Electrochemistry

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- 1. Which of the following statement is correct?
 - Cathode is -ve terminal in both, glavanic (a)and electrolytic cells.
 - Anode is +ve terminal in both, galvanic and (b) electrolytic cells.
 - Cathode and anode are -ve terminal in (c) electrolytic and galvanic cell.
 - (d) Cathode and anode are +ve terminal in electrolytic and galvanic cell.
- 2. Based on the cell notation for a spontaneous reaction, at the anode :

 $Ag(s) |AgCl(s)| Cl^{-}(aq) || Br^{-}(aq) || Br_{2}(l) |C(s)|$

- (a) AgCl gets reduced
- (b) Ag gets oxidized
- (c) Br⁻ gets oxidized
- (d) Br₂ gets reduced
- 3. For the electrochemical cell

Pt (s) | H₂(g) | H⁺(1M) |, Cu (s) which one of 1 atm

the following statements is true?

- (a) H^+ ions are formed at anode and Cu is deposited at cathode.
- (b) H₂ is liberated at cathode and Cu is deposited at anode.
- (c) Oxidation occurs at cathode.
- (d) Reduction occurs at anode.
- 4. The standard reduction potentials at 298K for the following half reactions are given against each

$$\operatorname{Zn}^{2+}(aq) \quad 2e^{-} \longrightarrow \operatorname{Zn}(s); \quad -0.762 \text{ V}$$

 $\operatorname{Cr}^{3+}(aq) \quad 3e^{-} \longrightarrow \operatorname{Cr}(s); \quad -0.740 \text{ V}$

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

$$\operatorname{Fe}^{3+}(aq) e^{-} = \operatorname{Fe}^{2}(aq); 0.770 \mathrm{V}$$

Which is the strongest reducing agent?

Tricky

(a) Zn(s)(b) Cr(s)

(d) $Fe^{3+}(aq)$ (c) $H_{2}(g)$

5. Zn gives H_2 gas with H_2SO_4 and HCl but not with HNO₃ because Tricky

- (a) Zn acts as oxidizing when reacts with HNO_3 .
- (b) HNO_3 is weaker acid than H_2SO_4 and HCl.
- (c) In electrochemical series Zn is above hydrogen.
- (d) NO_3^{-1} is reduced in preference to hydronium.
- A smuggler could not carry gold by depositing iron on the gold surface since
- (a) gold is denser
- (b) iron rusts
- gold has higher reduction potential than iron (c)
- (d) gold has lower reduction potential than iron
- 7. Standard electrode potential for Sn⁴⁺ / Sn²⁺ couple is + 0.15 V and that for the Cr^{3+} / Cr couple is -0.74V. These two couples in their standard state are connected to make a cell. The cell potential will be:

(a)
$$+1.19$$
 V (b) $+0.89$ V

- (d) +1.83 V (c) +0.18 V
- A solution contains Fe²⁺, Fe³⁺ and I⁻ ions. This solution was treated with iodine at 35°C. E° for $\text{Fe}^{3+}/\text{Fe}^{2+}$ is + 0.77 V and E° for $I_2/2I^- = 0.536$ V. The favourable redox reaction is :
 - (a) I₂ will be reduced to I^-
 - (b) There will be no redox reaction
 - (c) I^- will be oxidised to I_2
 - (d) Fe^{2+} will be oxidised to Fe^{3+}

 $P = Cu^{2+} (0.0001 \text{ M})/Cu(s)$ $Q = Cu^{2+} (0.1 \text{ M})/Cu(s)$ $R = Cu^{2+} (0.01 \text{ M})/Cu(s)$ $S = Cu^{2+} (0.001 \text{ M})/Cu(s)$ If the standard reduction potential of Cu²⁺/Cu

- is +0.34 V, the reduction potentials in volts of the above electrodes follow the order.
- (a) P > S > R > Q(b) S > R > Q > P(c) R > S > Q > P(d) Q > R > S > P
- Critical Thinking

- **10.** Following cell has EMF 0.7995V.
 - Pt $|H_2(1 \text{ atm})|HNO_3(1M)||AgNO_3(1M)|Ag$ If we add enough KCl to the Ag cell so that the final Cl⁻ is 1M. Now the measured emf of the cell is 0.222V. The K_{sp} of AgCl would be –

Toughnut

- (a) $1 \times 10^{-9.8}$ (b) $1 \times 10^{-19.6}$
- (c) 2×10^{-10} (d) 2.64×10^{-14} 11. Standard cell voltage for the cell Pb | Pb²⁺ || Sn²⁺ | Sn is - 0.01 V. If the cell is to
 - exhibit $E_{cell} = 0$, the value of $[Sn^{2+}]/[Pb^{2+}]$ should be antilog of –

(a)
$$+0.3$$
 (b) 0.5 (c) 1.5 (d) -0.5

- 12. An electrochemical cell is set up as: Pt; H_2 (1atm)|HCl (0.1 M) || CH₃COOH (0.1 M)|H₂ (1atm); Pt. The e.m.f of this cell will not be zero, because Tricky
 - (a) the temperature is constant
 - (b) e.m.f depends on molarities of acids used
 - (c) acids used in two compartments are different
 - (d) pH of 0.1 M HCl and 0.1 M CH₃COOH is not same
- 13. A variable, opposite external potential (E_{ext}) is applied to the cell

$$\label{eq:2n} \begin{split} &Zn|Zn^{2+}(1\ M)\,\|\,Cu^{2+}(1\ M)\,\|\,Cu, of\ potential\ 1.1\ V.\\ &When\ E_{ext}\,{<}\,1.1\ V \ and\ E_{ext}\,{>}\,1.1\ V, respectively\\ electrons\ flow\ from: \end{split}$$

- (a) anode to cathode in both cases
- (b) cathode to anode and anode to cathode
- (c) anode to cathode and cathode to anode
- (d) cathode to anode in both cases

14. The standard electrode potentials
$$E_{M/M}^{o}$$
 of

four metals A, B, C and D are -1.2 V, 0.6 V, 0.85 V and -0.76 V, respectively. The sequence of deposition of metals on applying potential is: (a) A, C, B, D (b) B, D, C, A

15. A current of 10.0 A flows for 2.00 h through an electrolytic cell containing a molten salt of metal X. This results in the decomposition of 0.250 mol of metal X at the cathode. The oxidation state of X in the molten salt is: (F = 96,500 C)

(a)
$$1+$$
 (b) $2+$ (c) $3+$ (d) $4+$

- **16.** Which of the following will form a cell with the highest voltage?
 - (a) 1M Ag, $1M \text{ Co}^2$
 - (b) 2MAg, $2MCo^2$
 - (c) 0.1 MAg, 2 MCo^2
 - (d) 2MAg, $0.1MCo^2$
- 17. The standard EMF for the cell reaction,

Zn $Cu^2 \longrightarrow Cu$ Zn² is 1.1 volt at 25°C. The EMF for the cell reaction, when 0.1 M Cu²⁺ and 0.1 M Zn²⁺ solutions are used, at 25°C is (a) 1.10V (b) 0.10V (c) -1.10V (d) -0.110V Consider the following reactions

18. Consider the following reactions

(i)
$$\operatorname{Cd}^{2+}(aq) + 2e^{-} \longrightarrow \operatorname{Cd}(s), E^{\circ} = -0.40 \,\mathrm{V}$$

(ii) $\operatorname{Ag}^+(aq) \quad e^- \longrightarrow \operatorname{Ag}(s), E^\circ = 0.80 \,\mathrm{V}$

For the galvanic cell involving the above reactions. Which of the following is not correct?



- (a) Molar concentration of the cation in the cathodic compartment changes faster than that of the cation in the anodic compartment.
- (b) E_{cell} increase when Cd^2 solution is diluted.
- (c) Twice as many electrons pass through the cadmium electrode as through silver electrode.
- (d) E_{cell} decreases when Ag solution is diluted.

19. Given that:
$$E_{Ag/Ag} = 0.80V$$
 and $[Ag^+] = 10^{-3}M$;

$$E_{Hg_2^2/Hg} = 0.785 V \text{ and } [Hg_2^{2+}] = 10^{-1} M$$

which is true for the cell reaction

$$2\text{Hg}(l) + 2\text{Ag}(aq) \rightarrow 2\text{Ag}(s) \text{Hg}_2^2(aq)?$$

- (a) The forward reaction is spontaneous
- (b) The backward reaction is spontaneous
- (c) $E_{cell} = 0.163 V$
- (d) $E_{cell} = 1.585 V$

20. Use the following standard electrode potentials, calculate ΔG° in kJ/mol for the indicated reaction:

5Ce⁴ (aq) Mn² (aq) 4H₂O(1)
$$\longrightarrow$$

5Ce³⁺(aq) MnO₄⁻(aq) 8H (aq)
MnO₄⁻(aq) 8H⁺(aq) 5e⁻ \longrightarrow

 Mn^2 (aq) $4H_2O(l); E^\circ = +1.51 V$

 $\operatorname{Ce}^{4+}(\operatorname{aq}) + \operatorname{e}^{-} \longrightarrow \operatorname{Ce}^{3}(\operatorname{aq}); \operatorname{E}^{\circ} = +1.61 \operatorname{V}$

(a) -9.65 (b) -24.3 (c) -48.25(d) -35.221. If 0.01 M solution of an electrolyte has a resistance of 40 ohms in a cell having a cell constant of 0.4 cm⁻¹, then its molar conductance

in $ohm^{-1} cm^2 mol^{-1}$ is

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

- 22. When a concentrated solution of an electrolyte is diluted
 - (a) its specific conductance increases.
 - (b) its equivalent conductivity decreases.
 - (c) its specific conductivity decreases and equivalent conductivity increases.
 - (d) both specific and equivalent conductivity increase.
- 23. The limiting equivalent conductivity of NaCl, KCl and KBr are 126.5, 150.0 and 151.5 S cm² eq⁻¹, respectively. The limiting equivalent ionic conductivity for Br⁻ is 78 S cm²eq⁻¹. The limiting equivalent ionic conductivity for Na⁺ ions would be :

(a) 128 (c) 49 (b) 125 (d) 50

- 24. The ionic conductivity of Ba^{2+} and Cl^{-} at infinite dilution are 127 and 76 ohm⁻¹ cm² eq⁻¹ respectively. The equivalent conductivity of BaCl₂ at infinity dilution (in $ohm^{-1}cm^2 eq^{-1}$) would be : (a) 203 (b) 279 (c) 101.5 (d) 139.5
- 25. Aluminium oxide may be electrolysed at 1000°C to furnish aluminium metal (At. Mass = 27 amu; 1 Faraday = 96,500 Coulombs). The cathode

reaction is- $Al^{3+} + 3e^- \rightarrow Al^{*}$

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

- (a) 5.49×10^1 C (b) 5.49×10^4 C
- (c) 1.83×10^7 C (d) 5.49×10^7 C

- When electric current is passed through acidified 26. water, 112 mL of hydrogen gas at STP collected at the cathode in 965 seconds. The current passed Tricky in amperes is
- (c) 0.1 (a) 1.0 (d) 2.0 (b) 0.5 27. On passing current through two cells, connected in series containing solution of AgNO₃ and CuSO₄, 0.18 g of Ag is deposited. The amount of the Cu deposited is:
 - (a) 0.529 g (b) 10.623 g
 - (c) $0.0529 \,\mathrm{g}$ (d) 1.2708 g
- 28. 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)
$$\frac{1000 \text{ x}}{\text{N}}$$
 (b) $\frac{1000}{\text{Nx}}$

(c)
$$\frac{1000N}{x}$$
 (d) $\frac{Nx}{1000}$

- For an electrolyte solution of 0.05 mol L^{-1} , 29. the conductivity has been found to be 0.0110 S cm^{-1} . The molar conductivity is
 - (a) $0.055 \,\mathrm{S} \,\mathrm{cm}^2 \,\mathrm{mol}^{-1}$
 - (b) $550 \,\mathrm{S} \,\mathrm{cm}^2 \,\mathrm{mol}^{-1}$
 - (c) $0.22 \,\mathrm{S} \,\mathrm{cm}^2 \,\mathrm{mol}^{-1}$
 - (d) $220 \,\mathrm{S} \,\mathrm{cm}^2 \,\mathrm{mol}^{-1}$
- The highest electrical conductivity of the following aqueous solutions is of
 - (a) 0.1 M difluoroacetic acid
 - (b) 0.1 M fluoroacetic acid
 - (c) 0.1 M chloroacetic acid
 - (d) 0.1 M acetic acid
- 31. A current of 2.0 A passed for 5 hours through a molten metal salt deposits 22.2 g of metal (At wt. = 177). The oxidation state of the metal in Toughnut

(a) +1 (b) +2 (c) +3 The specific conductivity of N/10 KCl solution 32. at 20°C is 0.212 ohm⁻¹ cm⁻¹ and the resistance of the cell containing this solution at 20°C is 55 ohm.

The cell constant is

- (a) $4.616 \,\mathrm{cm}^{-1}$ (b) $11.66 \,\mathrm{cm}^{-1}$
- (c) $2.173 \,\mathrm{cm}^{-1}$ (d) $3.324 \,\mathrm{cm}^{-1}$

- 30.

2

- **33.** The equivalent conductance at infinite dilution of a weak acid such as HF
 - (a) can be determined by extrapolation of measurements of dilute solutions of HCl, HBr and HI
 - (b) can be determined by measurement of very dilute HF solutions
 - (c) can be determined from measurements of dilute solutions of NaF, NaCl and HCl
 - (d) is an undefined quantity
- **34.** Conductance of 0.1 M KCl (conductivity = XOhm⁻¹cm⁻¹) filled in a conductivity cell is Y

Ohm⁻¹. If the conductance of 0.1 M NaOH filled

in the same cell is Z Ohm⁻¹, the molar conductance of NaOH will be **Toughnut**

(a)
$$10^3 \frac{XZ}{Y}$$
 (b) $10^4 \frac{XZ}{Y}$
(c) $10 \frac{XZ}{Y}$ (d) $0.1 \frac{XZ}{Y}$

35. In electrolytic reduction of a nitroarene with 50% current efficiency 20.50 g of the compound is reduced by 2×96500 C of electric charge. The molar mass of the compound is

- (c) 10.2 g (d) 20.5 g
- **36.** Electrolysis of NaCl solution with inert electrodes for certain period of time gave 600 cm^3 of 1.0 M NaOH in the electrolytic cell. During the same period 31.80 g of copper was deposited in a copper voltmeter in series with the electrolytic cell. What is the percent current efficiency in the electrolytic cell ? (At. wt. of Cu = 63. 6)

37. Given the ionic conductance of $\begin{bmatrix} COO^- \\ | \\ COO^- \end{bmatrix}$, K⁺, and COO^-

Na⁺ are 74, 50, and 73 cm² ohm⁻¹ eq⁻¹, respectively. The equivalent conductance at

infinite dilution of the salt \mid is COOK

- (a) $197 \,\mathrm{cm}^2 \,\mathrm{ohm}^{-1} \,\mathrm{eq}^{-1}$
- (b) $172 \text{ cm}^2 \text{ ohm}^{-1} \text{ eq}^{-1}$
- (c) $135.5 \text{ cm}^2 \text{ ohm}^{-1} \text{ eq}^{-1}$
- (d) $160.5 \text{ cm}^2 \text{ ohm}^{-1} \text{ eq}^{-1}$

38. On passing a current of 1.0 ampere for 16 min and 5 sec through one litre solution of $CuCl_2$, all copper of the solution was deposited at cathode. The strength of $CuCl_2$ solution was (Molar mass of Cu=63.5; Faraday constant = 96500 Cmol⁻¹)



- (a) 0.01 N (b) 0.01 M
- (c) 0.02 M (d) 0.2 N
- **39.** Electrode potential of the half cell Pt (s) | Hg (l) | Hg₂Cl₂(s) | Cl⁻ (aq) can be increased by :
 - (a) Increasing [Cl⁻]
 - (b) Decreasing [Cl⁻]
 - (c) Increasing $Hg_2Cl_2(s)$

(d) Decreasing Hg (l)

- **40.** Identify the correct statement :
 - (a) Corrosion of iron can be minimized by forming a contact with another metal with a higher reduction potential
 - (b) Iron corrodes in oxygen free water
 - (c) Corrosion of iron can be minimized by forming an impermeable barrier at its surface
 - (d) Iron corrodes more rapidly in salt water because its electrochemical potential is higher
- **41.** During the charging of lead storage battery, the reaction at anode is represented by :

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

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

$$4H^+$$
 $2e^-$

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

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

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

(a) NH₄ (b) N₂ (c) H₂ (d) Cl₂
43. Given
Fe³⁺ (aq) + e⁻
$$\rightarrow$$
 Fe²⁺ (aq); E^o = + 0.77 V

Al³⁺ (aq) + 3e⁻ \rightarrow Al(s); E^o = -1.66 V Br₂(aq) + 2e⁻ \rightarrow 2Br⁻; E^o = +1.09 V Considering the electrode potentials, which of the following represents the correct order of reducing power?

(a) $Fe^{2+} < Al < Br^-$ (b) $Br^- < Fe^{2+} < Al$ (c) $Al < Br^- < Fe^{2+}$ (d) $Al < Fe^{2+} < Br^-$ 44. How many electrons would be required to deposit 6.35 g of copper at the cathode during the electrolysis of an aqueous solution of copper sulphate? (Atomic mass of copper = 63.5 u, N_A = Avogadro's constant):

(a)
$$\frac{N_A}{20}$$
 (b) $\frac{N_A}{10}$ (c) $\frac{N_A}{5}$ (d) $\frac{N_A}{2}$

45. Given :

$$E^{o}_{\frac{1}{2}Cl_{2}/Cl^{-}}$$
 1.36 V, $E^{o}_{Cr^{3}/Cr}$ -0.74 V,

 $E^{o}_{Cr_{2}O_{7}^{2^{-}}/Cr^{3+}}$ 1.33 V, $E^{o}_{MnO_{4}^{-}/Mn^{2}}$ 1.51V The correct order of reducing power of the

species (Cr, Cr³⁺, Mn²⁺ and Cl⁻) will be :

- (a) $Mn^{2+} < Cl^{-} < Cr^{3+} < Cr$
- (b) $Mn^{2+} < Cr^{3+} < Cl^{-} < Cr$ (c) $Cr^{3+} < Cl^{-} < Mn^{2+} < Cr$

d)
$$Cr^{3+} < Cl^{-} < Cr < Mn^{2+}$$

46. The oxidation potential of a hydrogen electrode at pH = 10 and $p_{H_2} = 1$ atm is

(a)
$$-0.59 V$$
 (b) $0.00 V$

(d) 0.059V (c) +0.59 V

47. If the E°_{cell} for a given reaction has a negative value, then which of the following gives the correct relationships for the values of ΔG° and



(a)
$$\Delta G^{\circ} > 0; K_{eq} > 1$$
 (b) $\Delta G^{\circ} < 0; K_{eq} > 1$
(c) $\Delta G^{\circ} < 0; K_{eq} < 1$ (d) $\Delta G^{\circ} > 0; K_{eq} < 1$

- **48.** The EMF of the cell $Tl/Tl^{+}(0.001M) \parallel Cu^{2+}(0.01M)/Cu \text{ is } 0.83$. The cell EMF can be increased by
 - (a) Increasing the concentration of Tl^+ ions.
 - (b) Increasing the concentration of Cu^{2+} ions.
 - (c) Increasing the concentration of Tl $^+$ and Cu²⁺ ions.
 - (d) None of these
- **49.** Small quantities of solutions of compounds TX, TY and TZ are put into separate test tubes containing X, Y and Z solution. TX does not react with any of these. TY reacts with both X and Z. TZ reacts with X. The decreasing order of state of oxidation of the anions X⁻, Y⁻, Z⁻ is

(a) Y⁻, Z⁻, X⁻ (b) Z⁻, X⁻, Y

(c)
$$Y^-, X^-, Z^-$$
 (d) X^-, Z^-, Y^-

50. An unknown metal M displaces nickel from nickel (II) sulphate solution but does not displace manganese from manganese sulphate solution. Which order represents the correct order of reducing power?

(a)
$$Mn > Ni > M$$
 (b) $Ni > Mn > M$

(c)
$$Mn > M > Ni$$
 (d) $M > Ni > Mn$

51. What is the e.m.f for the given cell?

Cr | Cr³ (1.0M) || Co² (1.0M) | Co
(
$$E^{\circ}$$
 for Cr³ /Cr = -0.74 volt and E° for
Co² /Co = -0.28 volt)
(a) -0.46 volt (b) -1.02 volt

- (c) +0.46 volt (d) 1.66 volt
- The standard electrode potential (E°) for 52. OCl^{-}/Cl^{-} and $Cl^{-}/\frac{1}{2}Cl_{2}$ respectively are 0.94

V and -1.36 V. The E° value for OCl⁻ / $\frac{1}{2}$ Cl₂ will be

(a)	-0.42 V	(b)	-2.20 V
(c)	0.52 V	(d)	1.04V

For the reduction of silver ions with copper metal, 53. the standard cell potential was found to be +0.46V at 25°C. The value of standard Gibbs energy,

 ΔG° will be (F = 96500 C mol⁻¹)

(a)
$$-89.0 \text{ kJ}$$
 (b) -89.0 J
(c) -44.5 kJ (d) -98.0 kJ

In a cell that utilises the reaction 54.

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

- addition of H₂SO₄ to cathode compartment, will
- (a) increase the E and shift equilibrium to the right
- (b) lower the E and shift equilibrium to the right
- (c) lower the E and shift equilibrium to the left
- (d) increase the E and shift equilibrium to the left
- 55. For a cell reaction involving two electron change, the standard EMF of the cell is 0.295 V at 2°C. The equilibrium constant of the reaction at 25°C will be:

(a)	29.5×10^{-2}	(b)	10
(c)	$1 imes 10^{10}$	(d)	2.95×10^{-10}

56. On the basis of the information available from the reaction $\frac{4}{3}$ Al $O_2 \rightarrow \frac{2}{3}$ Al₂O₃, ΔG = -827 kJ mol^{-1} of O₂ the minimum e.m.f required to carry out an electrolysis of Al₂O₃ is $(F = 96500 \text{ C mol}^{-1})$ *為 Critical g

(a)
$$856V$$
 (b) $214V$ (c) $428V$ (d) $642V$

(a)

will be

(a)
$$8.56V$$
 (b) $2.14V$ (c) $4.28V$ (d) $6.42V$
57. The e.m.f. of a Daniell cell at 298 K is E₁.

$$\operatorname{Zn} \begin{vmatrix} \operatorname{ZnSO}_4 \\ (0.01 \text{ M}) \end{vmatrix} \begin{vmatrix} \operatorname{CuSO}_4 \\ (1.0 \text{ M}) \end{vmatrix} \operatorname{Cu}$$

When the concentration of ZnSO₄ is 1.0 M and that of $CuSO_4$ is 0.01 M, the e.m.f. changed to E_2 . What is the relationship between E_1 and E_2 ?

(a) $E_2 = 0 \neq E_1$ (b) $E_1 = E_2$

(c)
$$E_1 \quad E_2$$
 (d) $E_1 \quad E_2$

Mark the false statement? 58.

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

potential is
$$E = E^o - \frac{RT}{nF} \log a_{M^n}$$

- (d) The efficiency of a hydrogen-oxygen fuel cell is 23%
- The electrode potential $E_{(Zn^2/Zn)}$ of a zinc 59. electrode at 25°C with an aqueous solution of 0.1

M ZnSO₄ is $[E_{(Zn^2/Zn)} = -0.76$ V. Assume

$$\frac{2.303\text{RT}}{\text{F}} = 0.06 \text{ at } 298 \text{ K}].$$

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

- 60. The reduction potential (in volt) of a hydrogen electrode set up with a 2×10^{-2} M aqueous solution of a weak mono basic acid $(K_a = 5 \times 10^{-5})$ at one atmosphere and 25°C is (a) +0.09 (b) +0.18 (c) -0.09 (d) -0.18
- **61.** A solution of copper sulphate $(CuSO_4)$ is electrolysed for 10 minutes with a current of 1.5 amperes. The mass of copper deposited at the cathode (at. mass of Cu = 63u) is :
 - (a) 0.3892g (b) 0.2938g

(c) 0.2398g (d) 0.3928g

62. A battery is constructed of Cr and Na₂Cr₂O₇. The unbalanced chemical equation when such a battery discharges is following:

 $Na_2Cr_2O_7$ Cr H \rightarrow Cr³ H₂O Na If one Faraday of electricity is passed through the battery during the charging, the number of moles of Cr³⁺ removed from the solution is Tricky

$$\frac{4}{3}$$
 (b) $\frac{1}{3}$ (c) $\frac{3}{3}$ (d) $\frac{2}{3}$

The standard potentials of Ag⁺/Ag, Hg₂²⁺/2Hg, **63**. Cu^{2+}/Cu and Mg^{2+}/Mg electrodes are 0.80, 0.79,0.34 and -2.37 V, respectively. An aqueous solution which contains one mole per litre of the salts of each of the four metals is electrolyzed. With increasing voltage, the correct sequence of deposition of the metals at the cathode is

(a) Ag, Hg, Cu, Mg (b) Cu, Hg, Ag only (c) Ag, Hg, Cu only (d) Mg, Cu, Hg, Ag

- When electric current is passed through a cell **64**. having an electrolytic solution, the cations move towards the cathode and anions towards the anode. If anode is pulled out from the solution
 - (a) the cations and anions will move towards the cathode
 - (b) the cations will continue to move towards cathode and anions will stop moving
 - both the cations and anions will stop moving (c)
 - (d) the cations and anions will start moving randomly.
- 65. At 298K the standard free energy of formation of H_2O (1) is -237.20 kJ/mol while that of its ionisation into H⁺ ion and hydroxyl ions is 80 kJ/ mol, then the emf of the following cell at 298 K

Toughnut

[Take Faraday constant F = 96500 C]

 $H_2(g, 1 bar) | H^+(1M) | | OH^-(1M) | O_2(g, 1 bar)$ (b) 0.81 V (a) 0.40V

- (c) 1.23 V (d) -0.40 V
- A concentration cell is a galvanic cell in which **66**.



- (a) decrease in free energy in a spontaneous chemical process appears as electrical energy
- decrease in free energy in a spontaneous (b) physical process appears as electrical energy
- decrease in free energy in a spontaneous (c) physical or chemical process appears as electrical energy
- (d) a non-spontaneous physical or chemical process produces electrical energy.

- **67.** In the silver plating of copper, K[Ag(CN)₂] is used instead of AgNO₃. The reason is
 - (a) a thin layer of Ag is formed on Cu
 - (b) more voltage is required
 - (c) Ag⁺ ions are completely removed from solution
 - (d) less availability of Ag⁺ ions, as Cu cannot displace Ag from [Ag(CN)₂]⁻ ion
- **68.** 1.0 L each of a buffer containing 1 mole NH_3 and

1 mol of NH_4 were placed in the cathodic and anodic half-cells and 965 C of electricity was passed. If anodic and cathodic half cells reactions involve oxidation and reduction of water only as

$$2H_2O \longrightarrow 4H \qquad O_2 - 4e^{-3}$$

 $2H_2O + 2e^- \longrightarrow H_2$ 2OH Then pH of

Tricky

- (a) cathodic solution will increase
- (b) anodic solution will decrease
- (c) both the solutions will remain practically constant
- (d) both the solutions will increase
- 69. The emf of a particular voltaic cell with the cell

reaction Hg_2^2 $H_2 \implies 2Hg + 2H$ is 0.65 V. The maximum electrical work of this cell when 0.5 g of H_2 is consumed.

(a)
$$-3.12 \times 10^4 \text{ J}$$
 (b) $-1.25 \times 10^5 \text{ J}$
(c) $25.0 \times 10^6 \text{ J}$ (d) None of these

- **70.** A cell reaction would be spontaneous if the cell potential and $\Delta_r G$ are respectively :
 - (a) positive and negative
 - (b) negative, negative
 - (c) zero, zero
 - (d) positive, zero
- 71. The Nernst equation $E = E^{\circ} RT/nF \ln Q$ indicates that the Q will be equal to equilibrium constant K_c when :

(a)
$$E = E^{\circ}$$
 (b) $RT / nF = 1$

(c)
$$E = zero$$
 (d) $E^{\circ} = 1$

72. Thermodynamic efficiency of a cell is given by :

(a)
$$\frac{\Delta H}{\Delta G}$$
 (b) $\frac{nFE}{\Delta G}$ (c) $\frac{-nFE}{\Delta H}$ (d) nFE°

73. A graph was plotted between molar conductivity of various electrolytes (NaCl, HCl and NH₄OH) and \sqrt{c} (in mol L⁻¹). Correct set is :

Tricky



- (a) I (NaCl), II (HCl), III, (NH₄OH)
- (b) I (HCl), II (NaCl), III, (NH₄OH)
- (c) I (NH₄OH), II (NaCl), III, (HCl)
- (d) I (NH₄OH), II (HCl), III, (NaCl)

74. HNO₃ (aq) is titrated with NaOH(aq) conductometrically, graphical representation of the titration is :



- **75.** In the electrolysis of water, one faraday of electrical energy would liberate
 - (a) one mole of oxygen
 - (b) one gram atom of oxygen
 - (c) 8 g oxygen
 - (d) 22.4 lit. of oxygen
- 76. Resistance of a conductivity cell filled with a solution of an electrolyte of concentration 0.1 M is 100 Ω . The conductivity of this solution is 1.29 S m⁻¹. Resistance of the same cell when filled with 0.02 M of the same solution is 520 Ω . The molar conductivity of 0.02 M solution of electrolyte will be **Critical**

Thinking

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

77. $\Lambda_{\text{ClCH}_2\text{COONa}}$ 224 ohm⁻¹cm²g eq⁻¹, Λ_{NaCl} 38.2 ohm⁻¹cm²g eq⁻¹, $\Lambda_{\rm HCl}$ 203 ohm⁻¹cm²g eq⁻¹,

What is the value of $\Lambda_{\text{ClCH}_2\text{COOH}}$

- (a) $288.5 \text{ ohm}^{-1}\text{cm}^2\text{g eq}^{-1}$
- (b) $289.5 \text{ ohm}^{-1}\text{cm}^2\text{g eq}^{-1}$
- (c) $388.8 \text{ ohm}^{-1}\text{cm}^{2}\text{g eq}^{-1}$
- (d) $59.5 \text{ ohm}^{-1} \text{cm}^2 \text{g eq}^{-1}$
- 78. 1.08 g of pure silver was converted into silver nitrate and its solution was taken in a beaker. It was electrolysed using platinum cathode and silver anode. 0.01 Faraday of electricity was passed using 0.15 volt above the decomposition potential of silver. The silver content of the beaker after the above shall be
 - (a) 0 g (b) 0.108 g
 - (c) 1.08 g (d) None of these
- **79.** Specific conductance of 0.1 M sodium chloride solution is 1.06×10^{-2} ohm⁻¹ cm⁻¹. Its molar conductance in ohm⁻¹ cm² mol⁻¹ is
 - (a) 1.06×10^2 (b) 1.06×10^3
 - (c) 1.06×10^4 (d) 5.3×10^2
- **80.** In the electrolysis of CuCl₂ solution, the mass of the cathode increased by 3.2 g. What occured at the copper anode?
 - (a) 0.12 litre of Cl₂ was liberated
 - (b) 0.56 litre of O_2 was liberated
 - (c) $0.1 \text{ mol } \text{Cu}^{2+2}$ passed into the solution.
 - (d) $0.05 \text{ mol of } \text{Cu}^{2+}$ passed into the solution.
- 81. Given

 $\rm E_{Cu^2/Cu}=0.34V$, $\rm E_{Cu^2/Cu}=0.15V$ Standard electrode potential for the half cell Cu⁺/Cu is

(a) 0.38V (b) 0.53V (c) 0.19V (d) 0.49V 82. At 298 K, the standard reduction potentials are

> 1.51 V for MnO_4^- | Mn^{2+} , 1.36 V for Cl_2 | Cl^- , 1.07 V for $Br_2 | Br^-$, and 0.54 V for $I_2 | I^-$. At pH = 3, permanganate is expected to oxidize : (**DT**

$$\left(\frac{\mathbf{KI}}{\mathbf{E}} \quad 0.059 \mathbf{V}\right)$$

(a) Cl⁻, Br⁻ and I⁻ (b) Br⁻ and I⁻

(c) Cl^{-} and Br^{-} (d) I^- only

83. Electrode potentials (E°) are given below : $Cu^+/Cu = +0.52 V_{,}$ $F_{e}^{3+}/Fe^{2+} = +0.77 V_{e}^{3+}$ $\frac{1}{2}I_2(s)/I^-$ 0.54V,

 $Ag^{+}/Ag = +0.88V.$

Based on the above potentials, strongest oxidizing agent will be :

(a) Cu^+ (b) Fe^{3+} (c) Ag^+ (d) I_2 84. Which of the following is a highly corrosive salt?

- (b) HgCl₂ (a) Hg₂Cl₂ (d) PbCl₂ (c) FeCl₂
- 85. The Zn acts as sacrificial or cathodic protection to prevent rusting of iron because :
 - (a) E_{OP} of Zn E_{OP} of Fe
 - E_{OP} of Zn E_{OP} of Fe (b)
 - (c) E_{OP} of Zn E_{OP} of Fe
 - (d) Zn is cheaper than iron
- 86. Among the following cells:
 - (I) Leclanche cell
 - (II) Nickel-Cadmium cell
 - (III) Lead storage battery
 - (IV) Mercury cell
 - primary cells are
 - (a) I and II
 - (b) I and III (d) I and IV (c) II and III
- How long a current of 3 amperes has to be passed 87. through a solution of AgNO₃ to coat a metal surface of 80 cm² and 0.005 mm thick layer. Density of Ag is 10.5 g cm^{-3}
 - (a) 125.1 seconds (b) 12.5 seconds
 - (d) 200 seconds (c) 155.2 seconds
- 88. Al₂O₃ is reduced by electrolysis at low potentials and high currents. If 4.0×10^4 amperes of current is passed through molten Al₂O₂ for 6 hours, what mass of aluminium is produced? (Assume 100% current efficiency. At. mass of $Al = 27 \text{ g mol}^{-1}$)
 - (a) 8.1×10^4 g (b) 2.4×10^5 g
 - (c) 1.3×10^4 g (d) 9.0×10^3 g
- The resistance of 1 N solution of acetic acid 89. is 250 ohm, when measured in a cell of cell constant 1.15 cm⁻¹. The equivalent conductance (in $ohm^{-1} cm^2 equiv^{-1}$) of 1 N acetic acid will be (a) 4.6 (b) 9.2 (c) 18.4 (d) 0.023

Specific conductance of 0.1 *M* HA is 3.75×10^{-4} 90. ohm⁻¹ cm⁻¹. If λ^{∞} (HA) 250 ohm⁻¹ cm² mol⁻¹, the dissociation constant K_a of HA is :



(a) 1.0×10^{-5} (b) 2.25×10^{-4} (c) 2.25×10^{-5} (d) 2.25×10^{-13}

- **91.** Equivalent conductance at infinite dilution, λ° of NH₄Cl, NaOH and NaCl are 128.0, 217.8 and 109.3 $ohm^{-1} cm^2 eq^{-1}$ respectively. The equivalent conductance of 0.01 N NH₄OH is 9.30 ohm⁻¹ $cm^2 eq^{-1}$, then the degree of ionization of NH₄OH at this temperature would be
- (a) 0.04 (b) 0.1 (c) 0.39 (d) 0.62 92. I_{2} (s) |I⁻ (0.1 M) half cell is connected to a $H^+(aq)|H_2(1 bar)|$ Pt half cell and e.m.f. is found to be 0.7714 V. if $E_{I_2|I^-}$ 0.535 V, find the pH of

H⁺ | H₂ half-cell.

Toughnut

(a) 1 (c) 5⁶² (d) 7 (b) 3 93. Consider the following standard electrode potentials and calculate the equilibrium constant at 25° C for the indicated disproportion nation reaction :

$$3Mn^2$$
 (aq) $\longrightarrow Mn(s) 2Mn^3$ aq
 $Mn^{3+}(aq) + e^- \longrightarrow Mn^2$ (aq); E 1.51 V

 $Mn^{2+}(aq) + 2e^{-} \longrightarrow Mn(s); E -1.185 V$ (a) 1.2×10^{-43} (b) 2.4×10^{-73}

- (c) 6.3×10^{-92} (d) 1.5×10^{-62}
- 94. In a hydrogen-oxygen fuel cell, combustion of hydrogen occurs to
 - (a) produce high purity water
 - (b) create potential difference between two electrodes
 - (c) generate heat
 - (d) remove adsorbed oxygen from elctrode surfaces
- 95. The standard reduction potentials at 25°C of Li /Li, Ba² /Ba, Na /Na and Mg² /Mg are -3.03, -2.73, -2.71 and -2.37 volt respectively.

Which one of the following is the strongest oxidising agent?

- (a) Na^+ (b) Li^+ (c) Ba^{2+} (d) Mg^{2+}
- 96. The $\frac{E^{\circ}}{M^{3}}$ / $\frac{M^{2}}{M^{2}}$ values for Cr, Mn, Fe and Co are - 0.41, + 1.57, + 0.77 and + 1.97V respectively. For which one of these metals the change in oxidation state from +2 to +3 is easiest?

- 97. The solution of $CuSO_4$ in which copper rod is immersed is diluted to 10 times. The reduction electrode potential Toughnut
 - (a) Increases by 30 mV
 - (b) Decreases by 30 mV
 - (c) Increases by 59 mV
 - (d) Decreases by 59 mV
- 98. For a given reaction: $M(x+n) + ne^{-} \longrightarrow M^{x+}$, E_{red}^{-} is known along with $M^{(x+n)}$ and M^{x+} ion concentrations. Then
 - (a) n can be evaluated
 - (b) x can be evaluated
 - (c) (x+n) can be evaluated
 - (d) n, x, (x+n) can be evaluated
- The emf of the cell Pt $Cl_2(g) (P_1 atm)/Cl^-(aq)$ 99.
 - $(1M)/Cl_2(g)(P_2 atm)$ Pt will be positive when

Tricky

- (a) $P_1 = P_2$ (b) $P_1 < P_2$ (c) $P_1 > P_2$ (d) None of these **100.** Equivalent conductivity can be expressed in terms of specific conductance (κ) and concentration (N) in gram equivalent per dm⁻³ as: 1000

(a)
$$\kappa \times N$$
 (b) $\frac{\kappa \times 1000}{N}$
(c) $\frac{\kappa \times N}{1000}$ (d) $\kappa \times N \times 1000$

Answer KEYs

1	(c)	11	(a)	21	(d)	31	(c)	41	(b)	51	(c)	61	(b)	71	(c)	81	(b)	91	(a)
2	(b)	12	(d)	22	(c)	32	(b)	42	(c)	52	(a)	62	(c)	72	(c)	82	(b)	92	(b)
3	(a)	13	(c)	23	(d)	33	(c)	43	(d)	53	(a)	63	(c)	73	(b)	83	(c)	93	(c)
4	(a)	14	(c)	24	(b)	34	(b)	44	(c)	54	(a)	64	(d)	74	(a)	84	(b)	94	(b)
5	(d)	15	(c)	25	(d)	35	(a)	45	(a)	55	(c)	65	(a)	75	(c)	85	(b)	95	(d)
6	(c)	16	(d)	26	(a)	36	(c)	46	(c)	56	(b)	66	(b)	76	(a)	86	(d)	96	(c)
7	(b)	17	(a)	27	(c)	37	(c)	47	(d)	57	(b)	67	(d)	77	(c)	87	(a)	97	(b)
8	(c)	18	(c)	28	(b)	38	(a)	48	(b)	58	(c)	68	(c)	78	(a)	88	(a)	98	(a)
9	(d)	19	(b)	29	(d)	39	(b)	49	(a)	59	(b)	69	(a)	79	(a)	89	(a)	99	(b)
10	(a)	20	(c)	30	(a)	40	(c)	50	(c)	60	(d)	70	(a)	80	(d)	90	(c)	100	(b)



Hints & Solutions



- 1. (c) Cathode and anode are (-) ve terminals in 11. (a) Apply Nernst equation to the reaction electrolytic and galvanic cell.
- 2. **(b)** Ag becomes oxidized and Br- becomes reduced.

3. (a)
$$A: H_2 g \Longrightarrow 2H^+ aq 2e^-$$

$$C: Cu^{2+}$$
 aq $2e^- \Longrightarrow Cu s$

- (a) Since oxidation potential of Zn is highest 4. hence strongest reducing agent.
- 5. (d) The reduction potential of NO_3^- ion is more than H⁺ ion.
- (c) Gold having higher E_{Red}^{o} and oxidises 6. $Fe \longrightarrow Fe^2$.

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

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

$$E_{cell}^{o} = E_{cathode}^{o} - E_{anode}^{o}$$

$$=0.15 - (-0.74) = +0.8$$

= 0.15 - (-0.74) = +0.89 V(c) Given Fe³⁺/Fe²⁺ = +0.77 V 8. and $I_2/2I^-=0.536V$ $2 \left(e^{-2} + F e^{3+} \longrightarrow F e^{2+} \right) E^{\circ} = 0.77 V$ $2I^- \longrightarrow I_2 + 2e^- \qquad E^\circ = -0.536 V$ $2Fe^{3+} + 2I^{-} \xrightarrow{2} 2Fe^{2+} + I_{2}$ $E^{\circ} = E^{\circ}_{ox} + E^{\circ}_{red}$ = 0.77 - 0.536 = 0.164 VSo, reaction will take place.

9. (d)
$$E_{red} = E_{red}^{o} + \frac{0.591}{n} \log[M^{n+}]$$

Lower the concentration of Mⁿ⁺, lower is the reduction potential. Hence order of reduction potential is : Q > R > S > P

10. (a)
$$2Ag^+ + H_2 \longrightarrow 2H^+ + 2Ag$$

$$E \quad E \quad -\frac{0.0591}{2} \log \frac{[\text{H}]^2}{\text{P}_{\text{H}_2} \times [\text{Ag}]^2}$$

$$0.222 = 0.7995 - \frac{0.0591}{2} \log \frac{1}{[\text{Ag}]^2}$$

$$[\text{Ag}^+] = 10^{-9.8}$$

$$K_{sp} = [\text{Ag}^+] [\text{CI}^-] = (10^{-9.8}) \times (1) = 10^{-9.8}$$

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

$$E_{\text{cell}} = E^{\circ} + \frac{0.059}{2} \log \frac{[\text{Sn}^2]}{[\text{Pb}^2]}$$
$$\log \frac{[\text{Sn}^2]}{[\text{Pb}^2]} = \frac{0.01 \times 2}{0.059} = 0.3 \quad (\because E_{\text{cell}} = 0)$$

or
$$\frac{[\operatorname{Sn}^2]}{[\operatorname{Pb}^2]}$$
 antilog (0.3)

or

12. (d) For a concentration cell having different concentrations of ions.

$$E = -\frac{0.0591}{n} \log \frac{c_1}{c_2}$$

If all the concentrations are identical then obviously the cell voltage is zero. But as the pH of 0.1 M HCl (strong acid) and pH of 0.1M CH₃COOH is (weak acid) not same, therefore the cell voltage is not zero.

- EMF of galvanic cell = 1.1 volt 13. (c) If $E_{ext} < EMF$ then electron flows steadily from anode to cathode while If $E_{ext} > EMF$ then electron flows from cathode to anode as polarity is changed.
- 14. (c) As the value of reduction potential decreases the reducing power increases i.e. C < B < D < A(0.85).(0.6).(-0.76)(-1.2)
- 15. (c) According to Faraday's first law of electrolysis

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

Where $E =$ equivalent weight
= $\frac{\text{mol. mass of metal (M)}}{\text{avidation state of metal (w)}}$

oxidation state of metal (x)Substituting the value in the formula

$$W = \frac{M}{x} \times \frac{i \times t}{96500}$$

or
$$x = \frac{M}{W} \times \frac{i \times t}{96500} = \frac{10 \times 2 \times 60 \times 60}{96500 \times 0.250} = 3$$

[Given : no. of moles = $\frac{M}{W} = 0.250$]
Hence exidation state of metal is (+3)

Hence oxidation state of metal is (+3)

16. (d) Oxidation potential of Co is more than Ag, hence cell reaction will be

Co
$$2Ag \longrightarrow Co^2 2Ag$$

 $E_{cell} = E_{cell}^o - \frac{RT \ln [Co^2]}{nF [Ag]^2}$, the lesser the
value of the factor $\frac{[Co^2]}{[Ag]^2}$ greater will be
value of E_{max}

17. (a) Since concentration of ions is the same hence $E_{\text{cell}} = E_{\text{cell}}^{\text{o}}$

18. (c)
$$\operatorname{Cd}(s) + 2\operatorname{Ag}(aq) \longrightarrow$$

 $\operatorname{Cd}^2(aq) \operatorname{Ag}(s),$

$$E_{\text{cell}} = E_{\text{cell}} - \frac{0.059}{2} \log \frac{[\text{Cd}^2]}{[\text{Ag}]^2}$$

19. (b)
$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0592}{2} \log \frac{[\text{Hg}_2^2 \ (aq)]}{[\text{Ag} \ (aq)]^2}$$

= $(0.80 - 0.785) - \frac{0.0592}{2} \log \frac{10^{-1}}{(10^{-3})^2} = -0.133 \text{V}$
hence backward reaction is spontaneous.

20. (c) $E_{cell} = E_{RP RHS} - E_{RP LHS}$ $= 1.61 - 1.51 \Longrightarrow 0.10 \,\mathrm{V}$ $\Delta G^{\circ} \!=\! -nFE^{\circ} \! \Longrightarrow \! -5 \times 96500 \times 0.10 \, J$ $\Delta G^{\circ} = -48.25 \text{ kJ mol}^{-1}$

21. (d) Molarity=0.01 M ; Resistance=40 ohm;
Cell constant
$$\frac{l}{A} = 0.4 \text{ cm}^{-1}$$
.

Specific conductivity (κ)

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

 $\frac{1000\kappa}{\text{Molarity}}$ Molar conductance (\wedge_m)

$$\frac{1000 \times 0.01}{0.01} \quad 10^3 \, \text{ohm}^{-1} \, \text{cm}^2 \, \text{mol}^{-1}$$

22. (c) On dilution number of ions decrease in unit volume hence specific conductance decreases. But separation between ions also increase hence equivalent conductance increases.

23. (d)
$$\Lambda_{m}^{\infty}(\text{NaBr}) = \Lambda_{m}^{\infty}(\text{NaCl}) + \Lambda_{m}^{\infty}\text{KBr} - \Lambda_{m}^{\infty}(\text{KCl})$$

 $\lambda_{m}^{\infty}(\text{Na}^{+}) + \lambda_{m}^{\infty}(\text{Br}^{-})$ 126.5 151.5 - 150
 $\lambda_{m}^{\infty}(\text{Na}^{-})$ 50 S cm² eq⁻¹.
24. (b) Equivalent conductivity of BaCl₂
 $\Lambda_{m}^{\infty}(\text{BaCl}_{2}) = \lambda_{m}^{\infty}(\text{Ba}^{2+}) + 2\lambda_{m}^{\infty}(\text{Cl}^{-})$
 $= 127 + 2 \times 76 = 279 \text{ ohm}^{-1} \text{ cm}^{2} \text{ eq}^{-1}.$
25. (d) 1 mole of $e^{-} = 1\text{F} = 96500 \text{ C}$
 $27\text{g of Al is deposited by 3 × 96500 \text{ C}$
 $5120 \text{ g of Al will be deposited by}$
 $= \frac{3 \times 96500 \times 5120}{27}$ 5.49×10⁷ C
26. (a) 2H⁺ + 2e⁻ \longrightarrow H₂
 $E_{H}(\text{Eq. wt}) = \frac{2}{2} = 1$
 $1 \text{ g} = \frac{22400}{2} = 11200 \text{ mL (STP)}$
Total charge passed $= \frac{96500 \times 112}{11200} = 965$
 $Q = It = 965$
 $I = \frac{965}{965} = 1 \text{ amp.}$

$$\Rightarrow \frac{W_{Cu}}{0.18} = \frac{63.5}{2} \times \frac{1}{108}$$
$$\Rightarrow W_{Cu} = \frac{63.5 \times 18}{2 \times 108 \times 100} = 0.0529 \text{ g}$$

$$= \kappa = \frac{1}{x}.$$

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

29. (d)
$$\Lambda_{\rm m} = \frac{\kappa \times 1000}{M} = \frac{0.0110 \times 1000}{0.05}$$

= 220 S cm² mol⁻¹

=

30. (a) Difluoroacetic acid being strongest acid will furnish maximum number of ions showing highest electrical conductivity.

31. (c)
$$m = \frac{E.wt \times Q}{96500}$$
;
 $\therefore E.wt = \frac{m \times 96500}{Q} = \frac{22.2 \times 96500}{2 \times 5 \times 60 \times 60} = 60.3$
Oxidation state $\frac{At wt}{Eq. wt.} = \frac{177}{60.3} = 3$
32. (b) Cell constant
 $= \kappa \times R$
 $= 0.212 \text{ ohm}^{-1} \text{ cm}^{-1} \times 55 \text{ ohm}$
 11.66 cm^{-1}
33. (c) Kohlrausch Law, $\Lambda_{eq} (NaF) + \Lambda_{eq} (HCI) - \Lambda_{eq} (NaCI) = \Lambda_{eq} (HF)$
34. (b) Conductivity (X) = conductance (c) × cell
constant
 \therefore Cell constant $= \frac{X}{Y}$
Conductivity of NaOH $= \frac{X}{Y} \cdot Z$
Molar conductance of NaOH
 $= \frac{X}{Y} \cdot Z \times \frac{1000}{0.1} = \frac{XZ}{Y} \cdot 10^4$
39. (b)
35. (a) ArNO₂ 6H 6e⁻ \longrightarrow ArNH₂ 2H₂O
 $W = \frac{E.wt \times Q}{96500}$; 20.50 $= \frac{M}{6} \times \frac{2 \times 96500}{96500} \times \frac{50}{100}$
 \therefore M = 123.0 g
36. (c) Moles of NaOH formed $= \frac{600 \times 1}{100} = 0.6$
 $arrow = 1000 \text{ or geq of NaOH} = 0.6 \text{ geq of Cu deposited}$
 $arrow = \frac{31.8}{63.6/2} = 1.0$
 \therefore Current efficiency $= \frac{0.6 \times 100}{1} = 60\%$
37. (c) Total charge = 2
Number of equivalent of ion
 $= \frac{Charge on the ion}{Total charge}$
 \therefore Eq of $\binom{COO}{IOO^{-}} = \frac{2}{2} = 1$

Eq of Na⁺ =
$$\frac{1}{2}$$
, Eq of K⁺ = $\frac{1}{2}$
 $\therefore \lambda^{\circ}_{eq} \begin{pmatrix} COONa \\ | \\ COOK \end{pmatrix}$
= $\lambda^{\circ}_{eq} \begin{pmatrix} COO^{-} \\ | \\ COO^{-} \end{pmatrix} + \frac{1}{2}\lambda Na + \frac{1}{2}\lambda K$
= $74 + \frac{50}{2} + \frac{73}{2} = 135.5 \text{ ohm}^{-1} \text{ cm}^2 \text{ eq}^{-1}$
By Faraday's Ist law of electrolysis,
 $\frac{W}{E} = \frac{Q}{96500}$ (where Q it = charge of ion)
We know that no. of gram equivalent
 $\frac{W}{E} = \frac{\text{it}}{96500} \frac{1 \times 965}{96500} \frac{1}{100}$
(where i = 1 A, t = 16 \times 60 + 5 = 965 \text{ sec})

E 96500 96500 100 (where i = 1 A, $t = 16 \times 60 + 5 = 965$ sec.) Since, we know that Normality

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

E E
$$-\frac{0.0591}{2}\log\frac{1}{Cl^{-2}}$$
; as $Cl^{-}\downarrow, E\uparrow$

- (c) Corrosion of iron can be minimized by forming an impermeable barrier at its surface.
- **41.** (b) During charging, the lead storage battery behaves like an electrolytic cell. So, at anode the reaction is

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

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

- 43. (d) Reducing character decreases down the series. Hence the correct order is $Al < Fe^{2+} < Br^{-}$
- 44. (c) Cu \longrightarrow Cu⁺⁺ + 2e⁻ i.e, to deposit 1 mole of Cu at cathode from Cu²⁺ SO₄²⁻ solution = 2 moles of electrons are required i.e, To deposit

$$6.35g = \frac{6.35}{63.5} \times 2 \quad \frac{2}{10} \quad \frac{1}{5}$$
 moles

Thus total no. of electrons required = $\frac{N_A}{5}$

- 45. (a) Lower the value of reduction potential higher will be reducing power hence the correct order will be $Mn^{2+} < Cl^- < Cr^{3+} < Cr$
- **46.** (c) $E_{oxidation} = 0.059 \text{ pH} = 0.059 \times 10 = 0.59 \text{ V}$
- 47. (d) Standard Gibbs free energy is given as $\Delta G^{\circ} = -nE^{\circ}F$ If $E^{\circ}_{cell} < 0$ i.e. - ve $\Delta G^{\circ} > 0$ Further $\Delta G^{\circ} = -RT \ell n K_{eq}$
- **48. (b)** The oxidation potential

 $\frac{1}{\text{Concentration of ions}}$ and reduction potential concentration of ions. The cell voltage can be increased by decreasing the concentration of ions around anode or by increasing the concentration of ions around cathode

- **49.** (a) Oxidising tendency $\frac{1}{\text{Electrode potential}}$
 - $TX \longrightarrow No reaction$ $TY \longrightarrow X, Z$ $TZ \longrightarrow X$ $\Rightarrow order of electrode potential is$ TY < TZ < TX $\Rightarrow Order of oxidation of the anion is$ $Y^{-}>Z^{-}>X^{-}$
- 50. (c) Oxidation potential of M is more than Ni and less than Mn. Hence reducing power Mn > M > Ni

51. (c)
$$E_{Cr}^{\circ} + C_{Cr}^{3+} = -0.74 \text{ V}, E_{Co}^{\circ} + C_{Co}^{2+} = -0.28 \text{ V}$$

The given cell reaction is
 $Cr | Cr^{3} | 1.0 \text{ M} | | Co^{2} | (1.0 \text{ M}) | Co$
 \therefore Cr is anode and Co is cathode
 $E_{cell}^{\circ} = E_{C}^{\circ} - E_{A}^{\circ} = -0.28 - (-0.74)$
 $= -0.28 + 0.74 = +0.46 \text{ V}$
52. (a) $OCl^{-} \longrightarrow Cl^{-}; E^{\circ} = 0.94 \text{ V}(I)$
 $Cl^{-} \longrightarrow \frac{1}{2}Cl_{2} e^{-}; E^{\circ} = 1.36 \text{ V}(II)$
 $Add(I) + (II) OCl^{-} \longrightarrow \frac{1}{2}Cl_{2}; E^{\circ} = 0.94 - 1.36$
 $= -0.42 \text{ V}$
53. (a) $Cu + 2Ag^{+}(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$
Here, $n = 2$, $E_{cell}^{\circ} = +0.46 \text{ V}$

$$\Delta G^{\circ} = -nE^{\circ}F$$
$$= \frac{-2 \times 0.46 \times 96500}{1000} \text{ kJ} \simeq -89 \text{ kJ}$$

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

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

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

$$E^{\circ}_{\text{cell}} = \frac{2.303 \text{ RT}}{\text{nF}} \log K_{\text{eq}} - \frac{0.0591}{n} \log K_{\text{eq}}$$

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

or $\log K_{-} = \frac{2 \times 0.295}{10} - 10$

or
$$\log K_{eq} = \frac{2 \times 0.255}{0.0591}$$
 10
or $K_{eq} = 1 \times 10^{10}$

56. (b)
$$\Delta G = -nEF$$

For 1 mole of Al, $n = 3$
 \therefore for $\frac{4}{3}$ mole of Al, $n = 3 \times \frac{4}{3}$ 4
According to question,
 $-827 \times 1000 = -4 \times E \times 96500$
 $E \quad \frac{827 \times 1000}{4 \times 96500} = 2.14V$
 \therefore minimum e.m.f. required = 2.14V
57. (b) Cell reaction is, Zn Cu² \rightarrow Zn² Cu
 $E_{cell} = E_{cell}^{\circ} - \frac{RT}{nF} ln \frac{[Zn^2]}{[Cu^2]}$
Greater the factor $\left[\frac{(Zn^2)}{(Cu^2)}\right]$, less is the
EMF
Hence $E_1 > E_2$

58. (c) Correct Nernst equation is

$$E \quad E^{\circ} \quad \frac{2.303 RT}{nF} \log a_{\mathrm{M}^n}$$
 .

59. (b) For
$$\operatorname{Zn}^{2+} \longrightarrow \operatorname{Zn}$$

$$E_{\text{Zn}^2/\text{Zn}} = E_{\text{Zn}^2/\text{Zn}} - \frac{2.303\text{RT}}{nF} \log \frac{\text{Zn}}{[\text{Zn}^2]}$$

$$= -0.76 - \frac{0.06}{2} \log \frac{1}{0.1} = -0.76 - 0.03$$
$$E_{\text{Zn}^2} / \text{Zn} = -0.79\text{V}$$

60. (d) $2H^+ + 2e^- \longrightarrow H_2$ $E_{cell} = 0.06 \log [H^+] = 0.06 \times \log \sqrt{K_a C}$

$$=0.06 \times \log 10^{-3} = -0.18$$
 V

61. (b) W = Zitwhere Z = Electrochemical equivalentEq. wt. of copper = $\frac{63}{2}$ 31.5

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

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

62. (c) Reduction half reaction : $Cr_2O_7^{2-} 6e^- 14H \longrightarrow 2Cr^3 7H_2O$ Oxidation half reaction : $Cr \longrightarrow Cr^3 3e^-$ Overall reaction : $Cr_2O_7^{2-} Cr 14H 3e^- \longrightarrow 3Cr^3 7H_2O$

3F of electricity = 3 moles of
$$Cr^{3+}$$

1F of electricity = $\frac{3}{3}$ moles of Cr^{3+}

- 63. (c) More the reduction potential, more is the deposition of metals at cathode. Cation having E^0 value less than 0.83V (reduction potential of H₂O) will not deposit from aqueous solution. Hence correct order of A deposition of the metal at the cathode is Ag > Hg > Cu
- 64. (d) On pulling out one of the electrodes the electric field applied to the solution disappears and hence the ions start moving randomly.
- 65. (a) Cell reaction cathode :

$$H_2O(l) \quad \frac{1}{2}O_2(g) \quad 2e^- \longrightarrow 2OH^-(aq)$$

anode:
$$H_2(g) \longrightarrow 2H^+(aq) = 2e^{-1}$$

H₂O(1)
$$\frac{1}{2}$$
O₂(g) H₂(g) \longrightarrow 2H⁺(aq) 2OH⁻(aq)
Also we have

$$\begin{split} H_{2}(g) + \frac{1}{2}O_{2}(g) &\longrightarrow H_{2}O(l); \\ \Delta G_{f} &= -237.2 \text{ kJ/mole} \\ H_{2}O(l) &\longrightarrow H^{+}(aq) \quad OH^{-}(aq); \\ \Delta G & 80 \text{ kJ/mol} \\ \text{Hence for cell reaction} \\ \Delta G & -237.2 + (2 \times 80) = -77.20 \text{ kJ/mol} \end{split}$$

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

- **66.** (b) In concentration cell the spontaneous process is physical in nature involving transfer of matter from higher concentration to lower concentration in indirect manner.
- 67. (d) In the silver plating of copper, K[Ag(CN)₂] is used instead of AgNO₃. Copper being more electropositive readily precipitate silver from their salt solution

 $Cu + 2AgNO_3 \longrightarrow Cu(NO_3)_2 + Ag$ whereas in K[Ag(CN)₂] solution a complex anion [Ag(CN)₂]⁻ is formed and hence Ag⁺ are less available in the solution and therefore copper cannot displace Ag from its complex ion.

68. (c) Due to buffer action the pH will remain practically constant.

69. (a)
$$W_{\text{max}} = -n.FE;$$

l

$$W_{\rm max} = -2 \times 96500 \times 0.65 = -1.25 \times 10^5 \, {\rm J}$$

0.5g H₂ 0.25 mole

Hence W_{max}

$$=-1.25 \times 10^5 \times 0.25 -3.12 \times 10^4 \text{ J}$$

70. (a) For spontaneous $\Delta_r G^\circ = -nFE^\circ < 0$ $\therefore E^\circ > 0$

71. (c) When
$$Q = K_C$$
, $E = 0$

- 72. (c) Thermodynamic efficiency of a cell is given by the ratio of Gibbs function change to enthalpy change in the overall cell reaction.
 73. (b) Ionic molar conductivity of H⁺ is very high
 - **(b)** Ionic molar conductivity of H^+ is very high and NH_4OH is a weak electrolyte.
- 74. (a) Molar conductivity of H⁺ and OH⁻ are very high as compare to other ions. Initially conductance of solution sharply decreases due to consumption of free H⁺. After complete neutralization further slightly increases due to presence of OH⁻.

75. (c) According to the definition 1 F or 96500 C is the charge carried by 1 mol of electrons when water is electrolysed

> $2H_2O \longrightarrow 4H^+ + O_2 + 4e^-$ So, 4 Faraday of electricity liberate = 32 g of O₂. Thus 1 Faraday of electricity liberate

81.

$$=\frac{32}{4}$$
g of O₂ = 8 g of O₂

76. (a) Conductivity
$$(\kappa) = \frac{\ell}{R.a.}$$

 $\kappa = \frac{l}{R.a}$; cell constant. $\left(\frac{l}{a}\right) = 1.29 \times 100 = 129$
Again conductivity of 0.02M solution

Again conductivity of 0.02M solution

$$\kappa \quad \frac{\kappa \times 129}{520} \times 129$$
$$\Lambda_{\rm m} \quad \frac{\kappa \times 1000}{\rm M} \quad \frac{129}{520} \times \frac{1000}{0.02}$$

$$1.24 \times 10^{-4} \text{ S m}^2 \text{ mol}^{-1}$$

- 77. (c) $CICH_2COONa + HCl \rightarrow CICH_2COOH$ NaCl
 - $\Lambda_{\text{CICH}_{2}\text{COONa}} + \Lambda_{\text{HCI}} = \Lambda_{\text{CICH}_{2}\text{COOH}} + \Lambda_{\text{NaCI}}$ $224 + 203 = \Lambda_{\text{CICH}_{2}\text{COOH}} + 38.2$ $\Lambda_{\text{CICH}_{2}\text{COOH}} = 427 38.2$ $= 388.8 \text{ ohm}^{-1}\text{cm}^{2}\text{g eq}^{-1}$

78. (a) Ag
$$e^- \longrightarrow Ag_{1F} \longrightarrow 108g$$

1 F = 1 mole of electrons = 96500 C 0.01F = 1.08 g Ag; Ag left = 1.08 - 1.08 = 0 g

79. (a)
$$\Lambda_{\rm m} = \frac{\kappa \times 1000}{\rm M} = \frac{1.06 \times 10^{-2} \times 1000}{0.1}$$

 1.06×10^2

80. (d) The amount. of copper deposited at cathode by reduction of
$$Cu^2$$
 ions is $\frac{3.2}{63}$ 0.05 moles.

The same amount 0.05 mole of Cu^2 must pass into solution from anode by oxidation

(b)
$$\operatorname{Cu}^2 + e^- \longrightarrow \operatorname{Cu}$$
;
 $E_1^o = 0.15 \mathrm{V}; \Delta G_1^o - n_1 E_1^o F$
 $\operatorname{Cu}^2 + 2e \longrightarrow \operatorname{Cu};$
 $E_2^o = 0.34 \mathrm{V}; \Delta G_2^o - n_2 E_2^o F$
On subracting eq.(i) from eq. (ii) we get
 $\operatorname{Cu} + e^- \longrightarrow \operatorname{Cu}; \Delta G^o = \Delta G_2^o - \Delta G_1^o$
 $-n E^o F = -(n_2 E^o F - n_1 E_1^o F)$
 $E^o \frac{n_2 E_2^o F - n_1 E_1^o F}{nF}$
 $= \frac{2 \times 0.34 - 0.15}{1} = 0.53 \mathrm{V}$

82. (b)
$$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$$

$$E=1.51-\frac{0.059}{5}\log\frac{[Mn^2]}{[MnO_4^-][H]^8}$$

Taking Mn^{2+} and MnO_4^- in standard state i.e. 1 M,

$$E=1.51 - \frac{0.059}{5} \times 8\log \frac{1}{[H^+]}$$
$$=1.51 - \frac{0.059}{5} \times 8 \times 3 \quad 1.2268 \text{ V}$$

Hence at this pH, MnO_4^- will oxidise only Br⁻ and I⁻ as SRP of Cl₂/Cl⁻ is 1.36 V which is greater than that for MnO_4^-/Mn^{2+} .

83. (c) Higher the value of reduction potential stronger will be the oxidising hence based on the given values Ag⁺ will be strongest oxidizing agent.

$$\textbf{84. (b)} \quad \mathrm{E}_{\mathrm{Hg}^2 \ |\mathrm{Hg}} \quad \mathrm{E}_{\mathrm{Hg}^2 \ |\mathrm{Hg}} \quad \mathrm{E}_{\mathrm{Fg}^2 \ |\mathrm{Fe}} \quad \mathrm{E}_{\mathrm{Pb}^2 \ |\mathrm{Pb}}.$$

85. (b) E_{OP} of Zn E_{OP} of Fe.

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

Mass = Volume × density

$$\therefore$$
 Mass of Ag to be deposited

$$= \frac{80 \times 0.005}{10} \times 10.5 \quad 0.42 \text{ g}$$
Amount deposited = $\frac{i \times t \times E.wt}{96500}$
 $\therefore 0.42 = \frac{108 \times 3 \times t}{98500} \therefore t = 125.1 \text{ seconds}$
88. (a) $\therefore Q = i \times t$
 $\therefore Q = 4.0 \times 10^4 \times 6 \times 60 \times 60 \text{ C} = 8.64 \times 10^8 \text{ C}$
Now since 96500 C liberates 9 g of Al
 8.64×10^8 C liberates
 $\frac{9}{96500} \times 8.64 \times 10^8 \text{ g Al} = 8.1 \times 10^4 \text{ g of Al}$
89. (a) $\kappa \frac{\text{Cell constant}}{\text{R}} \frac{1.15}{250}$
 $\Lambda_{\text{eq}} \frac{1.15}{250} \times \frac{1000}{1} \quad 4.6$
90. (c) $\lambda_m \frac{1000\kappa}{0.1} \frac{1000 \times 3.75 \times 10^{-4}}{0.1} \quad 3.75;$
 $\alpha \frac{\lambda_m}{\lambda_m^{\infty}} \frac{3.75}{250} \quad 1.5 \times 10^{-2};$
 $K_a = C\alpha^2 = 0.1 \times (1.5 \times 10^{-2})^2 = 2.25 \times 10^{-5}$
91. (a) $\Lambda_{\text{eq}}^{\infty} (\text{NH}_4\text{ OH}) \quad \Lambda_{\text{eq}}^{\infty} (\text{NH}_4\text{ CI}) \quad \Lambda_{\text{eq}}^{\infty} (\text{NaOH})$
 $-\Lambda_{\text{eq}}^{\infty} (\text{NaCI})$
129.8 217.8-109.3 238.3 ohm^{-1} \text{ cm}^2 \text{ eq}^{-1}
 $\alpha \quad \frac{\Lambda_{\text{eq}}}{\Lambda_{\text{eq}}^{\infty}} \quad \frac{9.30}{238.3} \quad 0.04$
92. (b) The cell reaction is

87. (a) Volume = Area \times thickness

H₂ g I₂ s \rightleftharpoons 2H⁺ aq 2I⁻ aq 0.7714 = 0.535 - $\frac{0.0591}{2} \log \frac{[H^+]^2 [I^-]^2}{P_{H_2}}$ pH = 3

93. (c)
$$2Mn^2 \longrightarrow 2Mn^3 + 2e^-, \Delta G_1$$

 $\frac{Mn^{2+} + 2e^- \rightarrow Mn, \Delta G_2}{3Mn^2 \text{ aq } \rightarrow Mn \text{ s } 2Mn^3 \text{ aq}}$
 $-2 \times F \times E_3^\circ = -2 \times F \times [-1.51] - 2 \times F \times (-1.185)$
 $E_3 = -2.695$
 $E_3^\circ = \frac{0.0591}{2} \log K_{eq}; K_{eq} \approx 6.3 \times 10^{-92}$.

- 94. (b) In $H_2 O_2$ fuel cell, the combustion of H_2 occurs to create potential difference between the two electrodes
- **95.** (d) Higher the reduction potential, stronger is the oxidising agent.
- 96. (c) The given values show that Cr has highest negative value of $E^{\circ}_{M^{3+}/M^{2+}}$ i.e., Cr is the strongest reducing agent among given metals therefore its oxidation will be easiest.

97. (b)
$$(E_{Cu^2 | Cu})_1 = E_{Cu^2 | Cu}^0 = \frac{0.059}{2}$$

 $\log[Cu^{2^+}]_1$

$$[Cu^{2+}(aq) + 2e^{-} \longrightarrow Cu(s)] \text{ and}$$

$$(E_{Cu^{2}}|_{Cu})_{2} = E_{Cu^{2}}^{0}|_{Cu} = \frac{0.059}{2} \log [Cu^{2}]_{2}$$

$$\Rightarrow \Delta E_{Cu^{2}}|_{Cu} = -\frac{0.059}{2} \log \left(\frac{C_{2}}{C_{1}}\right)$$

$$= -\frac{0.059}{2} \log 10 = -0.030 \text{ V} = -30 \text{ mV}$$

$$= -\frac{0.059}{2} \log 10 = -0.030 \text{ V} = -30 \text{ mV}$$

98. (a)
$$E_{red} = E_{red} = \frac{0.059}{n} \log \frac{[M^{(x-1)}]}{[M^x]}$$

99. (b) The cell reaction is

$$Cl_2(g)(P_2 \text{ atm}) \longrightarrow Cl_2(aq)(P_1 \text{ atm})$$

 $E_{cell} = -\frac{0.0592}{2} \log \frac{P_1}{P_2} - \frac{0.0592}{2} \log \frac{P_2}{P_1}$
 E_{cell} will be positive when $P_2 - P_1$.

100. (b)
$$\Lambda_{eq} = \frac{\kappa \times 1000}{N}$$