2$                       c)  $c_2 > c_1$                       d) None
105. The standard EMF of decinormal calomel electrode is  $0.268$  V. The EMF is  
 a)  $-0.298$  V                      b)  $0.327$  V                      c)  $-0.327$  V                      d)  $0.298$  V
106. Given electrode potentials are  
 $\text{Fe}^{3+} + e^- \longrightarrow \text{Fe}^{2+}; \quad E^\ominus = 0.771$  V  
 $\text{I}_2 + 2e^- \longrightarrow 2\text{I}^\ominus; \quad E^\ominus = 0.536$  V  
 $E^\ominus_{\text{cell}}$  for the cell reaction,  
 $2\text{Fe}^{3+} + 2\text{I}^\ominus \longrightarrow \text{Fe}^{2+} + \text{I}_2$  is  
 a)  $(2 \times 0.771 - 0.536) = 1.006$  V                      b)  $(0.771 - 0.5 \times 0.536) = 0.503$  V  
 c)  $(0.771 - 0.536) = 0.235$  V                      d)  $(0.536 - 0.771) = -0.236$  V
107. The volume is obtained at cathode and anode on electrolysis of aqueous sodium succinate are  
 a) Anode: Ethene +  $\text{CO}_2$ , Cathode :  $\text{H}_2$                       b) Anode: Ethyne +  $\text{CO}_2$ , Cathode :  $\text{H}_2$   
 c) Anode: Ethene +  $\text{H}_2$ , Cathode : Na                      d) Anode: Ethyne +  $\text{H}_2$ , Cathode : Na
108. How much will the reduction potential of a hydrogen electrode change when its solution initially at  $\text{pH} = 0$  is neutralized to  $\text{pH} = 7$ ?  
 a) Increase by  $0.059$  V                      b) Decrease by  $0.059$  V                      c) Increase by  $0.41$  V                      d) Decrease by  $0.41$  V
109. The increase in the molar conductivity of  $\text{HCl}$  with dilution is due to

- a) Increase in the self ionization of water  
 c) Decrease in the self ionization of water
- b) Hydrolysis of HCl  
 d) Decrease in the interionic forces
110. Among Na, Hg, S, Pt, and graphite which can be used as electrode in electrolytic cell having aqueous solutions?  
 a) Na and S  
 b) Hg, Pt and S  
 c) Na, Hg, and S  
 d) Hg, Pt, and graphite
111. What weight of copper will be deposited by passing 2 faradays of electricity through a cupric salt (atomic weight of Cu = 63.5)?  
 a) 2.0  
 b) 3.175  
 c) 63.5  
 d) 127.0
112. The thermodynamic efficiency of cell is given by  
 a)  $\Delta H/\Delta G$   
 b)  $\frac{nFE}{\Delta G}$   
 c)  $\frac{-nFE}{\Delta H}$   
 d)  $nFE^\ominus$
113. Calculate the potential of the following cell:  

$$\text{Pt} \left| \begin{array}{cc} \text{Co}^{2+}, \text{Co}^{3+} \\ (2.0 \text{ M}) \quad (1.0 \text{ M}) \end{array} \right| \left| \begin{array}{ccc} \text{Cr}^{3+}, \text{Cr}_2\text{O}_7^{2-}, \text{H}^\oplus \\ (1.0 \text{ M}) \quad (4.0 \text{ M}) \quad (1.0 \text{ M}) \end{array} \right| \text{Pt}$$

$$E^\ominus_{\text{Co}^{2+}|\text{Co}^{3+}} = -2\text{V}, E^\ominus_{\text{Cr}_2\text{O}_7^{2-}|\text{Cr}^{3+}} = +1.0 \text{ V}$$
 a) 1.024 V  
 b) -1.024 V  
 c) 0.976 V  
 d) -0.976 V
114. From the following information, calculate the solubility product of AgBr  
 $\text{AgBr}(s) + e^- \rightarrow \text{Ag}(s) + \text{Br}^\ominus(\text{aq}); E^\ominus = 0.07 \text{ V}$   
 $\text{Ag}^\oplus(\text{aq}) + e^- \rightarrow \text{Ag}(s); E^\ominus = 0.80 \text{ V}$   
 a)  $4 \times 10^{-13}$   
 b)  $4 \times 10^{-10}$   
 c)  $4 \times 10^{-17}$   
 d)  $4 \times 10^{-7}$
115. Given the standard potential of the following at 25°C  
 $\text{MnO}_2 \rightarrow \text{Mn}^{3+}; E^\ominus = 0.95 \text{ V}$   
 $\text{Mn}^{3+} \rightarrow \text{Mn}^{2+}; E^\ominus = 1.51 \text{ V}$   
 The standard potential of  $\text{MnO}_2 \rightarrow \text{Mn}^{2+}$  is  
 a) -0.56 V  
 b) -2.46 V  
 c) -1.23 V  
 d) 1.23 V
116. Given the following cell at 25°C  

$$\text{Pt} \left| \begin{array}{c} \text{H}_2 \\ (1 \text{ atm}) \end{array} \right| \left| \begin{array}{c} \text{CH}_3\text{COOH} \\ (10^{-3} \text{ M}) \end{array} \right| \left| \begin{array}{c} \text{NaOH} \\ (10^{-3} \text{ M}) \end{array} \right| \left| \begin{array}{c} \text{H}_2 \\ (1 \text{ atm}) \end{array} \right| \text{Pt}$$
 What will be the potential of the cell?  
 Given  $\text{p}K_a$  of  $\text{CH}_3\text{COOH} = 4.74$   
 a) -0.42 V  
 b) 0.42 V  
 c) -0.19 V  
 d) 0.19 V
117. The standard reduction potential of  $\text{Li}^\oplus|\text{Li}$ ,  $\text{Ba}^{2+}|\text{Ba}$ ,  $\text{Na}^\oplus|\text{Na}$  and  $\text{Mg}^{2+}|\text{Mg}$  are -3.05, -2.73, -2.71, and -2.37 volts, respectively. Which one of the following is strongest oxidizing agent?  
 a)  $\text{Na}^\oplus$   
 b)  $\text{Li}^\oplus$   
 c)  $\text{Ba}^{2+}$   
 d)  $\text{Mg}^{2+}$
118.  $\text{Pt}(\text{H}_2)(1 \text{ atm})|\text{H}_2\text{O}$ , electrode potential at 298 K is  
 a) -0.2364 V  
 b) -0.4137 V  
 c) 0.4137 V  
 d) 0.00 V
119. For the electrochemical cell,  $(\text{M}|\text{M}^\oplus)||(\text{X}^\ominus|\text{X}), E^\ominus_{(\text{M}^\oplus|\text{M})} = 0.44 \text{ V}$  and  $E^\ominus_{(\text{X}|\text{X}^\ominus)} = 0.334 \text{ V}$ . From this data, one can deduce that  
 a)  $\text{M} + \text{X} \rightarrow \text{M}^\oplus + \text{X}^\ominus$  is a spontaneous reaction  
 b)  $\text{M}^\oplus + \text{X}^\ominus \rightarrow \text{M} + \text{X}$  is a spontaneous reaction  
 c)  $E_{\text{cell}} = 0.77 \text{ V}$   
 d)  $E_{\text{cell}} = -0.77$
120. Copper can be deposited from acidified copper sulphate and alkaline cuprous cyanide. If the same current is passed for a definite time:  
 a) The amount of copper deposited from acidic copper sulphate will be higher  
 b) The amount of copper deposited from alkaline cuprous cyanide will be higher  
 c) The same amount of copper will be deposited  
 d) None of these
121. The gas evolved at the anode when  $\text{K}_2\text{SO}_4(\text{aq})$  is electrolyzed between Pt electrodes is



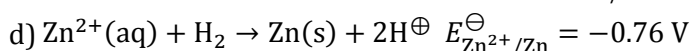
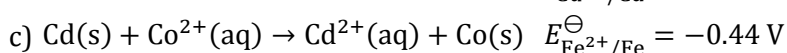
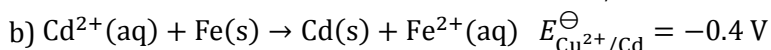
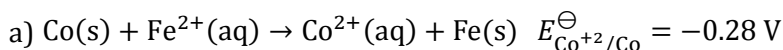
135.  $\text{Cu}^{\oplus} + e^{-} \longrightarrow \text{Cu}$ ,  $E^{\ominus} = x_1$  volt  
 $\text{Cu}^{2+} + 2e^{-} \longrightarrow \text{Cu}$ ,  $E^{\ominus} = x_2$  volt, then for  
 $\text{Cu}^{2+} + e^{-} \longrightarrow \text{Cu}^{\oplus}$ ,  $E^{\ominus}$  (volt) will be  
 a)  $x_1 - 2x_2$                       b)  $x_1 + 2x_2$                       c)  $x_1 - x_2$                       d)  $2x_2 - x_1$
136. On carrying out the electrolysis of acidified water, the volume of hydrogen liberated at STP condition is 22.4 L. The volume of oxygen liberated is  
 a) 22.4 L                      b) 44.8 L                      c) 11.2 L                      d) 2.24 L
137. Ionic strength of 0.4 M  $\text{CaCl}_2$  is  
 a) 1.2                      b) 1.0                      c) 0.9                      d) 0.8
138. Total charge required to convert three moles of  $\text{Mn}_3\text{O}_4$  to  $\text{MnO}_4^{\ominus}$  in present of alkaline medium  
 a) 10 F                      b) 20 F                      c) 30 F                      d) 40 F
139. The mass of copper that will be deposited at cathode in electrolysis of 0.2 M solution of copper sulphate when a quantity of electricity equal t that required to liberate 2.24 L of hydrogen from 0.1 M aqueous  $\text{H}_2\text{SO}_4$  is passed (atomic mass of Cu = 63.5) will be  
 a) 1.59 g                      b) 3.18 g                      c) 6.35 g                      d) 12.70 g
140.  $\text{Ag}|\text{Ag}^{\oplus}(1\text{ M})|\text{Ag}^{\oplus}||\text{Ag}$   
 1 L solution      1 L solution  
 0.5 F electricity in the LHS (anode) and 1 F of electricity in the RHS (cathode) in first passed making them independent electrolytic cells at 298 K. EMF of the cell after electrolysis will be  
 a) Increased                      b) Decreased                      c) No change                      d) Time is also required
141. During the electrolysis of the aqueous solution of copper sulphate using Pt electrodes, the reaction taking place at anode electrode is  
 a)  $\text{Cu}^{2+} + 2e^{-} \rightarrow \text{Cu}$                       b)  $\text{Cu} \rightarrow \text{Cu}^{2+} + 2e^{-}$   
 c)  $2\text{H}_2\text{O} \rightarrow 4\text{H}^{\oplus} + \text{O}_2 + 4e^{-}$                       d)  $\text{H}_2\text{O} + e^{-} \longrightarrow \overset{\ominus}{\text{O}}\text{H} + 1/2\text{H}_2$
142. Silver is removed electrolytically from 200 mL of a 0.1 N solution of  $\text{AgNO}_3$  by a current of 0.1 A. How long will it take to remove half of the silver from the solution?  
 a) 0.1 s                      b) 100 s                      c) 965 s                      d) 9650 s
143. The EMF of Ni-Cad battery is dependent of:  
 a)  $\text{Cd}(\text{OH})_2$                       b)  $\text{Ni}(\text{OH})_2$                       c)  $\overset{\ominus}{\text{O}}\text{H}$                       d) None of these
144. A current 965 A is passed for 1 s through 1 L solution of 0.02 N  $\text{NiSO}_4$  using Ni electrodes. What is the new concentration of  $\text{NiSO}_4$ ?  
 a) 0.01 N                      b) 0.01 M                      c) 0.002 M                      d) 0.02 M
145. In acid medium,  $\text{MnO}_4^{\ominus}$  is an oxidizing agent  
 $\text{MnO}_4^{\ominus} + 8\text{H}^{\oplus} + 5e^{-} \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$   
 If  $\text{H}^{\oplus}$  ion concentration is doubled, electrode potential of the half cell  $\text{MnO}_4^{\ominus}, \text{Mn}^{2+}|\text{Pt}$  will  
 a) Increase by 28.36 mV                      b) Decrease by 28.36 mV  
 c) Increase by 14.23 mV                      d) Decrease by 142.30 mV
146. The values of  $\Lambda_{\text{eq}}^{\infty}$  for  $\text{NH}_4\text{Cl}$ ,  $\text{NaOH}$  and  $\text{NaCl}$  are, respectively, 149.74, 248.1, and 126.4  $\text{ohm}^{-1} \text{cm}^2 \text{eq}^{-1}$ . The value of  $\Lambda_{\text{eq}}^{\infty} \text{NH}_4\text{OH}$  is  
 a) 371.44  
 b) 271.44  
 c) 71.44  
 d) It cannot be calculated from the data given
147. Ionic strength of 0.2 M  $\text{Na}_3\text{PO}_4$  is  
 a) 2.8                      b) 1.2                      c) 1.0                      d) 1.8
148. A current strength of 1.0 A is passed for 96.5 s through 100 mL of a solution of 0.05 M  $\text{KCl}$ . The concentration of the final solution with respect to  $\overset{\ominus}{\text{O}}\text{H}$  ions is

- a) 0.005 M                      b) 0.05 M                      c) 0.01 M                      d) 0.001 M
149.  $\text{Pt}(\text{Cl}_2)(p_1)|\text{HCl}(0.1 \text{ M})|(\text{Cl}_2)(p_2)$ , Pt cell reaction will be endergonic if  
 a)  $p_1 = p_2$                       b)  $p_1 > p_2$                       c)  $p_2 > p_1$                       d)  $p_1 = p_2 = 1 \text{ atm}$
150. A solution containing 1 mol per liter of each  $\text{Cu}(\text{NO}_3)_2$ ,  $\text{AgNO}_3$ , and  $\text{Hg}_2(\text{NO}_3)_2$  is being electrolyzed by using inert electrodes. The values of standard electrode potentials in volts (reduction potential) are  
 $\text{Ag}|\text{Ag}^\oplus = +0.80$ ,  $2\text{Hg}|\text{Hg}_2^{2+} = +0.79$   
 $\text{Cu}|\text{Cu}^{2+} = +0.34$ ,  $\text{Mg}|\text{Mg}^{2+} = -2.37$   
 With increasing voltage, the sequence of deposition of metals at the cathode will be  
 a) Ag, Hg, Cu, Mg                      b) Mg, Cu, Hg, Ag                      c) Ag, Hg, Cu                      d) Cu, Hg, Ag

### Multiple Correct Answers Type

151. On the electrolysis of a  $10^{-6} \text{ M}$  HCl solution :  
 a)  $\text{O}_2$  gas is produced at the anode  
 b)  $\text{Cl}_2$  gas is produced at the anode  
 c)  $\text{H}_2$  gas is produced at the cathode  
 d)  $\text{O}_2$  gas is produced at the cathode
152. Which of the following changes will increase the EMF of the cell:  
 $\text{Co}(\text{s})|\text{CoCl}_2(\text{M}_1)||\text{HCl}(\text{M}_2)||(\text{H}_2, \text{g}) \text{Pt}$   
 a) Increase the volume of  $\text{CoCl}_2$  solution from 100 mL to 200 mL  
 b) Increase  $\text{M}_2$  from 0.01 M to 0.50 M  
 c) Increase the pressure of the  $\text{H}_2(\text{g})$  from 1.0 to 2.0 atm  
 d) Increase  $\text{M}_1$  from 0.01 M to 0.50 M
153. A solution containing  $\text{Na}^\oplus$ ,  $\text{NO}_3^\ominus$ ,  $\text{Cl}^\ominus$  and  $\text{SO}_4^{2-}$  ions, all at unit concentrations, is electrolyzed between nickel anode and platinum cathode. As the current is passed through the cell  
 a) pH of the cathode increases                      b) Oxygen is the major product at anode  
 c) Nickel is deposited at cathode                      d) Chlorine is the major product at anode
154. Which of the following statements is/are incorrect?  
 a) Reduction occurs at the cathode in both galvanic and electrolytic cells  
 b) Oxidation takes place at the cathode in both galvanic and electrolytic cells  
 c) The anode is the negative terminal and the cathode is the positive terminal in a galvanic cell  
 d) The anode is the negative terminal and the cathode is the positive terminal for an electrolytic cell
155. To observe the effect of concentration on the conductivity, electrolytes of different natures are taken in two vessels A and B: A contains weak electrolyte, e.g.,  $\text{NH}_4\text{OH}$  and B contains strong electrolyte, e.g.,  $\text{NaCl}$ . In both containers, the concentration of respective electrolyte is increased and the conductivity observed:  
 a) In A conductivity increases, in B conductivity decrease  
 b) In A conductivity decreases while, in B conductivity decrease  
 c) In both A and B conductivity increases  
 d) In both A and B conductivity decreases
156. Which of the following statements is/are correct?  
 a) The cell constant of an electrolytic cell is measured as the product:  $kR$  rather using  $l/a$   
 b) As an electrolytic solution is diluted, its conductance, equivalent conductance, and molar conductance, and molar conductance increase  
 c) Kohlrausch's law may be applied to calculate molar conductance at infinite dilution for both weak and strong electrolytes  
 d) Kohlrausch's law may also be applied at any concentration of the electrolyte
157. The unit(s) of conductance is/are :  
 a) Ohm                      b)  $\text{Ohm}^{-1}$                       c) Mho                      d) Siemens
158. How can the oxidation potential of a hydrogen electrode be decreased?





170. Which of the following statements is / are correct?

- a) The electrolysis of molten NaOH liberates  $\text{O}_2$  at the anode
- b) The electrolysis of molten NaOH liberates  $\text{O}_2$  at the cathode
- c) Molten NaCl conducts electricity due to the presence of  $\text{Na}^+$  and  $\text{Cl}^-$  ions
- d) the electrolysis of molten KCl produces  $\text{Cl}_2$  at the cathode

171. When an aqueous solution of  $\text{CaCl}_2$  is electrolyzed using inert electrodes, which of the following is (are) true?

- a) Calcium deposits on cathode
- b) Calcium deposits an anode
- c) Chlorine is liberated on anode
- d) Calcium hydroxide precipitates near cathode on prolonged hydrolysis

172. Given:

$$E_{\text{Ag}^{\oplus}|\text{Ag}}^{\ominus} = 0.80 \text{ V}, E_{\text{Mg}^{2+}|\text{Mg}}^{\ominus} = -2.37 \text{ V};$$

$$E_{\text{Cu}^{2+}|\text{Cu}}^{\ominus} = 0.34 \text{ V}, E_{\text{Hg}^{2+}|\text{Hg}}^{\ominus} = 0.79 \text{ V}$$

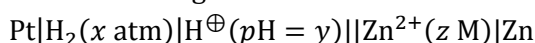
Which of the following statements is/are incorrect?

- a)  $\text{AgNO}_3$  can be stored in copper vessel
- b)  $\text{Cu(NO}_3)_2$  can be stored in magnesium vessel
- c)  $\text{CuCl}_2$  can be stored in silver vessel
- d)  $\text{HgCl}_2$  can be stored in copper vessel

173. The conductivity of a solution depends upon:

- a) Its dilution
- b) No. of ions present in solution
- c) Nature of solvent
- d) Nature of solute

174. In the following electrochemical cell:



$E_{\text{cell}}^{\ominus} = E_{\text{cell}}$ . This will be possible when

- a)  $[\text{Zn}^{2+}] = [\text{H}^{\oplus}] = 1 \text{ M}$  and  $p_{\text{H}_2} = 1 \text{ atm}$
- b)  $[\text{Zn}^{2+}] = 0.01 \text{ M}$ ,  $[\text{H}^{\oplus}] = 0.1$ , and  $p_{\text{H}_2} = 1 \text{ atm}$
- c)  $[\text{Zn}^{2+}] = 1 \text{ M}$ ,  $[\text{H}^{\oplus}] = 0.1 \text{ M}$ , and  $p_{\text{H}_2} = 0.01 \text{ atm}$
- d)  $[\text{Zn}^{2+}] = [\text{H}^{\oplus}] = 0.1 \text{ M}$  and  $p_{\text{H}_2} = 0.1 \text{ atm}$

175. Pb can reduce which of the following species?

- a)  $\text{Cu}^{2+}$
- b)  $\text{Br}_2$
- c)  $\text{Ag}^+$
- d)  $\text{Zn}^{2+}$

176. For which, the formula  $\alpha = \frac{\lambda_w}{\lambda_{\infty}}$  does not hold good?

- a)  $\text{CH}_3\text{COOH}$
- b)  $\text{HCl}$
- c)  $\text{HClO}_4$
- d)  $\text{CH}_3\text{NH}_2$

177. The passage of current through the solution of a certain electrolyte results in the liberation of  $\text{H}_2$  at the cathode and chlorine at the anode. The solution in the container could be :

- a)  $\text{NaCl(aq)}$
- b)  $\text{CuCl}_2(\text{aq})$
- c)  $\text{KCl(aq)}$
- d)  $\text{MgCl}_2(\text{aq})$

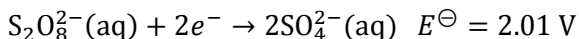
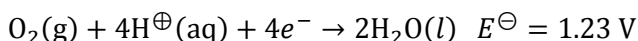
178. The term infinite dilution refers for:

- a)  $\alpha \rightarrow 1$  for weak electrolytes
- b) An electrolyte is 100% dissociated
- c) All interionic effects disappear

- d) When equivalent conductance of an electrolyte becomes constant
179. For a strong electrolyte, equivalent conductance increases slowly with dilution and can be expressed by the relationship:  $\Lambda_m = \Lambda_m^\circ - A\sqrt{c}$   
Which electrolyte(s) have same value of  $A$ ?  
a) NaCl                              b)  $\text{CaCl}_2$                               c)  $\text{ZnCl}_2$                               d)  $\text{MgSO}_4$
180. The existence of ions in strong electrolytes in solid state is confirmed by :  
a) They conduct current in molten state  
b) X-ray method  
c) Dipole moment measurement  
d) None of the above
181. A current of 2.68 A is passed for 1.0 hour through an aqueous solution of  $\text{CuSO}_4$  using copper electrodes  
Which of the following statements is/are correct?  
a) Increase in the mass of cathode = 3.174 g  
b) Decrease in the mass of anode = 3.174 g  
c) No change in the mass of electrodes  
d) The ratio between the change in the mass of cathode to anode is 1:2
182. Which of the following statements is/are correct?  
a)  $\text{F}_2$  is the strongest oxidizing agent                              b) Li is the strongest reducing agent  
c)  $\text{Li}^\oplus$  is the weakest oxidizing agent                              d)  $\text{F}_2$  has at highest reducing potential
183. Which modification(s) is/are necessary to determine resistance of solutions by usual method of conductance measurements?  
a) A.C. should be used  
b) A conductivity cell is used  
c) Galvanometer is replaced by head phone or magic eye arrangement  
d) None of the above
184. In which of the following cells salt bridge is not needed?  
a)  $\text{Pb}|\text{PbSO}_4(\text{s})|\text{H}_2\text{SO}_4|\text{PbO}_2(\text{s})|\text{Pb}$                               b)  $\text{Cd}|\text{CdO}(\text{s})|\text{KOH}(\text{aq})|\text{NiO}_2(\text{s})|\text{Ni}$   
c)  $\text{Fe}(\text{s})|\text{FeO}(\text{s})|\text{KOH}(\text{aq}).\text{Ni}_2\text{O}_3(\text{s})|\text{Ni}$                               d)  $\text{Zn}|\text{ZnSO}_4|\text{CuSO}_4|\text{Cu}$
185. On electrolysis, which of the following arrangements will lead to oxygen being evolved at the anode?  
a) Dilute  $\text{H}_2\text{SO}_4$  with copper electrodes  
b) Fused NaOH with an iron cathode and a nickel anode  
c) Aqueous  $\text{AgNO}_3$  solution with platinum electrodes  
d) Dilute  $\text{H}_2\text{SO}_4$  with platinum electrodes
186. Which is/are weak electrolyte(s)?  
a)  $\text{Mg}(\text{OH})_2$                               b)  $\text{Zn}(\text{OH})_2$                               c) NaOH                              d)  $\text{CH}_3\text{COOH}$
187. Which of the following statements is/are correct?  
a) The electrolysis of concentrated  $\text{H}_2\text{SO}_4$  at  $0 - 5^\circ\text{C}$  using a Pt electrode produces  $\text{H}_2\text{S}_2\text{O}_8$   
b) The electrolysis of a brine solution produces  $\text{NaClO}_3$  and NaClO  
c) The electrolysis of a  $\text{CuSO}_4$  solution using Pt electrodes causes the liberation of  $\text{O}_2$  at anode and the deposition of copper at cathode  
d) All electrolytic reactions are redox reactions
188. For the cell  $\text{Tl}|\text{Tl}^+(0.001M)||\text{Cu}^{2+}(0.1M)|\text{Cu}$ .  $E_{\text{cell}}^\circ$  at  $25^\circ\text{C}$  is 0.83 V, which can be increased :  
a) By increasing  $[\text{Cu}^{2+}]$     b) By increasing  $[\text{Tl}^+]$     c) By decreasing  $[\text{Cu}^{2+}]$     d) By decreasing  $[\text{Tl}^+]$
189. Hydrogen gas will reduce :  
a) Heated cupric oxide  
b) Heated ferric oxide  
c) Heated stannic oxide  
d) Heated aluminium oxide
190. Peroxodisulphate salts (e.g.,  $\text{Na}_2\text{S}_2\text{O}_8$ ) are strong oxidizing agents used as bleaching agents for fats, oils, etc



Given:



Which of the following statements is (are) correct?

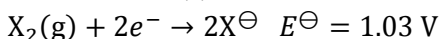
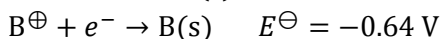
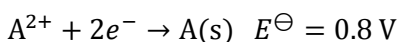
- a) Oxygen gas can oxidize sulphate ion to per-oxo disulphate ion ( $\text{S}_2\text{O}_8^{2-}$ ) in acidic solution
  - b)  $\text{O}_2(\text{g})$  is reduced to water
  - c) Water is oxidized to  $\text{O}_2$
  - d)  $\text{S}_2\text{O}_8^{2-}$  ions are reduced to  $\text{SO}_4^{2-}$  ions
191. A KI solution containing starch, turns blue on the addition of chlorine. Which of the following statements explain(s) this?
- a) The standard reduction potential of  $\text{Cl}_2$  is more than that of  $\text{I}_2$ .
  - b) The standard oxidation potential of  $\text{Cl}_2$  is more than that of  $\text{I}_2$
  - c) The product formed when  $\text{Cl}_2$  combines with starch is blue in colour.
  - d) The product formed when  $\text{I}_2$  combines with starch is blue in colour
192. During the working of a galvanic cell and with the passage of time
- a) Spontaneity of the cell reaction decreases,  $E_{\text{cell}}$  decreases
  - b) Reaction quotient  $Q$  decreases,  $E_{\text{cell}}$  increases
  - c) Reaction quotient  $Q$  increases,  $E_{\text{cell}}$  decreases
  - d) At equilibrium,  $Q = K_{\text{eq}}$ ,  $E_{\text{cell}} = 0$
193. In a solution of  $\text{CH}_3\text{COOH}$ , we have at infinite dilution:
- a)  $\text{CH}_3\text{COO}^{-}$  ion
  - b)  $\text{H}^{+}$  ion
  - c)  $\text{CH}_3\text{COOH}$  molecule
  - d) None of these
194. A current is passed through 500 mL of an aqueous solution of  $\text{CaI}_2$ . After sometime, it is observed, it is observed that 50 millimoles of  $\text{I}_2$  have been formed. Which of the following statements is (are) correct?
- a) The number of faradays of charge passed through the solution is 0.10 F
  - b) The volume of dry H at STP that has been formed during electrolysis is 1120 mL
  - c) The pH of the solution is nearly 0.7
  - d) The mass of calcium produced is 2.0 g
195. Iron can be prevented from rusting by
- a) Connecting iron to more electropositive metal – a case of cathodic protection
  - b) Connecting iron to more electropositive metal – a case of anodic protection
  - c) Connecting iron to less electropositive metal – a case of anodic protection
  - d) Connecting iron to less electropositive metal – a case of cathodic protection
196. Which of the following conditions will increase the voltage of the cell represented by the following equation?
- $$\text{Zn}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu}(\text{s})$$
- a) Decrease in temperature
  - b) Increase in pressure
  - c) Decrease in the concentration of  $\text{Zn}^{2+}(\text{aq})$  ions
  - d) Increase in the concentration of  $\text{Cu}^{2+}(\text{aq})$  ions
197. Electrolysis of aqueous solutions of which of the following substances results in only the decomposition of water?
- a) Potassium chloride
  - b) Zinc sulphate
  - c) Potassium hydroxide
  - d) Sodium phosphate
198. During discharging of a lead storage battery
- a) The reaction at anode is  $\text{Pb} \rightarrow \text{Pb}^{2+} + 2\text{e}^{-}$
  - b) The reaction taking place at cathode is  $\text{Pb}^{2+} + 2\text{e}^{-} \rightarrow \text{Pb}$
  - c) The overall reaction is  
$$\text{Pb} + \text{PbO}_2 + 4\text{H}^{\oplus} + 2\text{SO}_4^{2-} \rightarrow 2\text{PbSO}_4 + 2\text{H}_2\text{O}$$
  - d) The reaction taking place at cathode is  
$$\text{PbO}_2 + 4\text{H}^{\oplus} + 2\text{e}^{-} \rightarrow \text{Pb}^{2+} + 2\text{H}_2\text{O}$$
199. In electrolysis of very dilute  $\text{NaOH}$  solution, using platinum electrodes.
- a) Na is obtained at cathode
  - b)  $\text{H}_2$  is evolved at anode
  - c)  $\text{O}_2$  is produced at anode
  - d)  $\text{H}_2$  is produced at cathode

200. If  $A + B \rightleftharpoons C + D$ ;  $K_C = K_1$  and  $E^\ominus = a$  V  
 $2A + 2B \rightleftharpoons 2C + 2D$ ;  $K_C = K_2$  and  $E^\ominus = b$  V  
 Then  
 a)  $a = b$                                       b)  $K_2 = K_1^2$                                       c)  $a = 2b$                                       d)  $b = a^2$
201. Select the wrong relations :  
 a)  $\Delta S = \left(\frac{\partial E}{\partial T}\right)_p \times nF$                       b)  $-\Delta S = \left(\frac{\partial E}{\partial T}\right)_p \times nF$                       c)  $\left(\frac{\partial E}{\partial T}\right)_p = \left(\frac{\partial \Delta S}{\partial T}\right)$                       d)  $\left(\frac{\partial E}{\partial T}\right)_p = \frac{\Delta H + nEF}{T \times nF}$
202. Consider the cell:  
 $\text{Cd(s)}|\text{Cd}^{2+}(1.0\text{ M})||\text{Cu}^{2+}(1.0\text{ M})|\text{Cu(s)}$   
 If we wish to make a cell with a more positive voltage using the same substances, we should  
 a) Increase both  $[\text{Cd}^{2+}]$  and  $[\text{Cu}^{2+}]$  to 2.0 M  
 b) Decrease the  $[\text{Cd}^{2+}]$  to 0.1 M  
 c) Increase the  $[\text{Cu}^{2+}]$  to 2.0 M  
 d) Decrease both the  $[\text{Cd}^{2+}]$  and  $[\text{Cu}^{2+}]$  to 0.01 M
203. Which of the following statements regarding rusting of iron is/are correct?  
 a) It takes place in moist air  
 b) It is stopped in  $\text{CO}_2$  atmosphere  
 c) It produces Fe (III) oxide  
 d) It is an electrochemical process
204. Given:  
 Oxidation:  $\text{H}_2\text{O}_2 \rightarrow \text{O}_2 + 2\text{H}^\oplus + 2e^-$      $E^\ominus = -0.69$  V;  
 $2\text{F}^\ominus \rightarrow \text{F}_2 + 2e^-$                        $E^\ominus = -2.87$  V;  
 Reduction:  $\text{H}_2\text{O}_2 + 2\text{H}^\oplus + 2e^- \rightarrow 2\text{H}_2\text{O}$      $E^\ominus = 1.77$  V;  
 $2\text{I}^\ominus \rightarrow \text{I}_2 + 2e^-$                        $E^\ominus = -0.54$  V;  
 Which of the following statements is/are correct?  
 a)  $\text{H}_2\text{O}_2$  behaves as an oxidant for  $\text{I}^\ominus$                                       b)  $\text{H}_2\text{O}_2$  behaves as a reduction for  $\text{I}_2$   
 c)  $\text{H}_2\text{O}_2$  behaves as an oxidant for  $\text{F}^\ominus$                                       d)  $\text{H}_2\text{O}_2$  behaves as a reductant for  $\text{F}_2$
205. For the electrochemical cell,  $(\text{M}|\text{M}^\oplus)||(\text{X}^\ominus|\text{X})$ ,  $E_{(\text{M}^\oplus|\text{M})}^\ominus = 0.44$  V, and  $E_{(\text{X}^\ominus|\text{X})}^\ominus = -0.33$  V  
 a)  $\text{M} + \text{X} \rightarrow \text{M}^\oplus + \text{X}^\ominus$  is the spontaneous reaction  
 b)  $\text{M}^\oplus + \text{X}^\ominus \rightarrow \text{M} + \text{X}$  is the spontaneous reaction  
 c)  $E_{\text{cell}}^\ominus = 0.77$  V  
 d)  $E_{\text{cell}}^\ominus = -0.77$  V
206. Which of the following is/are correct? Given, the half-cell emf is  $E_{\text{Cu}^{2+}/\text{Cu}}^\ominus = 0.337$ ,  $E_{\text{Cu}^+/\text{Cu}}^\ominus = 0.521$   
 a)  $E_{\text{Cu}^+/\text{Cu}^{2+}}^\ominus + E_{\text{Cu}^+/\text{Cu}}^\ominus$  is positive                                      b)  $\text{Cu}^+$  disproportionates  
 c) Cu and  $\text{Cu}^{2+}$  compropportionates                                      d) All of the above
207. The tarnishing of silver ornaments in atmosphere is due to  
 a)  $\text{Ag}_2\text{O}$                                       b)  $\text{Ag}_2\text{S}$                                       c)  $\text{Ag}_2\text{CO}_3$                                       d)  $\text{Ag}_2\text{SO}_4$
208. The factor(s) which influence(s) the conductance of solutions?  
 a) Solute-solute interaction  
 b) Solute-solvent interaction  
 c) Temperature  
 d) Solvent-solvent interaction
209. Rusting of iron is catalyzed by  
 a)  $\text{H}^\oplus$   
 b) Dissolved  $\text{CO}_2$  in water  
 c)  $\text{O}_2$   
 d) Impurities present in Fe
210. If same quantity of electricity is passed through three electrolytic cells containing  $\text{FeSO}_4$ ,  $\text{Fe}_2(\text{SO}_4)_3$ , and

Fe(NO<sub>3</sub>)<sub>3</sub>, then

- a) The amount of iron deposited in FeSO<sub>4</sub> and Fe<sub>2</sub>(SO<sub>4</sub>)<sub>3</sub> are equal
- b) The amount of iron deposited in FeSO<sub>4</sub> is 1.5 times of the amount of iron deposited in Fe(NO<sub>3</sub>)<sub>3</sub>
- c) The amount of iron deposited in Fe<sub>2</sub>(SO<sub>4</sub>)<sub>3</sub> and Fe(NO<sub>3</sub>)<sub>3</sub> are equal
- d) The same amount of gas is evolved in all three cases of the anode

211. Given:



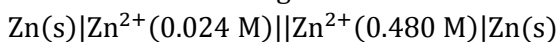
Which of the following statements is/are correct?

- a) X<sub>2</sub>(g) will oxidize both (A) and (B)
- b) A<sup>2+</sup> will oxidize B
- c) The reaction  
 $2X^{\ominus}(1.0 \text{ M}) + A^{2+}(1.0 \text{ M}) \rightarrow X_2(1 \text{ atm}) + A(s)$  will be spontaneous
- d) The oxidizing power of A<sup>2+</sup>, B<sup>+</sup>, and X<sub>2</sub>(g) is in the order X<sub>2</sub> > A<sup>2+</sup> > B<sup>+</sup>

212. In the atmosphere of industrial smog, copper corrodes to form

- a) Basic copper carbonate
- b) Copper sulphide
- c) Basic copper sulphate
- d) Copper oxide

213. Consider the following concentration cell:



Which of the following statements is (are) correct?

- a) The EMF of the cell at 25°C is nearly 0.038 V
- b) The EMF of the cell at 25°C is nearly -0.038 V
- c) If water is added in LHE, so that the [Zn<sup>2+</sup>] is reduced to 0.012 M, the cell voltage increases
- d) If water is added in LHE, so that the [Zn<sup>2+</sup>] is reduced to 0.012 M, the cell voltage remains same

214. In the following electrochemical cell:



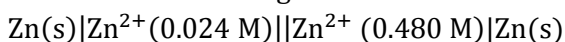
$E_{\text{cell}} = E_{\text{cell}}^{\ominus}$ . This will be when

- a) [Zn<sup>2+</sup>] = [H<sup>+</sup>] = 1 M and p<sub>H<sub>2</sub></sub> = 1 atm
- b) [Zn<sup>2+</sup>] = 0.01 M, [H<sup>+</sup>] = 0.1 M, and p<sub>H<sub>2</sub></sub> = 1 atm
- c) [Zn<sup>2+</sup>] = 1 M, [H<sup>+</sup>] = 0.1 M, and p<sub>H<sub>2</sub></sub> = 1 atm
- d) None of the above

215. A gas X at 1 atm is bubbled through a solution containing a mixture of M Y<sup>⊖</sup> and 1 M Z<sup>⊖</sup> at 25°C. If the reduction potential is Z > Y > X, then

- a) Y will oxidize X only
- b) Y will oxidize Z only
- c) Z will oxidize X and Y
- d) Z will reduce both X and Y

216. Consider the following concentration cell:



Which of the following statement is/are correct?

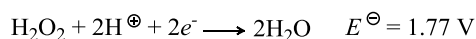
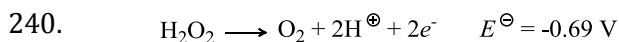
- a) The EMF of the cell at 25°C is nearly +0.039 V
- b) The EMF of the cell at 25°C is nearly -0.039 V
- c) If water is added in LHE, so that the [Zn<sup>2+</sup>] is reduced to 0.012 M, the cell voltage increases
- d) If water is added in LHE, so that the [Zn<sup>2+</sup>] is reduced to 0.012 M, the cell voltage decreases

217. When 4.0 A of current is passed through a 1.0 L, 0.10 M Fe<sup>3+</sup>(aq) solution for 1.0 hour, it is partly reduced to Fe(s) and partly of Fe<sup>2+</sup>(aq). The correct statement(s) is (are):

- a) 0.10 mol of electrons are required to convert all Fe<sup>3+</sup> to Fe<sup>2+</sup>
- b) 0.025 mol of Fe(s) will be deposited
- c) 0.075 mol of iron remains as Fe<sup>2+</sup>
- d) 0.050 mol of iron remains as Fe<sup>2+</sup>

218. Which of the following changes involves oxidation?
- The conversion of ferrous sulphate to ferric sulphate
  - The conversion of  $\text{H}_2\text{S}$  to  $\text{S}$
  - The conversion of  $\text{Cl}_2$  to  $\text{NaCl}$
  - The conversion of  $\text{Zn}$  to  $\text{ZnSO}_4$
219. Consider the cell:  
 $\text{Pt} | \text{H}_2(p_1 \text{ atm}) | \text{H}^\oplus(x_1 \text{ M}) || \text{H}^\oplus(x_2 \text{ M}) | \text{H}_2(p_2 \text{ atm}) | \text{Pt}$   
 The cell reaction be spontaneous if
- $p_1 = p_2$  and  $x_1 > x_2$
  - $p_1 = p_2$  and  $x_1 < x_2$
  - $x_1 = x_2$  and  $p_1 > p_2$
  - $x_1 = x_2$  and  $p_1 > p_2$
220. The amount of ion discharged during electrolysis is directly proportional to :
- Resistance
  - Time
  - Current
  - Chemical equivalent of the ion
221. Of the following, the metal(s) that cannot be obtained by electrolysis of the aqueous salt solution is/are :
- Ag
  - Mg
  - Cu
  - Al
222. In which of the following cells, EMF is greater than  $E_{\text{cell}}^\ominus$ ?
- $\text{Pt}, \text{H}_2(\text{g}) | \text{H}^\oplus(\text{pH} = 5) || \text{H}^\oplus(\text{pH} = 3) | \text{H}_2(\text{g}), \text{Pt}$
  - $\text{Zn}(\text{s}) | \text{Zn}^{2+}(0.2 \text{ M}) || \text{Cu}^{2+}(0.1 \text{ M}) | \text{Cu}(\text{s})$
  - $\text{Cr}(\text{s}) | \text{Cr}^{3+}(0.1 \text{ M}) || \text{Cu}^{2+}(0.2 \text{ M}) | \text{Cu}(\text{s})$
  - $\text{Pt}, \text{H}_2(\text{g}) | \text{H}^\oplus(\text{pH} = 4) || \text{H}^\oplus(\text{pH} = 6) | \text{H}_2(\text{g}) | \text{Pt}$
223. Which of the following changes will cause the free energy of a cell reaction to decrease?  
 $\text{Zn} | \text{ZnSO}_4(\text{aq})(x_1 \text{ M}) || \text{HCl}(\text{aq})(x_2 \text{ M}) / \text{H}_2(\text{g}), \text{Pt}$
- Increase in the volume of  $\text{HCl}$  solution from 100 mL to 200 mL
  - Increase in the pressure of hydrogen from 1 atm to 2 atm
  - Increase in molarity  $x_2$  from 0.1 to 1 M
  - Decrease in molarity  $x_1$  from 1 M to 0.1 M
224. If,  $A + B \rightleftharpoons C + D$ ;  $K_C = K_1$  and  $E^\circ = aV$   
 $2A + 2B \rightleftharpoons 2C + 2D$ ;  $K_C = K_2$  and  $E^\circ = bV$  then,
- $a = b$
  - $K_2 = K_1^2$
  - $a = 2b$
  - $b = a^2$
225. On passing 0.5 mol of electrons through  $\text{CuSO}_4$  and  $\text{Hg}_2(\text{NO}_3)_2$  solutions in series using inert electrodes
- 0.5 mol of  $\text{Cu}$  is deposited
  - 0.5 mol of  $\text{Hg}$  is deposited
  - 0.125 mol of  $\text{O}_2$  is produced
  - 0.5 mol of  $\text{O}_2$  is produced
226. Which of the following aqueous solutions remain alkaline after electrolysis?
- $\text{CH}_3\text{COONa}$
  - $\text{KNO}_3$
  - $\text{NaCl}$
  - $\text{LiF}$
227. The effect of dilution on conductances shows:
- Increase in specific conductance
  - Increase in equivalent conductance
  - Increase in molar conductance
  - Decrease in conductance
228. Coulomb is the quantity of charge defined as :
- One ampere of current passing for 1 sec
  - The charge which deposits 0.001118 of  $\text{Ag}$  on cathode
  - The charge which deposits electrochemical equivalence of metal
  - 1/2 ampere current for two second
229. Standard electrode potential of three metals  $X, Y$  and  $Z$  are  $-1.2 \text{ V}, +0.5 \text{ V}$  and  $-3.0 \text{ V}$  respectively. The reducing power of these metals will be :

- a)  $X > Y > Z$                       b)  $Y > Z > X$                       c)  $Y > X > Z$                       d)  $Z > X > Y$
230. During the electrolysis of aqueous zinc nitrate
- a) Zinc plates out at the cathode                      b) Zinc plates out at the anode  
 c) Hydrogen gas  $H_2$  is evolved at the anode                      d) Oxygen gas  $O_2$  is evolved at the anode
231. Which of the following statements is / are correct about strong electrolyte?
- a) It conducts electric current in solution  
 b) It possesses ions even in the solid state  
 c) It dissociates into ions when dissolved in water because the force of repulsion increases  
 d) The conductivity of an electrolytic solution is directly proportional to the number of ions in unit volume
232. Which are true for a standard hydrogen electrode?
- a) The hydrogen ion concentration is 1 M  
 b) Temperature of  $25^\circ C$   
 c) Pressure of hydrogen is 1 atmosphere  
 d) It contains a metallic conductor which does not adsorb hydrogen
233. During electrolysis,  $O_2(g)$  is evolved at anode in
- a) Dilute  $H_2SO_4$  with Pt electrode  
 b) Aqueous  $AgNO_3$  with Pt electrode  
 c) Dilute  $H_2SO_4$  with Cu electrode  
 d) Fused NaOH with an Fe cathode and Ni anode
234. Select the correct statements if 9.65 A current is passed for 1 hour through the cell:  
 $Ag|Ag^{\oplus}(1 M)||Cu^{2+}(1 M)|Cu$
- a) Ag will oxidize to  $Ag^{\oplus}$  and new  $[Ag^{\oplus}] = 1.36 M$   
 b)  $Ag^{\oplus}$  will reduce to Ag and new  $[Ag^{\oplus}] = 0.64 M$   
 c)  $Cu^{2+}$  will reduce to Cu and new  $[Cu^{2+}] = 0.82 M$   
 d) Cu will oxidize to  $Cu^{2+}$  and new  $[Cu^{2+}] = 0.82 M$
235. The EMF of the following cell:  
 $Cd(s)|CdCl_2(0.10 M)|AgCl(s)|Ag(s)$  is 0.6915 V at  $0^\circ C$  and 0.6753 V at  $25^\circ C$ . The  $\Delta H$  of reaction in kJ at  $25^\circ C$  is
- a) -176                      b) -234.7                      c) 123.5                      d) -167.6
236. Which of the following cells is/are rechargeable or secondary cell(s)?
- a) Ni - Cd cell  
 b) Mercury cell  
 c) Lead storage cell  
 d) Lithium battery
237. Rusting of iron is catalysed by :
- a)  $H^+$   
 b) Dissolved  $CO_2$  in water  
 c)  $O_2$   
 d) Impurities present in Fe
238. For  $I_2 + 2e^- \rightarrow 2I^{\ominus}$ , standard reduction potential = +0.54 V. For  $2Br^{\ominus} \rightarrow Br_2 + 2e^-$ , standard oxidation potential = -1.09 V. For  $Fe \rightarrow Fe^{2+} + 2e^-$ , standard oxidation potential = +0.44 V. Which of the following reaction is (are) spontaneous?
- a)  $Br_2 + 2I^{\ominus} \rightarrow 2Br^{\ominus} + I_2$                       b)  $Fe + Br_2 \rightarrow Fe^{2+} + 2Br^{\ominus}$   
 c)  $Fe + I_2 \rightarrow Fe^{2+} + 2I^{\ominus}$                       d)  $I_2 + 2Br^{\ominus} \rightarrow 2I^{\ominus} + Br_2$
239. 100 mL of buffer of 1 M  $NH_3(aq)$  and 1 M  $NH_4^{\oplus}(aq)$  are placed in two compartments of a voltaic cell separately. A current of 1.5 A is passed through both cells for 20 min. If only electrolysis of water takes place, then
- a) pH of LHE half cell will increase                      b) pH of RHE half cell will increase  
 c) pH of both half cell will increase                      d) pH of both half cell will decrease



Which of the following statements is (are) correct?

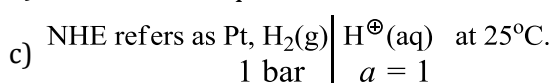
- a)  $\text{H}_2\text{O}_2$  behaves as an oxidant for  $\text{I}_2/\text{I}^{\ominus}$                       b)  $\text{H}_2\text{O}_2$  behaves as an reductant for  $\text{I}_2/\text{I}^{\ominus}$   
 c)  $\text{I}^{\ominus}/\text{I}_2$  behaves as an reductant for  $\text{H}_2\text{O}_2$                       d) None of these is correct

241. Which of the following statements is / are correct?

- a) Zn is precipitated by the addition of Cu powder to a  $\text{ZnSO}_4$  solution  
 b)  $\text{AgNO}_3$  solution can be stored in a copper container  
 c) When  $\text{Cl}_2$  is passed through  $\text{KBr(aq)}$ , the solution becomes coloured  
 d) The addition of a crystal of  $\text{I}_2$  turns a  $\text{KBr}$  solution violet

242. Select the correct statement(s) about NHE

- a)  $E^{\ominus}$  of NHE is arbitrarily assumed to be zero  
 b)  $E^{\ominus}$  of NHE is equal to zero



- d) NHE is very susceptible to dissolved  $\text{O}_2$ ,  $\text{H}_2\text{S}$  and all other reducing agents

243. One faraday is the amount of charge :

- a) That liberates 1 g – equivalent of a metal from its solution  
 b) 96515 coulomb  
 c) That liberates 31.78 g of Cu  
 d) That liberates 1/2 g-atom of Cu

244. The temperature coefficient of the cell is  $\left(\frac{\partial E}{\partial T}\right)_p$ . Choose the correct statement(s)

- a) When  $\left(\frac{\partial E}{\partial T}\right)_p = 0$ , then  $\Delta H = -nF E$   
 b) When  $\left(\frac{\partial E}{\partial T}\right)_p < 0$ , then  $|nFE| > |\Delta H|$   
 c) When  $\left(\frac{\partial E}{\partial T}\right)_p < 0$ , then  $|nFE| > |\Delta H|$  Exothermic reaction  
 d) When  $\left(\frac{\partial E}{\partial T}\right)_p = 0$ , then  $|\Delta H| > |nEF|$  Exothermic reaction

245. During electrolysis of aqueous  $\text{CuBr}_2$  using Pt electrode,

- a)  $\text{Br}_2(\text{g})$  is evolved at anode  
 b)  $\text{Cu(s)}$  is deposited at cathode  
 c)  $\text{Br}_2(\text{g})$  is evolved at anode and  $\text{H}_2(\text{g})$  at cathode  
 d)  $\text{H}_2(\text{g})$  is evolved at anode

3.ELECTROCHEMISTRY

**: ANSWER KEY :**

1) a	2) b	3) a	4) a	a, c
5) d	6) a	7) c	8) a	13) a, c 14) a, b, d 15) a, b 16)
9) c	10) c	11) a	12) b	b, d
13) a	14) c	15) c	16) c	17) a, d 18) a, b, d 19) b, c 20)
17) d	18) c	19) d	20) b	a, c
21) b	22) b	23) b	24) a	21) c, d 22) a, b, d 23) a, b, c 24)
25) d	26) c	27) c	28) b	a, b, c, d
29) a	30) d	31) c	32) b	25) a, b, c 26) b, c 27) a, c, d 28)
33) a	34) b	35) c	36) a	a, b, c, d
37) b	38) c	39) c	40) b	29) b, c 30) a, b, c 31) a, b 32)
41) c	42) b	43) a	44) b	a, b, c, d
45) d	46) d	47) d	48) a	33) a, b, c 34) a, b, c 35) b, c, d 36)
49) b	50) c	51) a	52) a	a, b, d
53) d	54) b	55) a	56) a	37) a, c, d 38) a, d 39) a, b, c 40)
57) b	58) d	59) c	60) a	c, d
61) d	62) a	63) d	64) c	41) a, d 42) a, c, d 43) a, b 44)
65) b	66) c	67) a	68) a	a, b
69) b	70) d	71) a	72) c	45) a, c 46) c, d 47) c, d 48)
73) c	74) c	75) a	76) a	a, c, d
77) b	78) a	79) b	80) b	49) c, d 50) a, b 51) a, d 52)
81) b	82) c	83) b	84) d	b, c
85) a	86) c	87) b	88) a	53) a, c, d 54) a, d 55) b, c 56)
89) a	90) c	91) d	92) c	a, b
93) c	94) d	95) b	96) a	57) a, b 58) a, b, c, d 59) a, b, c, d
97) a	98) d	99) d	100) c	60) b, c, d
101) a	102) c	103) a	104) c	61) a, b, d 62) a, c 63) a, c 64)
105) b	106) c	107) a	108) d	a, b
109) d	110) d	111) c	112) c	65) a, c 66) a, c 67) a, b, c 68)
113) d	114) a	115) d	116) a	a, b, d
117) d	118) c	119) b	120) b	69) b, c 70) b, c, d 71) b, d 72) a
121) a	122) a	123) b	124) d	73) c, d 74) a, b 75) b, c 76)
125) a	126) c	127) d	128) a	a, c
129) b	130) b	131) b	132) c	77) b, c 78) a, b, c, d 79) c, d
133) c	134) b	135) d	136) c	80) a, d
137) a	138) c	139) c	140) c	81) a, b, c, d 82) a, b, c 83) a, b
141) c	142) d	143) d	144) b	84) a, c
145) a	146) b	147) b	148) c	85) d 86) a, c, d 87) a, b, c, d
149) c	150) a	1) a, c	2)	88) a, b, c
a, b	3) a, d	4) b, d		89) b 90) a, c 91) c, d 92)
5) c	6) a, b, c	7) b, c, d	8)	a, c, d
b, c				93) a, b, c, d 94) a, b, c 95) a, b
9) a, c, d	10) a, b, c	11) a, b, c	12)	

**: HINTS AND SOLUTIONS :**1 **(a)**

The smaller the reduction potential of a substance, the more is its reducing power. ( $Y > Z > X$ )

2 **(b)**

$$\alpha = \frac{\Lambda_m}{\Lambda_m^0} = \frac{7.8}{390} = 0.02$$

$$K_a(\text{CH}_3\text{COOH}) = \frac{c\alpha \cdot c\alpha}{c - c\alpha}$$

$$= \frac{c\alpha^2}{1 - \alpha} = \frac{0.04 \times (0.02)^2}{1 - 0.02} \approx 1.6 \times 10^{-5}$$

$$\text{Use: } pK_a + pK_b = pK_w = 14$$

$$\Rightarrow pK_b = 14 - pK_a = 14 - 4.8 = 9.2$$

3 **(a)**

Equivalent point = 15 mL  $\text{H}_2\text{SO}_4$  (from graph)

Using mEq of acid = mEq of base

$$\Rightarrow (2 \times 0.2) \times 15 = (1 \times M_{\text{KOH}}) \times 20 \Rightarrow M_{\text{KOH}} = 0.3 \text{ M}$$

4 **(a)**

$$\kappa = \frac{1}{R} G^* = G G^* = \frac{1}{250} 1.15 = 0.0046 \text{ S cm}^{-1}$$

$$\Lambda_{\text{eq}} = \frac{\kappa 1000}{N} = \frac{0.00461000}{1} = 4.6 \text{ ohm}^{-1} \text{ cm}^2 \text{ eq}^{-1}$$

5 **(d)**

$$G^* = \frac{l}{a} = \frac{1.5 \text{ cm}}{0.75 \text{ cm}^2} = 2.0 \text{ cm}^{-1}$$

6 **(a)**

Since the reduction potential of  $\text{Br}_2|2\text{Br}^\ominus$  is greater than the reduction potential of  $\text{I}_2|2\text{I}^\ominus$ , so  $\text{I}^\ominus$  ions will reduce  $\text{Br}_2$

7 **(c)**

$$E_{\text{cell}} = \frac{0.0591}{1} \log \frac{[\text{H}^\oplus]_{\text{RHS}}}{[\text{H}^\oplus]_{\text{LHS}}}$$

$$[\text{For } \text{H}_2\text{O}, [\text{H}^\oplus] = [\text{OH}^\ominus] = 10^{-7} \text{ M}]$$

$$= 0.059 \log \frac{10^{-3}}{10^{-7}} = 0.59 \times 4 = 0.236 \text{ V}$$

$$\text{Or } E_{\text{cell}} = -0.059 (\text{pH}_c - \text{pH}) = -0.059(3 - 7) = 0.236 \text{ V}$$

8 **(a)**

$$E = E^\ominus - 0.059 \text{ pH}$$

$$= 0.699 - 0.059 \times 10 = 0.109 \text{ V}$$

9 **(c)**

$$\text{pOH} = 4, \text{pH} = 14 - 4 = 10$$

$$E = -0.059 \times 10 = -0.59 \text{ V}$$

10 **(c)**

**Faraday's first law**

The amount of substance liberated at the electrode is directly proportional to the quantity of electricity passed

Electrochemical equivalent (ECE) is the amount of substance liberated on passing 1 C of electricity

The ECEs of hydrogen, silver, and copper are 0.00001036, 0.001118, and 0.0003292, respectively

**Faraday's second law**

If same quantity of electricity (same current for same time) is passed through different electrolytes, the amount of substance liberated at different electrodes is in the ratio of their respective electrochemical equivalent (or equivalent weights)

11 **(a)**

$$\text{Initial mEq of } \text{Cu}^{2+} = 500 \times 1$$

$$\text{Electricity passed} \equiv \frac{2 \times 3600}{96500} = 0.0746 \text{ Eq}$$

$$\equiv 74.6 \text{ mEq}$$

$$\text{mEq of } \text{CuCl}_2 \text{ remaining} \equiv 500 - 74.6 = 425.4 \text{ mEq}$$

$$\therefore N = \frac{425.5}{500} = 0.85$$

12 **(b)**

Let  $A$  be the atomic weight of metal

$$1 \text{ mol of metal} = A \text{ g}$$

$$96500 \text{ C} \Rightarrow A \text{ g}$$

$$9.65 \times 10 \times 60 \Rightarrow \frac{A}{96500} \times 9.65 \times 10 \times 60$$

$$\therefore A \times 0.06 = 3 \text{ g}$$

$$A = \frac{3}{0.06} = 50$$

14 **(c)**

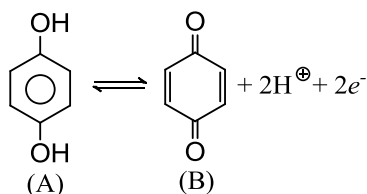
In this electrode,

$[A] = [B]$  in quinhydrone electrode

$$\text{Hence, } Q = [\text{H}^\oplus]^2$$



Hence,  $Q = [H^{\oplus}]^2$



$$E = E^{\ominus} - \frac{0.0591}{2} \log [H^{\oplus}]^2$$

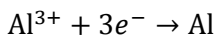
$$= E^{\ominus} - 0.0591 \log [H^{\oplus}]$$

$$= E^{\ominus} + 0.0591 \text{ pH}$$

$$= 1.30 + 0.059 \times 2$$

$$= 1.42 \text{ V}$$

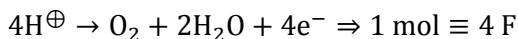
15 (c)



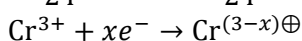
$$1 \text{ mol of Al} = 27 \text{ g} = 3 \text{ F}$$

$$\therefore 13.5 \text{ g of Al} = \frac{3}{27} \times 13.5 \equiv 1.5 \text{ F}$$

16 (c)



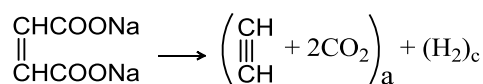
$$\Rightarrow \frac{0.6}{24} \text{ mol} \equiv 4 \times \frac{0.6}{24} \times 0.1 \text{ F}$$



$$\Rightarrow 1 \text{ mol} \equiv x \text{ F} \Rightarrow \frac{2.6}{52} \text{ mol} \equiv \frac{x}{20} \text{ F} \equiv 0.1 \text{ F} \Rightarrow x = 2$$

$$\Rightarrow \text{Oxidation state of Cr}^{(3-2)} = \text{Cr}^{+1} = +1$$

18 (c)

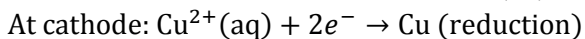
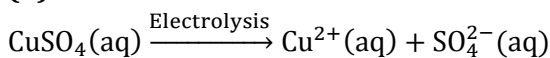


$$2 \text{ F} = 2 \times 96500 \text{ C} \Rightarrow 4 \times 22.4 \text{ L}$$

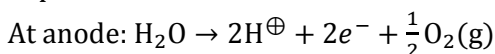
$$0.2 \times 965 \text{ C} \Rightarrow \frac{4 \times 22.4 \times 0.2 \times 965}{2 \times 96500}$$

$$\Rightarrow 0.0896 \text{ L} \Rightarrow 89.6 \text{ mL}$$

19 (d)

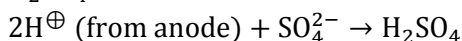


The blue colour of  $\text{CuSO}_4$  disappears due to the deposition of Cu on Pt electrode



Since oxidation potential  $\text{H}_2\text{O} >$  oxidation potential of  $\text{SO}_4^{2-}$ , so oxidation of  $\text{H}_2\text{O}$  occurs and  $\text{O}_2(\text{g})$  is evolved at anode

The colourless solution is due to the formation of  $\text{H}_2\text{SO}_4$  as follows:



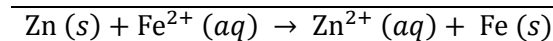
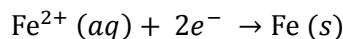
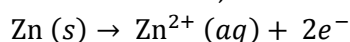
20 (b)

Carbon is a powerful reducing agent from X and Y but not from Z. Y is more reactive than X. Therefore, reactivity series:  $Z > Y > X$

21 (b)

For all  $\text{Zn} | \text{Zn}^{2+} (\text{A} = 0.1 \text{ M}) || \text{Fe}^{2+} (\text{A} = 0.01 \text{ M}) | \text{Fe}$ .

The cell reaction,



On applying Nernst equation,

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

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

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

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

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

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

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

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

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

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

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

22 (b)

Volume to be covered

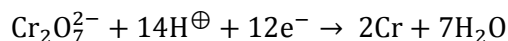
$$= (0.25 \times 10^4 \text{ cm}^2) \times (2 \times 10^{-3} \text{ cm}) = 5 \text{ cm}^3$$

$$\Rightarrow \text{Mass of Cr to be deposited} = 5 \times 7.19 \text{ g}$$

$$= 35.95 \text{ g}$$

$$\equiv \frac{35.95}{52} \text{ mol}$$

$$= 0.69 \text{ mol}$$

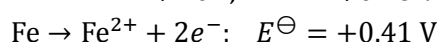
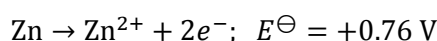
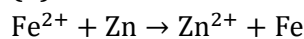


Now, 2 mol Cr deposited  $\equiv 12 \text{ F}$  electricity

$$\Rightarrow 0.59 \text{ mol Cr deposited} \equiv \frac{12}{2} \times 0.069 \text{ F}$$

$$= \frac{It}{96500} = \frac{75 \times t}{96500} \quad t = 1.48 \text{ h} \approx 1.5 \text{ h}$$

23 (b)



These are oxidation potentials

Reduction potentials are equal and opposite

Fe forms cathode and Zn forms anode

$$E_{\text{cell}}^{\ominus} = (E^{\ominus}_{\text{red}})_{\text{c}} + (E^{\ominus}_{\text{oxid}})_{\text{a}}$$

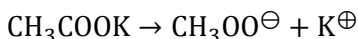
$$= (-0.41 + 0.76) \text{ V}$$

$$= 0.35 \text{ V}$$

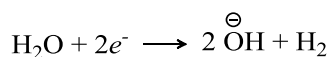
24 (a)

Follow from the definition

25 (d)

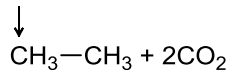
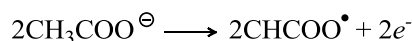


At cathode,  $\text{H}_2\text{O}$  undergoes reduction of give  $\text{H}_2$



$$2 \text{ F} \equiv 1 \text{ mol of } \text{H}_2$$

At anode,



$$2 \text{ F} \equiv 1 \text{ mol of } \text{C}_2\text{H}_6 \text{ and } 2 \text{ mol of } \text{CO}_2$$

Total volume at cathode and anode

$$= 4 \text{ mol} = 4 \times 22.4 \text{ L}$$

$$2 \text{ F} = 2 \times 96500 \text{ C} \Rightarrow 4 \times 22.4 \text{ L}$$

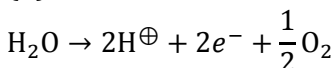
$$0.1 \times 965 \text{ C} \Rightarrow \frac{4 \times 22.4 \times 0.1 \times 965}{2 \times 96500}$$

$$\Rightarrow 0.0448 \text{ L} = 44.8 \text{ mL}$$

26 (c)

Factual statement

28 (b)



$$2\text{F} \equiv 1 \text{ mol of } \text{H}_2\text{O} = \frac{1}{2} \text{ mol of } \text{O}_2$$

$$\therefore 2 \times 96500 \text{ C} = 1.93 \times 10^5 \text{ C}$$

29 (a)

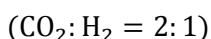
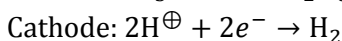
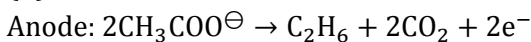
$$\text{By } \Delta S = \frac{\Delta H - \Delta G}{T};$$

$$\Delta H = -nF \left[ E - T \left( \frac{\delta E}{\delta T} \right)_p \right] \text{ and } \Delta G = -nEF$$

$$\Delta S = -nF \left( \frac{\delta E}{\delta T} \right)_p = 3 \times 96500 \times \left( -\frac{0.0006}{10} \right)$$

$$= -17.37 \text{ J K}^{-1}$$

30 (d)



31 (c)

$$E_{\text{cell}} = E_{\text{OP}_{\text{D}_2}} = E_{\text{RP}_{\text{H}_2}}$$

(OP = oxidation potential, RP = reduction potential)

$$= E_{\text{OP}_{\text{D}_2|\text{D}^{\oplus}}} - \frac{0.059}{2} \log[\text{D}^{\oplus}]^2 + E_{\text{RP}_{\text{H}^{\oplus}|\text{H}_2}} + 0.059 \log[\text{H}^{\oplus}]^2$$

$$= 0 = 0.003 - \frac{0.059}{2} \log \frac{[\text{D}^{\oplus}]^2}{[\text{H}^{\oplus}]^2}$$

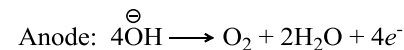
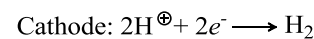
$$= \frac{[\text{D}^{\oplus}]}{[\text{H}^{\oplus}]} = 1.124$$

32 (b)

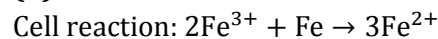
KCl is used since its conductivity is known accurately at various concentrations and at different temperatures

33 (a)

NaF:



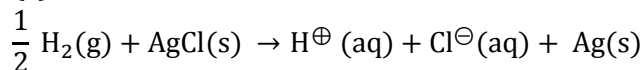
34 (b)



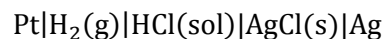
$$E_{\text{cell}}^{\ominus} = 0.77 - (-0.44) = 1.21 \text{ V}$$

$\Rightarrow \text{Fe}^{3+}$  and Fe will reduce

35 (c)



There is oxidation of hydrogen and reduction of silver therefore, it occurs in the following cell:



37 (b)

Factual statement

40 (b)

Factual statement

41 (c)

It is a concentration cell, therefore,  $E_{\text{cell}}^{\ominus} = 0$

$$\text{pH of } \text{HNO}_3 = -\log 10^{-1} = 1$$

$$\text{pOH of } \text{NaOH} = -\log 10^{-3} = 3$$

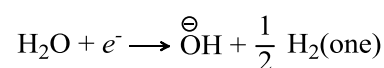
$$\therefore \text{pH of } \text{NaOH} = 14 - 3 = 11$$

$$\therefore E = -0.059(\text{pH}_{\text{c}} - \text{pH}_{\text{a}})$$

$$= -0.059(11 - 1) = -0.59 \text{ V}$$

42 (b)

At cathode:



At anode:



$\text{H}_2$  and  $\text{CO}_2$  at anode and  $\text{H}_2$  at cathode. So, number of gases as in (b)

43 (a)

$$1 \text{ F} = 96500 \text{ C} = 1 \text{ Eq of } \text{Fe}^{2+}$$

$$2 \times 3600 \text{ C} \equiv \frac{2 \times 3600}{96500} \equiv 0.0746 \text{ Eq}$$

$$\therefore V \times 0.1 \equiv 0.0746$$

$$V = 0.746$$

44 (b)

It is a concentration cell, therefore,  $E_{\text{cell}}^{\ominus} = 0$

For weak base  $\text{NH}_4\text{OH}$

$$\text{pOH} = \frac{1}{2}(4.74 - \log 10^{-3}) = 3.87$$

$$\therefore \text{pH} = 14 - 3.87 = 10.13$$

For strong base  $\text{NaOH}$

$$\text{pOH} = 3, \text{pH} = 14 - 3 = 11$$

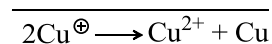
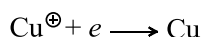
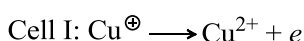
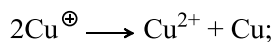
$$\therefore E_{\text{cell}} = -0.059(\text{pH}_c - \text{pH}_a)$$

$$= -0.059(11 - 10.13)$$

$$= -0.059 \times 0.87$$

$$= -0.05 \text{ V}$$

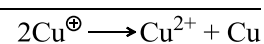
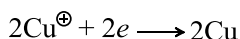
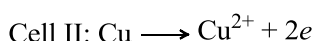
45 (d)



$$E^{\ominus} = -0.15 + 0.50$$

$$= +0.35 \text{ V}$$

$$n = 1$$

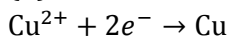


$$E^{\ominus} = -0.325 + 0.50$$

$$= +0.175 \text{ V}$$

$$n = 2$$

46 (d)



$$E_{\text{Cu}^{2+}|\text{Cu}} = E^{\ominus}_{\text{Cu}^{2+}|\text{Cu}} - \frac{0.059}{2} \log \frac{1}{[\text{Cu}^{2+}]}$$

$$= E^{\ominus}_{\text{Cu}^{2+}|\text{Cu}} - \frac{RT}{2F} \ln [\text{Cu}^{2+}]$$

$$\text{Intercept} = 0.34 \Rightarrow E^{\ominus}_{\text{Cu}^{2+}|\text{Cu}} = 0.34 \text{ V}$$

$$\Rightarrow E_{\text{Cu}^{2+}|\text{Cu}} = 0.34 + \frac{0.059}{2} \log 0.1 = 0.31 \text{ V}$$

$$\Rightarrow E_{\text{Cu}|\text{Cu}^{2+}} = -E_{\text{Cu}^{2+}|\text{Cu}} = -0.34 + \frac{0.059}{2} \text{ V}$$

47 (d)

$$\Lambda_m = \frac{\kappa \times 1000}{M} = \frac{0.0110 \times 1000}{0.05}$$

$$= 220 \text{ S cm}^2 \text{ mol}^{-1}$$

48 (a)

Hg is below H in electrochemical series

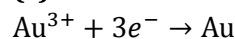
49 (b)

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

$$\text{From line OA} = E^{\ominus}_{\text{cell}} = 1.10 \text{ V}$$

$$\text{If } \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} = 10^{-2} \text{ M, then } E_{\text{cell}} = 1.1591 \text{ V}$$

50 (c)



$$\frac{W}{197/3} = \frac{\text{Charge}}{96500}$$

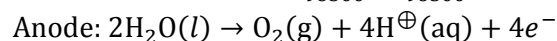
$$\therefore \text{Charge} = \frac{1.314 \times 3 \times F}{197} = 0.02 \text{ F}$$

51 (a)

Fused  $\text{NaCl}$  will give reduction of  $\text{Na}^{\oplus}$  ion to  $\text{Na}$  at cathode and oxidation of  $\text{Cl}^{\ominus}$  to  $\text{Cl}_2$  at anode

52 (a)

$$\text{Number of Faradays} = \frac{It}{96500} = \frac{2 \times 965}{96500} = 0.02$$



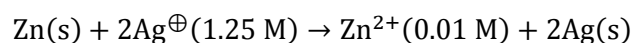
$$\Rightarrow 4 \text{ F electricity} = 1 \text{ mol O}_2 \equiv 22.4 \text{ L O}_2 \text{ at STP}$$

$$\Rightarrow 0.02 \text{ F electricity} = 22.4 \times \frac{0.02}{4}$$

$$= 0.112 \text{ L O}_2 \text{ at STP}$$

53 (d)

Write cell reaction:

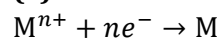


$$Q = \frac{[\text{Zn}^{2+}]}{[\text{Ag}^{\oplus}]^2} = \frac{0.01}{(1.25)^2} = 64 \times 10^{-3}$$

54 (b)

For a cell to be spontaneous ( $E_{\text{cell}} = +ve$  or  $\Delta G = -ve$  (exergonic)  $p_1 > p_2$  or  $c_2 > c_1$ . Factual statement

55 (a)



$$\text{Mole of metal deposited} = \frac{2.977}{106.4} = 0.27 \text{ mol}$$

$$1 \text{ mol of M deposited} = n \text{ Faraday}$$

$$= n \times 96500 \text{ C}$$

$$0.027 \text{ mol} = n \times 96500 \times 0.027$$

$$\therefore n \times 96500 \times 0.027 = 10800 \text{ C}$$

$$\therefore n = 4$$

56 (a)

Reduction and oxidation of  $\text{H}_2\text{O}$  occurs to give  $\text{H}_2$  at cathode and  $\text{O}_2$  at anode

58 (d)

Factual statement

59 (c)

It is a concentration cell, therefore,  $E^{\ominus}_{\text{cell}} = 0$

Acidic buffer at anode:

$$\therefore \text{pH} = \text{pK}_a + \log \left[ \frac{\text{Salt}}{\text{Acid}} \right]$$

$$= 4.74 + \log \frac{0.1}{0.01} = 4.74 + 1 = 5.74$$

Basic buffer at cathode:

$$\therefore \text{pOH} = \text{pK}_b + \log \left[ \frac{\text{Salt}}{\text{Base}} \right]$$

$$= 4.74 + \log \frac{0.2}{0.1}$$

$$= 4.74 + 0.03 = 5.04$$

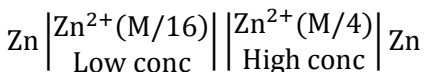
$$\begin{aligned} \therefore \text{pH} &= 14 - 5.04 = 8.96 \\ E_{\text{cell}} &= -0.59(\text{pH}_c - \text{pH}_a) \\ &= -0.059(8.96 - 5.74) \\ &= -0.059 \times 3.22 = -0.189 \text{ V} \end{aligned}$$

60 (a)

A salt of strong acid ( $S_A$ ) and strong base ( $S_B$ ), e.g.,  $\text{Ca}(\text{NO}_3)_2$  is completely ionized and hence acts as a strong electrolyte

61 (d)

For concentration cell,  $E^\ominus = 0$



$$E = -\frac{0.059}{2} \log \frac{(\text{Zn}^{2+})_a}{(\text{Zn}^{2+})_c}$$

$$= -\frac{0.059}{2} \log \frac{\text{M}/16}{\text{M}/4}$$

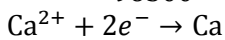
$$= -\frac{0.059}{2} \log \left( \frac{1}{4} \right)$$

Solve,  $E = 0.0178 \text{ V}$

62 (a)

$25 \times 10^{-3} \text{ A} \times 60 \text{ C}$  of charge carried by

$$= \frac{25 \times 10^{-3} \text{ A} \times 60 \text{ s}}{96500} = 1.55 \times 10^{-5} \text{ mol } e^-$$



2 mol of  $e^-$  will produce =  $6 \times 10^{23}$  atoms of Ca

$1.55 \times 10^{-5} \text{ mol } e^-$  will produce

$$= \frac{1.55 \times 10^{-5} \times 6 \times 10^{23}}{2}$$

$$= 4.68 \times 10^{18} \text{ atom of Ca}$$

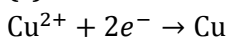
63 (d)

$\text{Cr}_2\text{O}_7^{2-}$  reduces to  $2\text{Cr}^{3+}$ , so must be represented at cathode.  $\text{H}_2$  is oxidized to  $2\text{H}^\oplus$ , so must be represented at anode. So the cell representation is as shown in (d)

65 (b)

Factual statement

66 (c)



$2 \text{ F} = 1 \text{ mol of Cu} = 63.5 \text{ g of Cu}$

$$0.5 \text{ F} = \frac{63.5}{2} \times 0.5 = 15.8 \text{ g}$$

67 (a)

$$\text{Zlt or Elt.} \left[ \begin{array}{l} \text{Ew of } \text{O}_2 = \frac{32}{4} \\ \text{n factor} = 4 \end{array} \right]$$

$$W_{(\text{O}_2)} = \frac{32 \times 2 \times 965}{4 \times 96500}$$

$$= 0.16 \text{ g} = 0.005 \text{ mol} = 0.112 \text{ L at STP}$$

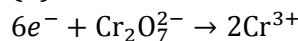
69 (b)

Specific resistance =  $x$

$$\therefore \text{Specific conductance (or conductivity)} = \kappa = \frac{1}{x}$$

$$\therefore \Lambda_{\text{eq}} = \frac{\kappa \times 1000}{xN}$$

70 (d)

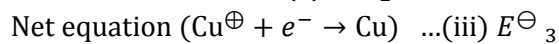
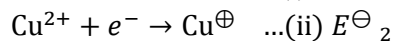
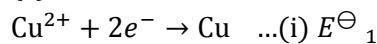


So charge =  $6 \times F = 6 \times 96500 \text{ C}$

71 (a)

Factual statement

72 (c)



Is obtained by equation (i) - (ii)

$$E^\ominus_3 = \frac{n_1 E^\ominus_1}{n_3} = \frac{2 \times 0.337 - 1 \times 0.153}{1} = 0.521 \text{ V}$$

73 (c)

$\text{H}_2 - \text{O}_2$  cell is called fuel cell

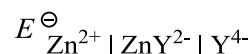
74 (c)

Factual statement of Faraday's second law of electrolysis

75 (a)

$$E^\ominus_{\text{Zn}^{2+}|\text{ZnY}^{2-}|\text{Y}^{4-}} = E^\ominus_{\text{Zn}^{2+}|\text{Zn}} + \frac{0.059}{2} \log K$$

$$K_f = \frac{[\text{ZnY}^{2-}]}{[\text{Y}^{4-}]} \Rightarrow K = \frac{1}{K_f}$$



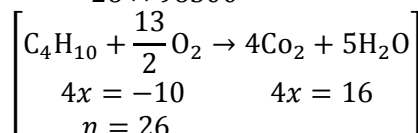
$$= -0.76 + \frac{0.059}{2} \log \frac{1}{3.2 \times 10^{16}} = -1.25 \text{ V}$$

76 (a)

$\Delta G = \Delta G^\ominus = -2.75 \times 10^6 \text{ J mol}^{-1}$  (as  $P = 1 \text{ bar}$  and  $T = 298 \text{ K}$ )

$$-\Delta G^\ominus = nE^\ominus F$$

$$E^\ominus = \frac{2.75 \times 10^6}{26 \times 96500}$$

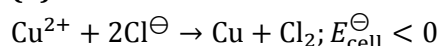


$$E^\ominus = 1.096 \text{ V}$$

$$\text{Also, } E^\ominus = \frac{0.059}{n} \log K_p \Rightarrow 1.096 = \frac{0.059}{26} \log K_p$$

$$K_p = 9.55 \times 10^{482}$$

77 (b)



$\Rightarrow$  Reactions which have  $E^\ominus_{\text{cell}} < 0$  (non-spontaneous) can be made to occur in an electrolytic cell

78 (a)

$$\text{Use: } \lambda_{\text{eq}} = 1000 \times \frac{\kappa}{c}$$

$$\Rightarrow \lambda_{\text{eq}} = \frac{1000 \times 26 \times 10^{-2}}{2} = 1.3 \times 10^2 \text{ S cm}^{-1}$$

81 (b)

$$\text{Use: } \Delta S = nF \left( \frac{dE_{\text{cell}}}{dT} \right)_P = 2 \times 96500 \times 3.4 \times 10^{-4} = 65.2 \text{ J K}^{-1}$$

82 (c)

In electrolytic cell, flow of electron is possible from cathode to anode through internal supply.

83 (b)

Only Fe has less reduction potential than Cu

84 (d)

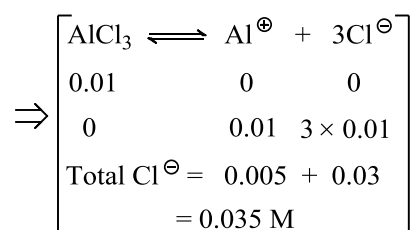
All are factual statements

85 (a)

$$\text{After mixing concentration of } [\text{NaCl}] = \frac{0.01}{1} = 0.005$$

$$\text{After mixing concentration of } [\text{AlCl}_3] = \frac{0.02}{2} = 0.01$$

$$[\text{Na}^{\oplus}] = 0.05 \text{ M } [\text{Al}^{3+}] = 0.01 \text{ M}, [\text{Cl}^{\ominus}] = 0.0035$$



$$\mu = \frac{1}{2} [0.005(1)^2 + 0.01(3)^2 + 0.35(1)^2] = 0.065$$

86 (c)

Total charge = 2

$$\text{Number of equivalent of ion} = \frac{\text{Charge on the ion}}{\text{Total charge}}$$

$$\therefore \text{Eq of } \left( \begin{array}{c} \text{COO}^{\ominus} \\ | \\ \text{COO}^{\ominus} \end{array} \right) = \frac{2}{2} = 1$$

$$\text{Eq of Na}^{\oplus} = \frac{1}{2}, \text{ Eq of K}^{\oplus} = \frac{1}{2}$$

$$\therefore \lambda_{\text{eq}}^{\circ} \left( \begin{array}{c} \text{COONa} \\ | \\ \text{COOK} \end{array} \right)$$

$$= \lambda_{\text{eq}}^{\circ} \left( \begin{array}{c} \text{COO}^{\ominus} \\ | \\ \text{COO}^{\ominus} \end{array} \right) + \frac{1}{2} \lambda_{\text{Na}^{\oplus}}^{\circ} + \frac{1}{2} \lambda_{\text{K}^{\oplus}}^{\circ}$$

$$= 74 + \frac{50}{2} + \frac{50}{2} = 135.5$$

87 (b)

At anode:

0.1 M acetic acid + 0.1 M sodium acetate  $\Rightarrow$  Acidic buffer

$$\therefore \text{pH}_a = \text{pK}_a + \log \left[ \frac{\text{Salt}}{\text{Acid}} \right] \quad (\text{pK}_a(\text{CH}_3\text{COOH}) = 5)$$

$$= 5 + \log \frac{0.1}{0.1} = 5$$

At cathode:

0.1 M formic acid + 0.1 M sodium formate  $\Rightarrow$  Acidic buffer

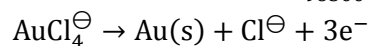
$$\therefore \text{pH}_c = \text{pK}_a + \log \left( \frac{0.1}{0.1} \right) \quad [\text{pK}_a(\text{HCOOH}) = 4]$$

$$\text{pH}_c = 4$$

$$E_{\text{cell}} = -0.059(\text{pH}_c - \text{pH}_a) = -0.059(4 - 5) = 0.059 \text{ V}$$

88 (a)

$$\text{Number of Faradays} = \frac{It}{96500} = \frac{0.3 \times 15 \times 60}{96500}$$



$$3F \equiv 1 \text{ mol Au}$$

$$\Rightarrow 2.8 \times 10^{-3} F \equiv \frac{2.8}{3} \times 10^{-3} \text{ mol Au}$$

$$\equiv \frac{2.8}{3} \times 197 \times 10^{-3} \text{ g Au}$$

$$= 0.184 \text{ g Au}$$

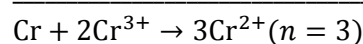
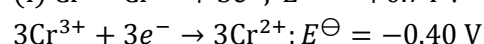
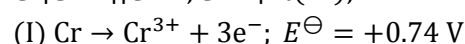
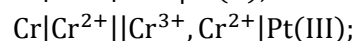
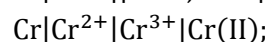
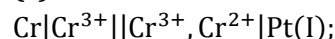
89 (a)

$$\text{Total mass of } \text{C}_2\text{H}_6 + 2\text{CO}_2 + \text{H}_2 \equiv 30 + 88 + 2 = 120 \text{ g}$$

$$\therefore 2 \times 96500 \text{ C} \Rightarrow 120 \text{ g}$$

$$0.01 \times 965 \text{ C} \Rightarrow \frac{120 \times 0.1 \times 965}{2 \times 96500} \Rightarrow 0.06 \text{ g}$$

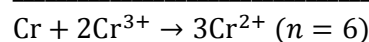
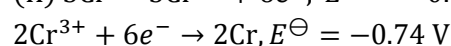
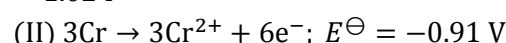
90 (c)



$$E^{\ominus} = 0.74 - 0.4 = 0.34 \text{ V}$$

$$\Delta G^{\ominus} = 3 \times 0.34 \times F$$

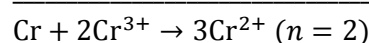
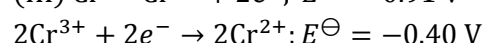
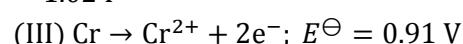
$$= 1.02 \text{ F}$$



$$E^{\ominus} = 0.91 - 0.74 = 0.17 \text{ V}$$

$$\Delta G^{\ominus} = 0.17 \times 6 \times F$$

$$= 1.02 \text{ F}$$

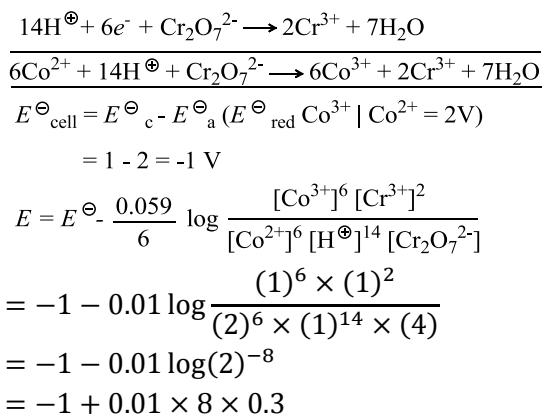


$$E^{\ominus} = 0.91 - 0.40 = 0.51 \text{ V}$$

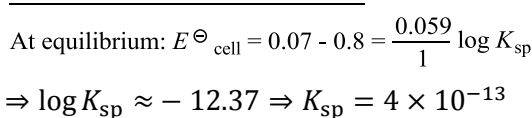
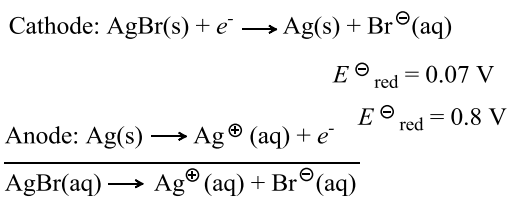
$$\Delta G^{\ominus} = 2 \times 0.51 \times F$$

- $= 1.02 \text{ F}$
- 91 **(d)**  
Factual statement
- 92 **(c)**  
 $E_{\text{oxidation}} = 0.059 \text{pH} = 0.059 \times 10 = 0.59 \text{ V}$
- 94 **(d)**  
 $ZIt \text{ or } EIt. \left[ \begin{array}{l} \text{Ew of } O_2 = \frac{32}{4} \\ \text{n factor} = 4 \end{array} \right]$   
 $W_{(O_2)} = \frac{32 \times 2 \times 965}{4 \times 96500}$   
 $0.224 \text{ mL } O_2 = 3.2 \times 10^{-4} \text{ g}$   
 $I = \frac{W}{Zt} = \frac{3.2 \times 10^{-4} \times 96500 \times 4}{32 \times 1} = 3.86 \text{ A}$
- 95 **(b)**  
 $(E_{\text{Cu}^{2+}|\text{Cu}})_1 = E_{\text{Cu}^{2+}|\text{Cu}}^\ominus + \frac{0.059}{2} \log[\text{Cu}^{2+}]_1$   
 $[\text{Cu}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Cu}(\text{s})]$   
and  $(E_{\text{Cu}^{2+}|\text{Cu}})_2 = E_{\text{Cu}^{2+}|\text{Cu}}^\ominus + \frac{0.059}{2} \log[\text{Cu}^{2+}]_2$   
 $\Rightarrow \Delta E_{\text{Cu}^{2+}|\text{Cu}} = -\frac{0.059}{2} \text{ V}$
- 96 **(a)**  
A saturated solution of  $\text{KNO}_3$  is used to make salt-bridge because the velocities of both  $\text{K}^\oplus$  and  $\text{NO}_3^\ominus$  are nearly the same. This keeps a uniform distribution of charges on both sides
- 97 **(a)**  
Efficiency  $= \frac{\Delta G^\ominus}{\Delta H^\ominus} = -\frac{nE^\ominus F}{\Delta H} = 80$   
 $E^\ominus = \frac{80 \times (-300) \times 10^3}{2 \times 96500 \times 100} = 1.24 \text{ V}$
- 98 **(d)**  
When a lead storage battery is discharged, sulphuric acid is consumed
- 99 **(d)**  
 $\text{Fe}^{3+} + 3e^- \rightarrow \text{Fe}; \Delta G_1 = -3F E_{\text{Fe}^{3+}|\text{Fe}}^\ominus$   
 $\text{Fe}^{2+} + 2e^- \rightarrow \text{Fe}; \Delta G_2 = -2F E_{\text{Fe}^{2+}|\text{Fe}}^\ominus$   
 $\text{Fe}^{3+} + e^- \rightarrow \text{Fe}^{2+}; \Delta G_3 = -3F E_{\text{Fe}^{3+}|\text{Fe}}^\ominus - 2E_{\text{Fe}^{2+}|\text{Fe}}^\ominus$   
 $E_{\text{Fe}^{3+}|\text{Fe}^{2+}} = 3E_{\text{Fe}^{3+}|\text{Fe}}^\ominus - 2E_{\text{Fe}^{2+}|\text{Fe}}^\ominus$   
 $= 3(-0.36) - 2(-0.439) \text{ V}$
- 100 **(c)**  
 $E_{\text{Cr}_2\text{O}_7^{2-}|\text{Cr}^{3+}}^\ominus = 1.33 \text{ V}; E_{\text{MnO}_2|\text{Mn}^{2+}}^\ominus = 1.23 \text{ V};$   
 $E_{\text{MnO}_4^\ominus|\text{Mn}^{2+}}^\ominus = 1.51 \text{ V}; E_{\text{Cl}_2|\text{Cl}^\ominus}^\ominus = 1.36 \text{ V};$   
 $\therefore E_{\text{MnO}_4^\ominus|\text{Mn}^{2+}}^\ominus$  is highest,  $\text{KMnO}_4$  is the strongest oxidizing agent
- 101 **(a)**  
 $K_{\text{Co}_2[\text{Fe}(\text{CN})_6]} = 2.06 \times 10^{-6} - 0.41 \times 10^{-6}$   
 $= 1.65 \times 10^{-6}$

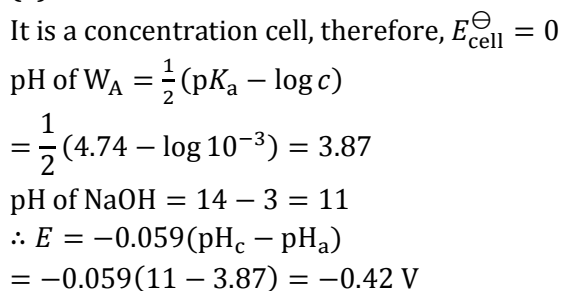
- $\Lambda_{\text{mCo}_2[\text{Fe}(\text{CN})_6]} = 2\Lambda_{\text{mCo}^{2+}} + \Lambda_{\text{m}[\text{Fe}(\text{CN})_6]^{4-}}$   
 $= 2 \times 86 + 444 = 616 \text{ S cm}^2 \text{ mol}^{-1}$   
 $\Rightarrow \Lambda_{\text{m}} = 1000 \frac{\kappa}{c} \Rightarrow c = 2.7 \times 10^{-6} \text{ M}$   
 $\Rightarrow K_{\text{spCo}_2[\text{Fe}(\text{CN})_6]} = 4c^3 = 7.87 \times 10^{-17} \text{ M}^3$
- 102 **(c)**  
Factual Statement
- 103 **(a)**  
 $\text{Pt}|\text{H}_2| \text{ Sodium acetate} + \text{CH}_2\text{COOH} (x:y)$   
 $\text{pH of acidic buffer} = \text{p}K_a + \log\left(\frac{\text{Salt}}{\text{Acid}}\right)$   
 $\therefore \text{pH}_1 = \text{p}K_a + \log\frac{x}{y}$   
Similarly,  
 $\text{pH}_2 = \text{p}K_a + \log\frac{y}{x}$   
 $\therefore E_1 = -0.059 \text{ pH}_1$   
 $E_2 = -0.059 \text{ pH}_2$   
 $(E_1 + E_2) = -0.059 \left( \text{p}K_a + \log\frac{x}{y} \right)$   
 $\quad \quad \quad - 0.059 \left( \text{p}K_a + \log\frac{y}{x} \right)$   
Solve  $\text{p}K_a = \frac{-(E_1 + E_2)}{2 \times 0.059}$
- 104 **(c)**  
 $E = E^\ominus - \frac{0.059}{2} \log\left(\frac{c_1}{c_2}\right) = \frac{0.0591}{2} \log\frac{c_2}{c_1}$   
To make  $\Delta G = -ve, E = +ve$   
Hence,  $c_2 > c_1$
- 105 **(b)**  
 $E = E^\ominus - \frac{0.059}{2} \log[\text{Cl}^\ominus]^2$   
 $= 0.268 - \frac{0.059}{2} \log(0.1)^2$   
 $= 0.268 - \frac{0.059}{2} \times 10^{-2}$   
 $= 0.268 + 0.059 = 0.327 \text{ V}$
- 108 **(d)**  
 $(E_{\text{H}^\oplus|\text{H}_2})_2 = -0.059 \text{pH}_1 = 0$   
and  $(E_{\text{H}^\oplus|\text{H}_2})_2 = -0.059 \times 7 = -0.41 \text{ V}$
- 109 **(d)**  
Factual statement
- 110 **(d)**  
Inert electrodes are used in electrolysis. Graphite, Hg, Pt
- 111 **(c)**  
 $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$   
 $2 \text{ F} \equiv 1 \text{ mol} = 63.5 \text{ g}$
- 112 **(c)**  
Thermodynamics efficiency  $= \frac{\Delta G}{\Delta H} = \frac{-nFE}{\Delta H}$
- 113 **(d)**  
 $\text{Co}^{2+} \rightarrow \text{Co}^{3+} + e^- \times 6$



114 (a)



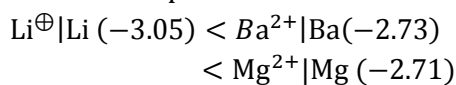
116 (a)



117 (d)

More is the reduction potential of an oxidizing agent (i.e., less -ve value) it has more tendency to undergo reduction and hence acts as a strong oxidizing agent

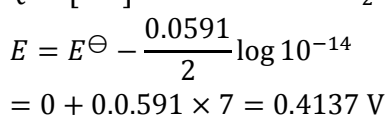
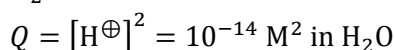
$\therefore$  Reduction potentials are as follows:



Hence,  $\text{Mg}^{2+}$  acts as the strongest oxidizing agent

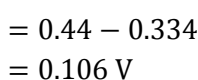
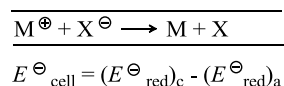
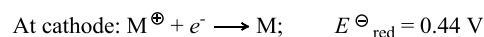
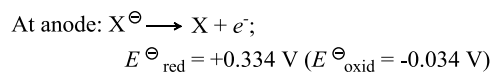
118 (c)

Since oxidation occurs



119 (b)

For reaction:



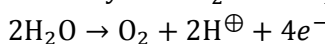
Since  $E^{\ominus}_{\text{cell}}$  is positive, hence the reaction is spontaneous

120 (b)

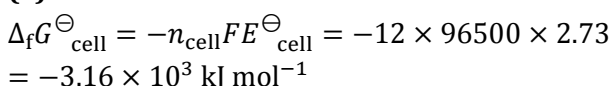
Alkaline solution will keep  $\text{H}^{\oplus}$  low so that there is a very small concentration. Hence, chances of  $\text{H}^{\oplus}$  getting reduced to  $\text{H}_2$  along with  $\text{Cu}$  are lesser in alkaline cuprous cyanide as compared to acidified copper sulphate

121 (a)

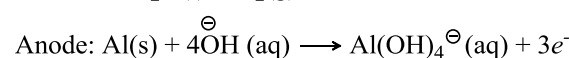
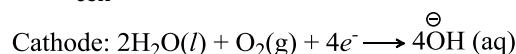
Electrolysis of  $\text{H}_2\text{O}$  takes place:



122 (a)



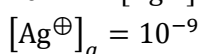
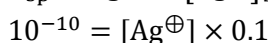
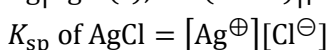
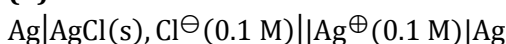
For  $n_{\text{cell}}$ , write the half cell reaction as:



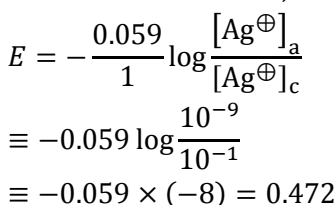
123 (b)

Since  $\text{Zn}$  acts as a sacrificial element, this implies  $\text{Zn}$  should have more tendency to get oxidized than  $\text{Fe}$

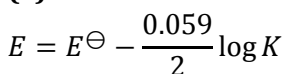
124 (d)



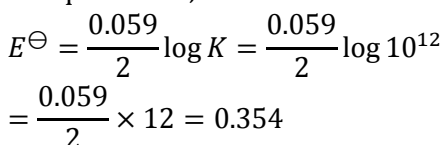
For concentration cell,  $E^{\ominus} = 0$



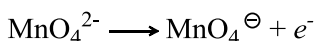
125 (a)



At equilibrium,  $E = 0$



126 (c)



$$1 \text{ F} = 1 \text{ mol of MnO}_4^{\ominus}$$

$$0.1 \text{ F} = 0.1 \text{ mol of MnO}_4^{\ominus}$$

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

128 (a)

$$E_{\text{cell}}^{\ominus} = 1.23 - (-0.44) = 1.67 \text{ V}$$

$$\Rightarrow \Delta G_{\text{cell}}^{\ominus} = nFE_{\text{cell}}^{\ominus} = -2 \times 96500 \times 1.67 \text{ J mol}^{-1}$$

$$= -332.3 \text{ kJ mol}^{-1}$$

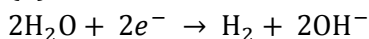
129 (b)

$$\frac{\text{Weight of H}_2}{\text{Ew of H}_2} = \frac{\text{Weight of Cu}}{\text{Ew of Cu}}$$

$$\frac{0.5}{1} = \frac{x}{63.5/2}$$

$$\therefore x = 15.9 \text{ g}$$

130 (b)



For 0.01 mole H<sub>2</sub>, 0.02 mole of electrons are consumed charge required

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

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

131 (b)

$$\kappa = 4.95 \times 10^{-5} \text{ S cm}^{-1}; \alpha = \frac{\Lambda_m}{\Lambda_m^{\circ}}$$

$$= 1000 \times \frac{\kappa}{C} = 1000 \times \frac{4.95 \times 10^{-5}}{0.001} = 49.5 \text{ S cm}^{-1}$$

$$\Rightarrow \alpha = \frac{49.5}{400} \Rightarrow \frac{1}{\alpha} = \frac{400}{49.5} \approx 8$$

132 (c)

Factual statement

133 (c)

Calomel electrode (Hg<sub>2</sub>Cl<sub>2</sub>)

134 (b)

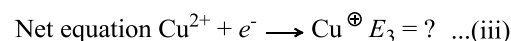
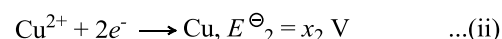
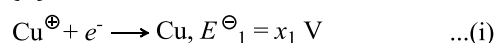
The smaller the reduction potential of a substance, the more is its reducing power and the less is its oxidizing power

The given order of reduction potential is Z > Y > X

The order of oxidizing power is Z > Y > X

Therefore, Y will oxidize X and not Z

135 (d)



Is obtained by equations (ii) - (i)

$$\therefore E_3 = \frac{n_2 E_2 - n_1 E_1}{n_3}$$

$$= \frac{2 \times x_1 - 1 \times x_1}{1} = 2x_2 - x_1$$

136 (c)

$$1 \text{ F} = 1 \text{ Eq of H}_2 = 11.2 \text{ L}$$

$$2 \text{ F} = 22.4 \text{ L}$$

$$1 \text{ F} = 1 \text{ Eq of O}_2 = 5.6 \text{ L}$$

$$2 \text{ F} = 11.2 \text{ L}$$

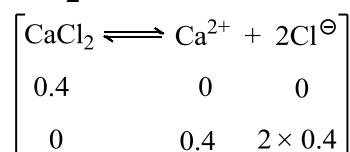
137 (a)

$$\text{Ionic strength} = \mu = \frac{1}{2} \sum c_i Z_i^2$$

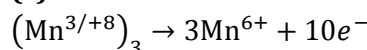
c<sub>i</sub> = concentration of ion; Z<sub>i</sub> = valency of ion

$$[\text{Ca}^{2+}] = 0.4 \text{ M}, [\text{Cl}^{\ominus}] = 0.8 \text{ M}$$

$$\mu = \frac{1}{2} [0.4(2)^2 + 0.8(1)^2] = 1.2$$



138 (c)



10 Faraday charge is required for conversion of 1 mol of Mn<sub>3</sub>O<sub>4</sub> to MnO<sub>4</sub><sup>2-</sup>

139 (c)

$$1 \text{ F} = 1 \text{ Eq of H}_2 = 11.2 \text{ L of H}_2$$

$$11.2 \text{ L of H}_2 = 1 \text{ Eq of H}_2$$

$$2.24 \text{ L of H}_2 = \frac{2.24}{11.2} = 0.2 \text{ Eq}$$

∴ 0.2 Eq of Cu will be deposited

$$1 \text{ Eq of Cu} = \frac{63.5}{2} \text{ g}$$

$$0.2 \text{ Eq of Cu} = \frac{63.5}{2} \times 0.2 = 6.35 \text{ g}$$

140 (c)

$$E = E^{\ominus} + \frac{0.0591}{1} \log \frac{[\text{Ag}^{\oplus}]_R}{[\text{Ag}^{\oplus}]_L}$$

$$= 0 + 0.0591 \log 2 = 0.0591 \times 0.301 \text{ V}$$

After current is passed [Ag<sup>⊕</sup>]<sub>R</sub> = 1 M

$$[\text{Ag}^{\oplus}]_L = 0.5 \text{ M}$$

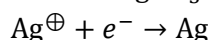
Hence, no change in EMF

141 (c)

At anode oxidation of H<sub>2</sub>O occurs since the oxidation potential of H<sub>2</sub>O is greater than the oxidation potential of SO<sub>4</sub><sup>2-</sup> ion

142 (d)

$$\text{mmoles AgNO}_3 = 0.1 \times 200 = 20$$



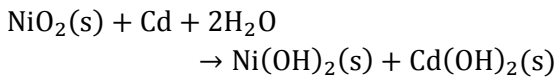
For 50% electrolysis, 10 mF is required

$$\Rightarrow 10 \text{ mF} = \frac{It}{96500} \Rightarrow t = 9650 \text{ s}$$

143 (d)

The net redox change:





144 (b)

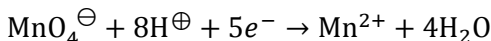
$$\frac{W}{E_w} = \frac{965 \times 1}{96500} = 0.01$$

Equivalent of  $\text{NiSO}_4$  present initially =  $1 \times 0.02 = 0.02$

If Ni electrodes are used no change in concentration of  $\text{NiSO}_4$ , i.e., 0.02 N or 0.01 M

If Pt electrodes are used, then equivalent of  $\text{NiSO}_4$  left = 0.01

145 (a)



$$E_{\text{MnO}_4^-|\text{Mn}^{2+}}$$

$$= E_{\text{MnO}_4^-|\text{Mn}^{2+}}^\ominus - \frac{0.059}{5} \log \left( \frac{[\text{Mn}^{2+}]}{[\text{MnO}_4^-][\text{H}^+]^8} \right)$$

$$= E_{\text{MnO}_4^-|\text{Mn}^{2+}}^\ominus - \frac{8}{5} \times 0.059 \text{ pH}$$

$$- \frac{0.059}{5} \log \left( \frac{[\text{Mn}^{2+}]}{[\text{MnO}_4^-]} \right)$$

$\Rightarrow [\text{H}^+]$  is doubled, i.e., pH is reduced by  $\log 2 \equiv 0.3$ , then  $E_{\text{MnO}_4^-|\text{Mn}^{2+}}$  will be changed

(increase) by

$$\frac{8}{5} \times 0.059 \times 0.3 \text{ V} = 28.36 \text{ mV}$$

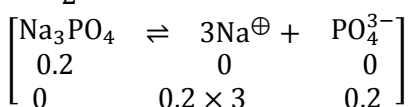
146 (b)

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

147 (b)

$$[\text{Na}^+] = 0.6 \text{ M} \quad [\text{PO}_4^{3-}] = 0.2$$

$$\mu = \frac{1}{2} [0.6(1)^2 + 0.2(3)^2] = 1.2$$



148 (c)

Aqueous KCl on electrolysis gives  $\text{Cl}_2$  at anode

and  $\text{H}_2$  at cathode and  $\text{OH}^-$  ions in the solution

$$\text{Equivalent of current} = \frac{1 \times 96.5}{96500} = 0.01 \text{ M}$$

$$\therefore [\text{OH}^-] = \frac{1 \text{ mEq}}{100 \text{ mL}} = 0.01 \text{ M}$$

149 (c)

When  $\Delta G = +ve$  (it called endergonic) or

$E_{\text{cell}} = -ve$  or non spontaneous reaction

This is only possible, when  $p_2 > p_1$  or  $c_1 > c_2$

150 (a)

Higher the oxidation potential, more easily it is oxidized and faster is the deposition of metal.

Decreasing order of  $E^\ominus$  oxidation is

$$E_{\text{Ag}|\text{Ag}^+(0.80)}^\ominus > E_{2\text{Hg}|\text{Hg}_2^{2+}(0.79)}^\ominus$$

$$> E_{\text{Cu}|\text{Cu}^{2+}(0.34)}^\ominus$$

$$> E_{\text{Mg}|\text{Mg}^{2+}(-2.37)}^\ominus$$

Hence, the sequence of deposition of metals at the cathode will be Ag, Hg, and Cu

152 (a,b)



$$E_{\text{cell}} = E_{\text{cell}}^\ominus - \frac{0.059}{2} \log \frac{[\text{Co}^{2+}]}{[\text{H}^+]^2} \quad (p_{\text{H}_2} \equiv Q_{\text{cell}})$$

a. Increase in the volume (i.e., dilution) of  $\text{CoCl}_2$  solution (i.e., decrease in  $[\text{Co}^{2+}]$ ) will cause  $Q_{\text{cell}}$  to decrease causing  $E_{\text{cell}}$  to increase

b. Increasing  $M_2$  will cause a reduction in  $Q$  thus increasing  $E_{\text{cell}}$

c. Increasing  $p_{\text{H}_2}$  will increase  $Q$  thus reducing  $E_{\text{cell}}$

d. Increasing  $M_1$  will increase  $Q$  thus reducing  $E_{\text{cell}}$

153 (a,d)

a. At cathode: Reduction of  $\text{Na}^+$  does not occur

but reduction of  $\text{H}_2\text{O}$  occurs to give  $\text{OH}^-$  and  $\text{H}_2(\text{g})$ , so pOH decreases and pH increases

At anode: Oxidation of  $\text{Cl}^-$  ions occurs to give  $\text{Cl}_2$

(g). Likewise, oxidation of  $\text{NO}_3^-$  and  $\text{SO}_4^{2-}$  does not occur but oxidation of  $\text{H}_2\text{O}$  occurs to give  $\text{H}^+$  ions and  $\text{H}_2(\text{g})$ . So pH decreases at anode

155 (c)

For strong and weak electrolytes, increasing concentration will increase conductance ( $G$ ), decrease molar conductance ( $\Lambda_m$ ) and equivalent conductance

( $\Lambda_{\text{eq}}$ ) and increases conductivity ( $\kappa$ )

158 (b,c)

In the given case,

$$E_{\text{oxi}} = - \frac{0.059}{1} \log \frac{[\text{H}^+]}{\sqrt{p_{\text{H}_2}}}$$

$\therefore$  Increase  $[\text{H}^+]$  or decrease  $p_{\text{H}_2}$ , to decrease  $E_{\text{oxi}}$

159 (a,c,d)

Correct statements:

a. Rust is  $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$

b. Correct

c. Saline water fastens the rusting

d. Pure metal undergo corrosion slower than impure metals

160 (a,b,c)

As temperature increases solubility decreases

$\therefore \kappa \downarrow; \Lambda_m \uparrow$

Specific conductance ( $\kappa \uparrow$ ) increases as concentration increases

167 (a,d)

$\Delta G = \Delta H - T \Delta S$  and  $\Delta G = \Delta H + T \left(\frac{\partial \Delta G}{\partial T}\right)_P$

$$\therefore \left(\frac{\partial \Delta G}{\partial T}\right)_P = \frac{\Delta G - \Delta H}{T} = -\frac{T \Delta S}{T} = -\Delta S$$

$$\therefore \Delta S = +nF \left(\frac{\partial E}{\partial T}\right)_P$$

Also,  $-nFE = \Delta H + T \times (-nF) \left(\frac{\partial E}{\partial T}\right)_P$

$$\therefore \left(\frac{\partial E}{\partial T}\right)_P = \frac{\Delta H + nFE}{T}$$

172 (a,b,d)

If  $E^\ominus_{\text{cell}} > 0$ , then the reaction is spontaneous

a.  $E^\ominus_{\text{cell}} = 0.8 - 0.34 = 0.46 \text{ V}$  ( $2\text{Ag}^\oplus + \text{Cu} \rightarrow 2\text{Ag} + \text{Cu}^{2+}$ )

b.  $E^\ominus_{\text{cell}} = 0.34 - (-2.37) = 2.81 \text{ V}$  ( $\text{Cu}^{2+} + \text{Mg} \rightarrow \text{Cu} + \text{Mg}^{2+}$ )

c.  $E^\ominus_{\text{cell}} = 0.34 - 0.8 = -0.46 \text{ V}$  ( $\text{Cu}^{2+} + 2\text{Ag} \rightarrow \text{Cu} + 2\text{Ag}^{2+}$ )

d.  $E^\ominus_{\text{cell}} = 0.79 - 0.34 = 0.45 \text{ V}$  ( $\text{Hg}^{2+} + \text{Cu} \rightarrow \text{Cu}^{2+} + \text{Hg}$ )

In (a), (b), and (d),  $E^\ominus_{\text{cell}} > 0$  which means that the respective solution in the respective vessel cannot be stored

174 (a,b,c,d)

$$E_{\text{cell}} = E^\ominus_{\text{cell}} - \frac{0.059}{2} \log \frac{[\text{H}^\oplus]^2}{[\text{Zn}^{2+}]p_{\text{H}_2}}$$

$$\text{Check for } \frac{[\text{H}^\oplus]^2}{[\text{Zn}^{2+}]p_{\text{H}_2}} = 1 \text{ for } E_{\text{cell}} = E^\ominus_{\text{cell}}$$

It is possible in all cases

175 (a,b,c)

A reducing agent will react spontaneously with any oxidizing agent stronger than the oxidizing agent with which it is coupled. Pb is placed above  $\text{Cu}^{2+}$ ,  $\text{Br}_2$  and  $\text{Ag}^+$  in the electrochemical series. Hence, it can reduce all of these

179 (b,c)

"A" depends on the type (magnitude of charge on cation and anion) of electrolyte for a given

temperature of solvent

**Electrolyte**      **Type**

NaCl                      1, -1

CaCl<sub>2</sub>                    [2, -1] same type,

ZnCl<sub>2</sub>                    [2, -1] hence same "A"

MgSO<sub>4</sub>                    2, -2

181 (a,b)

$$\text{Number of Faradays} \equiv \frac{It}{96500} = \frac{2.68 \times 1 \times 3600}{96500} = 0.1$$

At cathode:  $\text{Cu}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Cu}(\text{s})$

$\Rightarrow 2F \equiv 1 \text{ mol Cu}$

$\Rightarrow 0.1 F \equiv 0.05 \text{ mol Cu} \equiv 0.05 \times 63.5 \text{ g Cu}$

$= 3.175 \text{ g Cu deposited}$

At anode:  $\text{Cu}(\text{s}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2e^-$

$\Rightarrow 2 F \equiv 1 \text{ mole Cu}$

$\Rightarrow 0.1 F \equiv 0.05 \text{ mol Cu} \equiv 0.065 \times 63.5 \text{ g Cu}$

$= 3.175 \text{ g Cu used}$

182 (a,b,c,d)

Reduction potential of  $\text{F}_2$  is highest. Oxidation potential of Li is highest

184 (a,b,c)

Salt bridge is used to eliminate liquid junction potential arising due to different speed of ions present in cathodic and anodic compartments

187 (a,c,d)

Electrolysis of concentrated  $\text{H}_2\text{SO}_4$  at  $0 - 5^\circ\text{C}$  using Pt electrode produces  $\text{H}_2\text{S}_2\text{O}_8$

At anode:  $2\text{H}_2\text{SO}_4 \rightarrow \text{H}_2\text{S}_2\text{O}_8 + 2\text{H}^\oplus + 2e^-$

Statements (c) and (d) are factual statements

Statement (b) is wrong. Electrolysis of brine (aq. NaCl) gives  $\text{H}_2(\text{g})$  at cathode and  $\text{Cl}_2(\text{g})$

190 (c,d)

Create a cell with required cell reaction:

$$\text{O}_2 + \text{SO}_4^{2-} \rightarrow \text{S}_2\text{O}_8^{2-} + \text{H}_2\text{O}; E^\ominus_{\text{cell}} = 1.23 - 2.01 < 0$$

$\Rightarrow$  Spontaneous cell reaction: (reverse reaction),

i.e.,  $\text{H}_2\text{O} + \text{S}_2\text{O}_8^{2-} \rightarrow \text{O}_2 + \text{SO}_4^{2-}$

192 (a,c,d)

During the working of a galvanic cell, it reaches equilibrium from its standard conditions

$\Rightarrow Q \downarrow$  to reach  $K_{\text{eq}}; E_{\text{cell}} \downarrow$

Reaction becomes less spontaneous as  $E_{\text{cell}} \downarrow$

194 (a,b)

CaI <sub>2</sub> (aq): Cathode:	Ca <sup>2+</sup> ,	H <sup>⊕</sup>
	×	✓
Anode:	I <sup>⊖</sup> ,	OH <sup>⊖</sup>
	✓	×

$$\text{a. } \left[ \begin{array}{l} \text{Mmoles I}_2 = 50 \Rightarrow 100\text{mF passed} = 0.1 F \\ (2I^\ominus \rightarrow \text{I}_2 + 2e^-) \end{array} \right]$$

b. Also, mmoles of  $H_2 = \frac{1}{2} \times$  number of mF

$$= \frac{1}{2} \times 100 = 50 \Rightarrow V_{H_2} \text{ at STP}$$

$$= 50 \times 22.4 \text{ ml} \equiv 1120 \text{ mL}$$

100 mmoles  $OH^\ominus$  left

$$\Rightarrow [OH^\ominus] = \frac{100}{500} = 0.2$$

c.

$$\Rightarrow pOH = -\log 0.2 = 0.7$$

$$\Rightarrow pH = 13.3$$

195 (a,c)

**Cathodic protection:** A technique to control corrosion of a metal surface by making it work as a cathode of an electrochemical cell by placing in contact with the metal to be protected another more easily corroded metal to act as the anode of the electrochemical cell. Most commonly used to protect steel, water pipelines, and storage tanks

**Anodic protection:** A technique to control corrosion of a metal by making it work as anode developing a passive film on the metal

It used in extremely corrosive conditions and most extensively used to store and handle sulphuric acid container

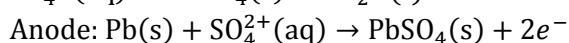
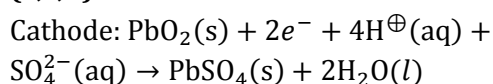
196 (c,d)

$E_{\text{cell}} = E^\ominus_{\text{cell}} - \frac{RT}{nF} \ln \frac{[Zn^{2+}]}{[Cu^{2+}]}$ . Decrease in  $[Zn^{2+}]$  or increase in  $[Cu^{2+}]$  will increase  $E_{\text{cell}}$

197 (c,d)

Electrolysis of  $ZnSO_4$  is done for the purification of Zn using Al cathode on which it gets reduced (factual statement)

198 (a,c,d)



199 (c,d)

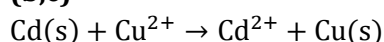
In the electrolysis of very dilute NaOH solution,  $H_2$  is evolved at the cathode while  $O_2$  is produced at anode

200 (a,b)

$$K_1 = \frac{[C][D]}{[A][B]} \text{ and } K_2 = \frac{[C]^2[D]^2}{[A][B]}$$

Also,  $E^\ominus$  is independent of stoichiometry

202 (b,c)



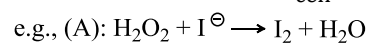
(1.0 M) (1.0 M)

$$E_{\text{cell}} = E^\ominus_{\text{cell}} - \frac{0.059}{2} \log \frac{[Cd^{2+}]}{[Cu^{2+}]}$$

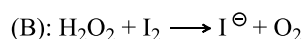
$E_{\text{cell}}$  increases when condition as in (b) and (c) are made

204 (a,d)

Create a cell with required equation (cell reaction) and find its  $E_{\text{cell}}^\ominus$



$$E_{\text{cell}}^\ominus = (1.77) - (0.54) > 0$$

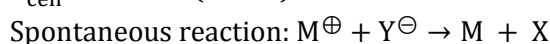


$$E_{\text{cell}}^\ominus = (0.54) - (0.69) < 0$$

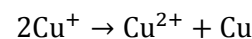
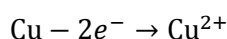
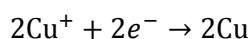
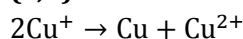
(Not possible to occur spontaneously) and so on

205 (b,c)

$$E_{\text{cell}}^\ominus = 0.44 - (-0.33) = 0.77 \text{ V}$$

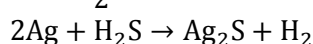
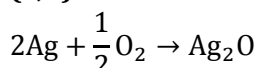


206 (a,b)



$$\therefore E^\ominus = \frac{2 \times 0.521 + 2(-0.337)}{2} = 0.184$$

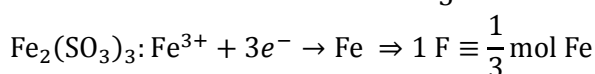
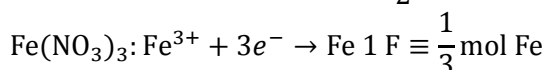
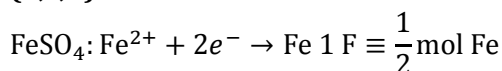
207 (a,b)



209 (a,b,c,d)

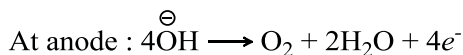
The most common example of corrosion is the rusting of iron. Rust is hydrated ferric oxide,  $Fe_2O_3 \cdot xH_2O$ . Some other examples include tarnishing of silver, development of green coating on copper, bronze, etc.

210 (b,c,d)



Amount of Fe deposited in  $Fe(NO_3)_3$  = Amount of Fe deposited in  $Fe_2(SO_4)_3$

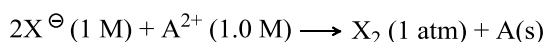
$$\text{Amount of Fe deposited in } Fe_2(SO_4)_3 = \frac{1}{2} : \frac{1}{3} = 1.5 : 1$$



In all cases same amount of gas is evolved at the anode

211 (a,b,d)

$E_{\text{cell}}^{\ominus}$  corresponding to the reaction is:



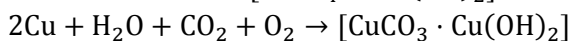
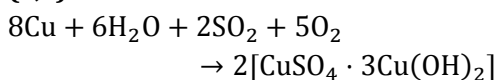
$$E_{\text{cell}}^{\ominus} = E_{\text{A}^{2+}|\text{A}}^{\ominus} - E_{\text{X}_2|2\text{X}^{\ominus}}^{\ominus} = 0.08 - 1.30 = -1.22 \text{ V}$$

$$E_{\text{cell}}^{\ominus} = E_{\text{A}^{2+}|\text{A}}^{\ominus} - E_{\text{X}_2|2\text{X}^{\ominus}}^{\ominus} = 0.08 - 1.30 = -1.22 \text{ V}$$

$E_{\text{cell}}$  is negative, so  $\Delta G^{\ominus}$  will be (+ve) and hence react (c) will not be feasible.

So, statement (c) is wrong. Statement (a), (b), and (d) are correct

212 (a,c)



213 (a,c)

For concentration cell,  $E_{\text{cell}}^{\ominus} = 0$

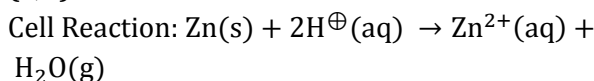
$$\text{a. } E_{\text{cell}} = 0 - \frac{0.059}{2} \log \frac{0.029}{0.48} = 0.038 \text{ V}$$

$$\text{c. } |\text{Zn}^{2+}| \text{ new in LHE} = 0.012$$

$$\therefore E'_{\text{cell}} = 0 - \frac{0.059}{2} \log \frac{0.012}{0.48} = 0.0472$$

$$\therefore E'_{\text{cell}} > E_{\text{cell}}$$

214 (a,b)



$$\text{Using: } E_{\text{cell}} = E_{\text{cell}}^{\ominus} - \frac{0.0591}{2} \log \left( Q_{\text{cell}} - \frac{p_{\text{H}_2} \cdot [\text{Zn}^{2+}]}{[\text{H}^{\oplus}]^2} \right)$$

$$\Rightarrow \text{For } E_{\text{cell}} = E_{\text{cell}}^{\ominus}, Q_{\text{cell}} = 1.0$$

$$\text{a. } Q_{\text{cell}} = \frac{1 \times 1}{1^2} = 1$$

$$\text{b. } Q_{\text{cell}} = \frac{1 \times 0.01}{(0.1)^2} = 1$$

$$\text{c. } Q_{\text{cell}} = \frac{1 \times 1}{(0.1)^2} = 100$$

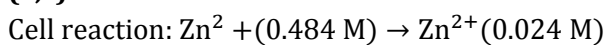
215 (a,c)

$E_Z^{\ominus}$  is highest  $\Rightarrow$  Strongest oxidizing agent

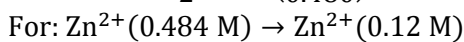
$E_X^{\ominus}$  is lowest  $\Rightarrow$  Weakest oxidizing agent

$\Rightarrow$  Y will oxidize X but not Z. Also, Z will oxidize both X and Y

216 (a,c)



$$E_{\text{cell}} = 0 - \frac{0.059}{2} \log \left( \frac{0.024}{0.480} \right) = 0.039 \text{ V}$$



$$\Rightarrow E_{\text{cell}} = 0 - \frac{0.059}{2} \log \left( \frac{0.012}{0.48} \right)$$

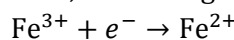
$$= 0.03(1.6) = 0.048 \text{ V}$$

217 (a,b,c)

$$\text{Number of Faradays} = \frac{4 \times 1 \times 3600}{96500} = 0.15$$

Initially, moles of  $\text{Fe}^{3+} = 0.1 \times 1 = 0.1$

First,  $\text{Fe}^{3+}$  will get reduced to  $\text{Fe}^{2+}$



1 F  $\equiv$  1 mol  $\text{Fe}^{3+}$  deposited

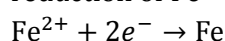
$\Rightarrow 0.15 \text{ F} \equiv 0.15 \text{ mol } \text{Fe}^{3+}$  deposited  $> \text{Fe}^{3+}$  available

Thus, 1 mol  $\text{Fe}^{3+} \equiv 1 \text{ F}$

$\Rightarrow 0.1 \text{ mol } \text{Fe}^{3+} \equiv 0.1 \text{ F}$  electricity is used

$\equiv 0.1 \text{ mol } \text{Fe}^{2+}$  produced

$\Rightarrow 0.15 - 0.1 = 0.05 \text{ F}$  electricity left for the reduction of  $\text{Fe}^{2+}$



2 F  $\equiv$  1 mol  $\text{Fe}^{2+}$

$\Rightarrow 0.05 \text{ F} \equiv \frac{0.05}{2} = 0.025 \text{ mol } \text{Fe}^{2+}$  reduced  $\equiv 0.025$

mol Fe deposited

$\Rightarrow \text{Fe}^{2+}$  left =  $0.1 - 0.025 = 0.075 \text{ mol}$

219 (b,c)

$$E_{\text{Red}} = -0.059 \left[ \text{pH} + \frac{\log p_{\text{H}_2}}{2} \right]$$

$$\Rightarrow E_{\text{cell}} = 0.059 \left[ \text{pH}_a - \text{pH}_c + \frac{1}{2} \log \frac{(p_{\text{H}_2})_a}{(p_{\text{H}_2})_c} \right] \text{ V}$$

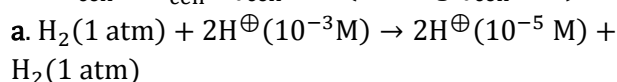
$E_{\text{cell}} > 0$ , when  $p_1 = p_2$  and  $x_1 < x_2$

Or when  $x_1 = x_2$  and  $p_1 > p_2$

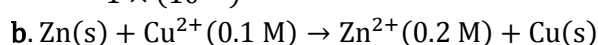
222 (a)

$$\text{Use: } E_{\text{cell}} = E_{\text{cell}}^{\ominus} - \frac{0.059}{n_{\text{cell}}} \log Q_{\text{cell}}$$

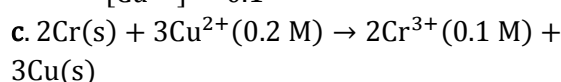
For  $E_{\text{cell}} > E_{\text{cell}}^{\ominus}$ ,  $Q_{\text{cell}} < 1$  (i.e.,  $\log Q_{\text{cell}} < 0$ )



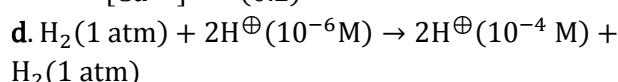
$$Q_{\text{cell}} = \frac{(10^{-5})^2 \times 1}{1 \times (10^{-3})^2} = 10^{-4} < 1$$



$$Q_{\text{cell}} = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} = \frac{0.2}{0.1} = 2 > 1$$



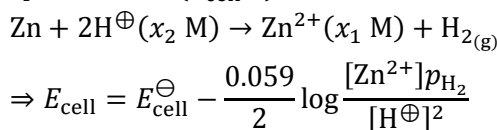
$$Q_{\text{cell}} = \frac{[\text{Cr}^{3+}]^2}{[\text{Cu}^{2+}]^3} = \frac{(0.1)^2}{(0.2)^3} > 1$$



$$Q_{\text{cell}} = \frac{(10^{-4})^2 \times 1}{1 \times (10^{-6})^2} = 10^4 > 1$$

223 (c,d)

Zn|Zn<sup>2+</sup>(x<sub>1</sub> M)||H<sup>+</sup>(x<sub>2</sub> M)|H<sub>2</sub>(g), Pt  
Free energy to decrease ⇒ cell is more spontaneous (E<sub>cell</sub> ↑)



ΔG = -ve or E<sub>cell</sub> = +ve

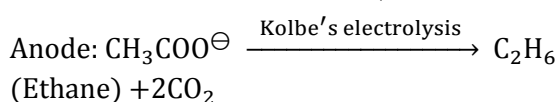
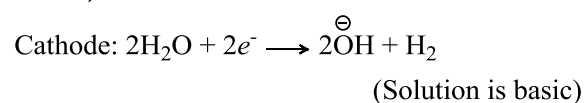
c. When x<sub>2</sub>(0.1 M) changed to 1 M

d. When x<sub>1</sub> (1 M) changed to 0.1 M

226 (a,c)

a. Reduction potential of H<sub>2</sub>O > Reduction potential of Na<sup>+</sup>

Hence,



b. K<sup>+</sup> does not undergo reduction but reduction of H<sub>2</sub>O occurs to give <sup>⊖</sup>O H ion and H<sub>2</sub>(g)

similarly, NO<sup>⊖</sup><sub>3</sub> ion does not undergo oxidation but oxidation of H<sub>2</sub>O occurs to give H<sup>+</sup> ions and O<sub>2</sub>(g)

H<sup>+</sup> and <sup>⊖</sup>O H ions get neutralized and pH = 7 (neutral solution)

c. Na<sup>+</sup> ions do not undergo reduction but

reduction of H<sub>2</sub>O occurs to give <sup>⊖</sup>O H ions and H<sub>2</sub>(g). (Hence, solution is basic)

Cl<sup>⊖</sup> undergo oxidation to give Cl<sub>2</sub>(g)

d. Same explanation as in (b)

229 (c,d)

It is not clear, the values given are E<sup>∘</sup><sub>RP</sub> or E<sup>∘</sup><sub>OP</sub>

Assuming E<sup>∘</sup><sub>OP</sub> order Y > X > Z

∴ Reducing power Y > X > Z

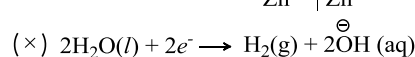
Assuming E<sup>∘</sup><sub>RP</sub> order Z > X > Y

∴ Reducing power Z > X > Y

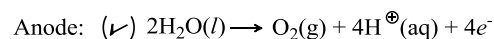
230 (a,d)

Aqueous Zn(NO<sub>3</sub>)<sub>2</sub>;

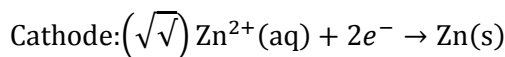
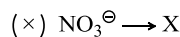
$$E_{\text{Zn}^{2+} | \text{Zn}}^{\ominus} = -0.76 \text{ V}$$



$$E_{\text{H}_2\text{O} | \text{H}_2}^{\ominus} = -0.83 \text{ V}$$

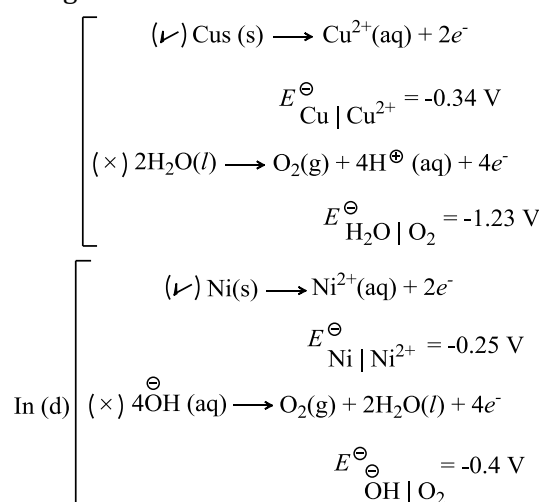


$$E_{\text{H}_2\text{O} | \text{O}_2}^{\ominus} = -1.23 \text{ V}$$



233 (a,b)

Clearly, O<sub>2</sub> will be evolved in (a) and (b) In (c), Cu will get oxidized



234 (a,c)

$$\frac{W}{E_w} = \frac{It}{96500} = \frac{9.65 \times 3600}{96500}$$

$$= 0.36 \text{ eq of Ag}^{\oplus}$$

$$= 0.36 \text{ eq of Cu}^{2+}$$

$$= 0.36 \text{ moles of Ag}^{\oplus}$$

$$= 0.18 \text{ mole of Cu}^{2+}$$

Now, Ag will oxidize to Ag<sup>+</sup> and Cu<sup>2+</sup> will reduce to Cu

235 (d)

**Note:** Cd(s)|CdCl<sub>2</sub>(0.1 M)|AgCl(s)|Ag(s) is *equiva*  
Cd(s)|Cd<sup>2+</sup>(0.1 M)||Cl<sup>⊖</sup>(0.2 M)|AgCl(s)|Ag(s)|

$$\text{Use: } \Delta H = nFG \left[ T \cdot \left( \frac{dE_{\text{cell}}}{dT} \right) - E_{\text{cell}} \right]$$

$$\Delta H_{25^\circ\text{C}} = 2 \times 96500$$

$$\left[ 298 \times \left( \frac{0.6753 - 0.6915}{25 - 0} \right) - 0.6753 \right]$$

$$= -167 - 67 \text{ kJ mol}^{-1}$$

236 (a,c,d)

Although every spontaneous redox reaction can be used as the basis of an electrochemical cell, yet every such electrochemical cell is not suitable for commercial purposes. This is because of the following two reasons:

a. Electrochemical cells using salt bridges have

high internal resistance. As a result, if we try to draw large current from them, their voltage drops sharply

b. They lack the compactness and ruggedness (rough handling) for probability

An electrochemical cell to be used as a commercial cell must fulfil the following two requirements:

a. It should have the compactness, lightness and ruggedness for probability

b. Its voltage should not drop much during use, i.e., drop in voltage should be negligible over the small interval of time during which it is being put to use

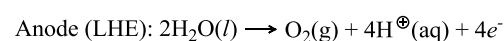
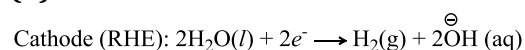
If a number of cells are connected in series, the arrangement is called a battery. However, the term battery is sometimes used even for a single cell. Strictly speaking, such a usage is not correct. Various commercial cells may be classified mainly into the following three types: (a) primary cells, (b) secondary cells, and (c) fuel cells

*Primary cells* are those in which the redox reaction occurs only once and the cell becomes dead after some time, and cannot be used again. Two common examples of this type are dry cell and mercury cell

*Secondary cells* are those which can be recharged by passing an electric current through them and hence can be used over and again. Two well-known examples of this type are lead storage battery and nickel-cadmium storage cell

*Fuel cells* are those in which the energy produced from the combustion of fuels such as  $H_2$ , CO,  $CH_4$ , etc., is directly converted into electrical energy. Such a conversion is possible because combustion reaction are also redox reactions

239 (b)



In RHE,  $OH^{\ominus}$  will be produced which will react with  $NH_4^{\oplus}$  (salt) to form  $NH_4OH$  (base), causing pOH to decrease or pH to increase

In LHE,  $H^{\oplus}$  will be produced with will react with  $NH_4OH$  (base) to form  $NH_4^{\oplus}$  (salt), causing pOH to increase or pH to decrease

242 (a,c,d)

Factual statement

244 (a,b,c)

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

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

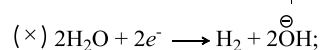
$$\Rightarrow \left( \frac{\partial E}{\partial T} \right)_p = \frac{\Delta H + nFE}{nF}$$

$$\text{Check for: } \left( \frac{\partial E}{\partial T} \right) = 0; > 0; < 0$$

245 (a,b)

Aqueous  $CuBr_2$ :

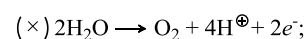
$$E_{Cu^{2+} | Cu}^{\ominus} = +0.34 \text{ V}$$



$$E_{H_2O | H_2}^{\ominus} = -0.83 \text{ V}$$



$$E_{Br^{\ominus} | Br_2}^{\ominus} = -1.09 \text{ V}$$



$$E_{H_2O | H_2}^{\ominus} = -1.23 \text{ V}$$

