Q1: NTA Test 01 (Single Choice)

Consider the following cell reaction

$$2Fe\,(s) + O_2\,(g) + 4H^+\,(aq) \rightarrow 2Fe^{2+}\,(aq) + 2H_2O\,(l)$$
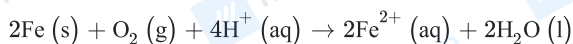
If $E_{cell} = E^o_{cell}$ at 25^oC and $\left[Fe^{2+}\right] = 10^{-3}M$, $P_{O_2} = 0.01$ atm and pH $= x$

Value of x is

(A) 1   (B) 2

(C) 3   (D) 4

Q2: NTA Test 02 (Single Choice)

The standard reduction potential for Cu^{2+}/Cu is $+\,0.34$ V. Calculate the reduction potential at pH $= 14$ for the above couple.

$$\left(K_{sp} \text{ of } Cu(OH)_2 = 1 \times 10^{-19}\right)$$

(A) -0.22 V   (B) $+0.22$ V

(C) -0.44 V   (D) $+0.44$ V

Q3: NTA Test 03 (Single Choice)

The equivalent conductivity of 0.1 M weak acid is 100 times lesser than that at infinite dilution. The degree of dissociation of weak electrolyte at 0.1 M is -

(A) 100   (B) 10

(C) 0.01   (D) 0.001

Q4: NTA Test 04 (Single Choice)

For sparingly soluble salt $\left[M\left(NH_3\right)_4 Br_2\right]H_2PO_2$, what will be the solubility product constant in $\left(\frac{mol}{litre}\right)^2$?

$$\left[\text{Given, } \lambda^0_{M(NH_3)_4Br_2^+} = 100 \text{ S m}^2 \text{ mol}^{-1}, \lambda^0_{H_2PO_2^-} = 50 \text{ S m}^2 \text{ mol}^{-1}\right]$$
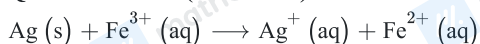
Take specific resistance as $200 \, \Omega \, cm$

(A) 1.11×10^{-11} (B) 1.11×10^{-3}

(C) 3.33×10^{-6} (D) None of these

Q5: NTA Test 05 (Single Choice)

The pressure of H_2 required to make the potential of hydrogen electrode zero in pure water at 298 K is:

(A) 10^{-14} atm   (B) 10^{-12} atm

(C) 10^{-10} atm   (D) 10^{-4} atm

Q6: NTA Test 06 (Numerical)

$$Ag\,(s) + Fe^{3+}\,(aq) \longrightarrow Ag^+\,(aq) + Fe^{2+}\,(aq)$$

Given standard electrode potentials- $E^o_{Fe^{3+}/Fe^{2+}} = +0.77$ V and $E^o_{Ag^+/Ag(s)} = +0.80$ V

If the reaction is feasible, enter 1.00 as answer elsewise enter 0.00.

Q7: NTA Test 07 (Single Choice)

Calculate ΔG^o for the following cell reaction

$$Zn\,(s) + Ag_2O\,(s) + H_2O\,(l) \rightarrow Zn^{2+}\,(aq) + 2Ag\,(s) + 2OH^-\,(aq)$$
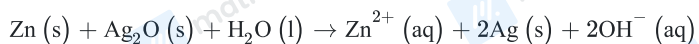
$$E^0_{Ag^+/Ag} = +0.80V \text{ and } E^0_{Zn^{+2}/Zn} = -0.76V, F = 96500$$

(Given: K_{sp} of $AgOH = 2 \times 10^{-8}$)

(A) $-305kJ/mol$ (B) $212\ kJ/mol$

(C) $305\ kJ/mol$ (D) $301\ kJ/mol$

Q8: NTA Test 08 (Single Choice)

The resistance of 0 .01 N solution of an electrolyte was found to be 220 ohm at 298 K using a conductivity cell with a cell constant of $0.88\ cm^{-1}$. The value of equivalent conductance of solution is-
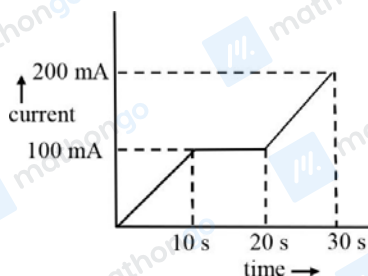
(A) 400 mho cm^2 g eq^{-1} (B) 295 mho cm^2 g eq^{-1}

(C) 419 mho cm^2 g eq^{-1} (D) 425 mho cm^2 g eq^{-1}

Q9: NTA Test 09 (Single Choice)

In a Cu-voltameter, mass deposited in 30 s is 'm' g. If the time-current graph is shown in the following figure:

What is the electrochemical equivalent of Cu?

(A) $m/2$ (B) $m/3$

(C) $m/4$ (D) $\frac{m}{63.5}$

Q10: NTA Test 12 (Numerical)

For the electrochemical cell, $Mg(s)\left|Mg^{2+}\,(aq, 1\ M)\right|\left|Cu^{2+}\,(aq, 1\ M)\right|Cu\,(s)$ the standard emf of the cell is 2.70 V at 300 K. When the concentration of Mg^{2+} is changed to x M, the cell potential changes to 2.67 V at 300 K. The value of x is _____.

(given, $\frac{F}{R} = 11500\ K\ V^{-1}$, where F is the Faraday constant and R is the gas constants, $\ln\,(10) = 2.30$)

Q11: NTA Test 13 (Numerical)

EMF of the following cell is 0.6 volt.

$$Ag\,(s)\,|AgBr\,(s)|KBr\,(0.01M)||AgNO_3\,(0.001\ M)|Ag(s)$$

K_{sp} of AgBr is expressed as 1×10^{-x}, x is

$\left[\text{Take}\frac{2.303RT}{F} = 0.06\ V\right]$

Q12: NTA Test 14 (Single Choice)

A current of 1.93 ampere is passed through 200 mL of 0.5 M Zinc sulphate (aq.) solution for 50 min with a current efficiency of 80%. If volume of solution remain constant, then $\left[Zn^{2+}\right]$ after deposition of Zn^{2+} is:

(A) 0.38 M   (B) 0.26 M

(C) 0.35 M   (D) 0.076 M

Q13: NTA Test 15 (Single Choice)

K_{sp} of $Al(OH)_3 = 10^{-36}$

and $E^o_{Al^{3+}/Al} = -1.66$ V

Reduction potential of Al^{3+}/Al couple at pH = 12 and 298 K is

(A) 1.07 V   (B) 2.25 V

(C) -1.07 V   (D) -2.25 V

Q14: NTA Test 16 (Single Choice)

Calculate the time required for complete decomposition of one mole of water using 2 ampere current. .

(A) 9650 sec   (B) 96500 sec

(C) 1.93×10^5 sec   (D) 38600 sec

Q15: NTA Test 17 (Numerical)

Four different solutions containing 1 M each of $Au^{+3}, Cu^{+2}, Ag^+, Li^+$ are being electrolysed by using inert electrodes. In how many samples, metal ions would be deposited at cathode?

[Given: $E^0_{Ag^+/Ag} = 0.8V, E^0_{Au^{+3}/Au} = 1.00V$, $E^0_{Cu^{+2}/Cu} = 0.34V, E^0_{Li^+/Li} = -3.03V$]

Q16: NTA Test 18 (Single Choice)

A current strength of 0.965 amperes is passed through excess fused $AlCl_3$ for 5 hours. How many litres of chlorine will be liberated at STP ?

$(F = 96500 \text{ C})$

(A) 2.016   (B) 1.008

(C) 11.2   (D) 20.16

Q17: NTA Test 19 (Numerical)

The cell, $Zn/Zn^{2+}(1M)||Cu^{2+}(1M)/Cu(E^0_{cell} = 1.10V)$ was allowed to be completely discharged at 298K. The relative concentration of Zn^{2+} to Cu^{2+} $\left(\frac{[Zn^{2+}]}{[Cu^{2+}]}\right)$ is 10^x . The value of x is:

(Take $\frac{2.303RT}{F} = 0.059$ Round off your answer up to one decimal)

Q18: NTA Test 20 (Numerical)

Consider an electrochemical cell: $A(s) \left| A^{n+} (aq, 2 \text{ M}) \right| \left| B^{2n+} (aq, 1 \text{ M}) \right| B(s)$. The value of ΔH^o for the cell reaction is twice that of ΔG^o at 300 K. If the emf of the cell is zero, the $\Delta S^\ominus \left(\text{in } JK - 1 \text{ mol}^{-1} \right)$ of the cell reaction per mole of B formed at 300 K is ____

(Given: $\ln(2) = 0.7, R = 8.3$ J K^{-1} mol^{-1}. H, S and G are enthalpy, entropy and Gibbs energy, respectively.)

Q19: NTA Test 21 (Numerical)

The molar conductivity of aqueous solution of a weak acid HA is 200 $Scm^2 mol^{-1}$ and its pH is 4. What is the value of pK_a of HA at 25^oC ?

(Given: $\lambda^o_m(NaA) = 100$ $Scm^2 mol^{-1}$, $\lambda^o_m(HCl) = 425$ $Scm^2 mol^{-1}$, $\lambda^o_m(NaCl) = (125$ $Scm^2 mol^{-1})\lambda^o_m(NaA) = 100$ $Scm^2 mol^{-1}$, λ^o_m)

     $(HCl) = 425$ $Scm^2 mol^{-1}$, $\lambda^o_m(NaCl) = 125$ $Scm^2 mol^{-1})$

Q20: NTA Test 22 (Single Choice)

The resistance of conductivity cell containing 0.001 M KCl solution at 298 K is 1500 Ω. What is the cell constant if conductivity of 0.001 M KCl solution at 298 K is 0.146×10^{-3} Scm^{-1}.

(A) 0.200 cm^{-1} (B) 0.219 cm^{-1}

(C) 0.195 cm^{-1} (D) 0.190 cm^{-1}

Q21: NTA Test 24 (Single Choice)

The temperature coefficient of the emf i.e., $\frac{dE}{dt} = -0.00065\ VK^{-1}$ for the cell $Cd|CdCl_2\ (1M)||AgCl\ (s)|Ag$ at 25^oC. Calculate the entropy changes ΔS 298 K for the cell reaction $Cd + 2\ AgCl \rightarrow Cd^{2+} + 2\ Cl^- + 2\ Ag$.

(A) $-105.5\ JK^{-1}$ (B) $-150.2\ JK^{-1}$

(C) $-75.7\ JK^{-1}$ (D) $-125.5\ JK^{-1}$

Q22: NTA Test 25 (Single Choice)

For an electrochemical cell, $Sn\ (s)\left|Sn^{2+}\ (aq,\ 1\ M)\right|\left|Pb^{2+}\ (aq,\ 1\ M)\right|Pb\ (s)$, the ratio $\frac{\left[Sn^{2+}\right]}{\left[Pb^{2+}\right]}$ when this cell attains equilibrium is

$\left(\text{Given: } E^0_{Sn^{2+}|Sn} = -0.14V, E^0_{Pb^{2+}|Pb} = -0.13V, \frac{2.303\ RT}{F} = 0.06\right)$

(A) 4.3 (B) 1

(C) -2.15 (D) 2.14

Q23: NTA Test 26 (Numerical)

pH of the anodic solution of the following cell is

$Pt,\ H_2\ (1\ atm)\left|H^+\ (xM)\right|\left|H^+\ (1\ M)\right|H_2(1\ atm),\ Pt$

if E_{cell}= 0.2364 V.

Q24: NTA Test 27 (Single Choice)

Given

$E^o_{Cl_2/Cl^-} = 1.36\ V,\ E^o_{Cr^{3+}/Cr} = -0.74\ V$

$E^o_{Cr_2O_7^{2-}/Cr^{3+}} = 1.33\ V,\ E^o_{MnO_4^-/Mn^{2+}} = 1.51\ V.$

Among the following, the strongest reducing agent is:

(A) Mn^{2+} (B) Cr^{3+}

(C) Cl^- (D) Cr

Q25: NTA Test 28 (Single Choice)

Metals which will not evolve H_2 gas with dil. HCl are

(A) Cu, Ag, Au (B) Cu, Zn, Al

(C) Fe, Ag, Pt (D) Hg, Mg, Pt

Q26: NTA Test 29 (Single Choice)

Reduction potentials of four elements P, Q, R, S is -2.90 V, 0.34 V, 1.2 V and -0.76 V. The decreasing order of reducing power is

(A) $P > Q > R > S$ (B) $S > R > Q > P$
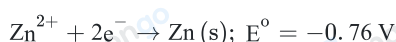
(C) $P > S > Q > R$ (D) $Q > S > R > P$

Q27: NTA Test 30 (Single Choice)

The standard electrode potentials, E^0 of Fe^{3+}/Fe^{2+} and Fe^{2+}/Fe at 300 K are +0.77 V and -0.44 V, respectively. The E^0 of Fe^{3+}/Fe at the same temperature is

(A) 1.21 V (B) 0.33 V

(C) -0.036 V (D) 0.036

Q28: NTA Test 31 (Single Choice)

Consider the following reduction processes

$Zn^{2+} + 2e^- \rightarrow Zn\ (s);\ E^o = -0.76\ V$

$Ca^{2+} + 2e^- \rightarrow Ca\,(s);\ E^o = -2.87\,V$

$Mg^{2+} + 2e^- \rightarrow Mg\,(s);\ E^o = -2.36\,V$

$Ni^{2+} + 2e^- \rightarrow Ni\,(s);\ E^o = -0.25\,V$
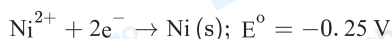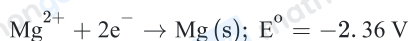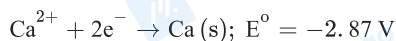
The reduction power of the metals increases in the order

(A) $Ca < Zn < Mg < Ni$

(B) $Ni < Zn < Mg < Ca$

(C) $Zn > Mg < Ni < Ca$

(D) $Ca < Mg < Zn < Ni$

Q29: NTA Test 32 (Numerical)
Copper sulphate solution (250 mL) was electrolysed using a platinum anode and a copper cathode. A constant current of 2 mA was passed for 16 min. It was found that after electrolysis the absorbance of the solution was reduced to 50% of its original value. Concentration of copper sulphate in the solution to begin with is $x \times 10^{-5}$. Find the valu of x

Q30: NTA Test 33 (Single Choice)
During an electrolysis of conc H_2SO_4, peroxydisulphuric acid $(H_2S_2O_8)$ and O_2 form in an equimolar amount. The moles of H_2 that will be formed simultaneously will be

(A) Thrice that of O_2

(B) Twice that of O_2

(C) Equal to that of O_2

(D) Half of the of O_2

Q31: NTA Test 34 (Numerical)
A current of 5.0 A flows for 4.0 h through an electrolytic cell containing a molten salt of metal M. This results in deposition of 0.25 mol of the metal M at the cathode. The oxidation state of M in the molten salt is + x. the value of 'x' is (1 Faraday = 96000 C mol^{-1})

Q32: NTA Test 35 (Numerical)
A cell contains two hydrogen electrodes. The negative electrode is in contact with a solution of 10^{-6} M hydrogen ion. The EMF of the cell is 0.118 V at 298 K. The concentration of H^+ ion at the positive electrode is 10^{-x} ,The value of 'x' is

Q33: NTA Test 36 (Single Choice)
The emf of the cell $Zn|Zn^{2+}\,(0.01M)||Fe^{2+}\,(0.001M)|$ Fe at 298 K is 0.2905 volt. Then the value of equilibrium constant for the cell reaction is

(A) $e^{0.32/0.0295}$

(B) $10^{0.32/0.0295}$

(C) $10^{0.26/0.0295}$

(D) $10^{0.32/0.0591}$

Q34: NTA Test 38 (Numerical)
Given that $\left(\text{in Scm}^2\ eq^{-1}\right)$ at $T = 298\,K$: Λ_{eq}^o for $Ba\,(OH_2)$, $BaCl_2$ and NH_4Cl are 228.8, 120.3 and 129.8 respectively. Specific conductance for 0.2 N NH_4OH solution is $4.766 \times 10^{-4}S\ cm^{-1}$, then value of pH of the given solution of NH_4OH will be nearly. (Take $\log 2 = 0.3$)

Q35: NTA Test 39 (Numerical)
Resistance of a 0.1M KCl solution in a conductance cell is 300 ohm and specific conductance of 0.1M KCl is $1.33 \times 10^{-2}ohm^{-1}cm^{-1}$. The resistance of 0.1M NaCl solution in the same cell is 400 ohm. The equivalent conductance of the 0.1M NaCl solution (in $ohm^{-1}cm^2/gmeq.$) is

Q36: NTA Test 40 (Single Choice)
$AgNO_3\,(aq)$ was added to an aqueous KCl solution gradually and the conductivity of the solution was measured. The plot of conductivity (Λ) versus the volume of $AgNO_3$ is
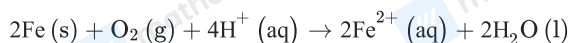
(A) (P)   (B) (Q)
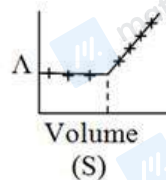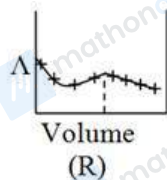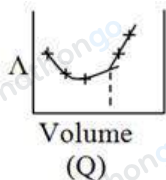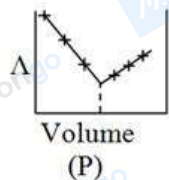
(C) (R)   (D) (S)

Q37: NTA Test 41 (Single Choice)

Consider the following cell reaction

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

$$E^o = 1.67\,V$$

At $\left[Fe^{2+}\right] = 10^{-3}\,M, P\,(O_2) = 0.1$ atm and pH $= 3$, the cell potential at 25^oC is

(A) 1.77 V   (B) 1.57 V

(C) 1.87 V   (D) 1.47 V

Q38: NTA Test 42 (Numerical)

$I_2\,(s)/I^-$ (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 $E^o_{I_2/I^-} = 0.535V$, find the pH of H^+/H_2 half cell.

Q39: NTA Test 43 (Numerical)

For the strong electrolytes NaOH, NaCl and $BaCl_2$ the molar ionic conductivities at infinite dilution are 250, 125 and 300 mho cm^2mol^{-1} respectively. The molar conductivity of $Ba(OH)_2$ at infinite dilution (mho cm^2mol^{-1}) is.

Q40: NTA Test 44 (Numerical)

Calculate the emf of the cell in which the following reaction takes place :

$$Ni(s) + 2\,Ag^+\,(0.002\,M) \longrightarrow Ni^{2+}\,(0.160\,M) + 2\,Ag(s)$$

Given that $E^o_{cell} = 1.05\,V$

Q41: NTA Test 45 (Single Choice)

If $E^0_{ClO_3^-/ClO_4^-} = -0.36$ V & $E^0_{ClO_3^-/ClO_2^-} = 0.33$ V at 300 K. The equilibrium concentration of perchlorate ion $\left(ClO_4^-\right)$ which was initially 1.0 M in ClO_3^- when the reaction starts to attain the equilibrium,

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

Given : Anti log(0.509) = 3.329

(A) 0.0236 M   (B) 0.0190 M

(C) 0.123 M   (D) 0.191 M

Q42: NTA Test 46 (Single Choice)

For the following electrochemical cell at 298 K,

$$Pt\,(s)\,|H_2\,(g, 1\,bar)|H^+\,(aq, 1M)|\,|M^{4+}\,(aq), M^{2+}\,(aq)|Pt\,(s)$$

$E_{cell} = 0.092V$ when $\frac{[M^{2+}(aq)]}{[M^{4+}(aq)]} = 10^x$

Given: $E^0_{M^{4+}|M^{2+}} = 0.151V$; $2.303\frac{RT}{F} = 0.059$ V

The value of x is

(A) -2 (B) -1

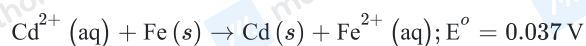(C) 1   (D) 2

Q43: NTA Test 46 (Single Choice)

Which substance is obtained in the solution on electrolysis of aq. $CuSO_4$ solution using graphite electrodes?

(A) $Cu(OH)_2$ (B) Na_2SO_4

(C) H_2SO_4 (D) H_2O

Q44: NTA Test 47 (Numerical)

A 1.0 M solution of Cd^{2+} is added to excess iron and the system is allowed to reach equilibrium. What is the concentration in mol of Cd^{2+}?

$$Cd^{2+}(aq) + Fe(s) \rightarrow Cd(s) + Fe^{2+}(aq); E^o = 0.037 \text{ V}$$

Given: $\log 18 = 1.25$

Report your answer upto two decimal places.

Q45: NTA Test 48 (Numerical)

0.02 equivalent of Ag was deposited in an electrolysis experiment. If same quantity of a electricity is passed through a gold solution, 1.314 g of gold is deposited. Find oxidation state of the gold. (Atomic mass of Au = 197)

Answer Keys

Q1: (A)	**Q2:** (A)	**Q3:** (C)
Q4: (A)	**Q5:** (A)	**Q6:** 0.00
Q7: (B)	**Q8:** (A)	**Q9:** (B)
Q10: 10	**Q11:** 15.00	**Q12:** (A)
Q13: (D)	**Q14:** (B)	**Q15:** 3
Q16: (D)	**Q17:** 37	**Q18:** -11.60
Q19: 4	**Q20:** (B)	**Q21:** (D)
Q22: (D)	**Q23:** 4	**Q24:** (D)
Q25: (A)	**Q26:** (C)	**Q27:** (C)
Q28: (B)	**Q29:** 8	**Q30:** (A)
Q31: 3	**Q32:** 4	**Q33:** (B)
Q34: 11.3	**Q35:** 100	**Q36:** (D)
Q37: (B)	**Q38:** 3	**Q39:** 50
Q40: 0.91	**Q41:** (D)	**Q42:** (D)
Q43: (C)	**Q44:** 0.05	**Q45:** 3

Solutions

Q1: (A) 1

For $E_{cell} = E_{cell}^{o}$

$\log Q$ must be zero

$Q = 1$

$Q = \dfrac{[Fe^{2+}]^2}{[H^+]^4 [P_{O_2}]} \Rightarrow \dfrac{10^{-6}}{(10^{-4x})(0.01)} = 1$
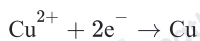
$10^{-4} = 10^{-4x}$

$x = 1$

Q2: (A) -0.22 V

When $pH = 14$, $\left[H^+\right] = 10^{-14}$ and $\left[OH^-\right] = 1$ M

$K_{sp} = \left[Cu^{2+}\right]\left[OH^-\right]^2 = 10^{-19}$

$\therefore \left[Cu^{2+}\right] = \dfrac{10^{-19}}{\left[OH^-\right]^2} = 10^{-19}$

The half cell reaction

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

$E = E^O - \dfrac{0.059}{2}\log \dfrac{1}{\left[Cu^{2+}\right]}$

$= 0.34 - \dfrac{0.059}{2}\log \dfrac{1}{10^{-19}} = -0.22$ V

Q3: (C) 0.01

The equivalent conductivity of 0.1 M weak acid is 100 times lesser than that at infinite dilution.
$\wedge_v = \dfrac{\wedge^0}{100}$
$\therefore$ Degree of dissociation $(\alpha) = \dfrac{\wedge_v}{\wedge^0} = \dfrac{\wedge^0}{100\wedge^0} = 0.01$

Q4: (A) 1.11×10^{-11}

Saturated solution corresponding to infinite dilution

$\wedge_m^o$ for salt $= \lambda^\circ_{M\left(NH_3\right)_4 Br_2^+} + \lambda^\circ_{H_2PO_2^-} = 100 \text{ S m}^2 \text{ mole}^{-1} + 50 \text{ S m}^2 \text{ mol}^{-1} = 150 \text{ S m}^2 \text{ mol}^{-1}$

$\kappa = \dfrac{1}{\rho} = \dfrac{1}{200}\Omega^{-1} \text{ cm}^{-1} = \dfrac{1}{2}\Omega^{-1}\text{m}^{-1}$

$\wedge_m^o \left(salt\right) = \dfrac{\kappa}{s} = \dfrac{\frac{1}{2}\Omega^{-1}\text{m}^{-1}}{s \text{ moles } \text{m}^{-3}}$

$$s\left(\text{moles m}^{-3}\right) = \frac{1}{2 \times 150} = \frac{1}{300} = 3 \cdot 33 \times 10^{-3}$$

$$s\left(\text{moles L}^{-1}\right) = 3 \cdot 33 \times 10^{-6}$$

$$K_{sp} = \left(3 \cdot 33 \times 10^{-6}\right)^2$$

$$= 1.11 \times 10^{-11} \text{ M}^2$$

Q5: (A) 10^{-14} atm

$\text{Pt}, \frac{H_{2(g)}}{H^+}$ Hydrogen electrode $E^o_{H_2/H^+} = 0.0$ Volt

$$H_{2(g)} \rightarrow 2H^+_{(aq)} + 2e^-$$

$\left[H^+\right] = 10^{-7}$ M at 25^oC (for Pure water)

$$E = \frac{-0.0591}{2}\log\left(\frac{[H^+]^2}{P_{H_2}}\right)$$

$$E = 0 = \log\frac{[H^+]}{P_{H_2}} = 0$$
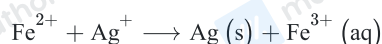
$$\therefore \left[H^+\right]^2 = P_{H_2} = \left(10^{-7}\right)^2$$

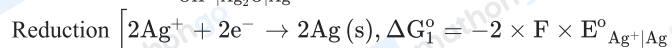$$\therefore P_{H_2} = 10^{-14} \text{ atm}$$

Q6: 0.00

$$\text{Ag}\,(s) + Fe^{3+}\,(aq) \longrightarrow Ag^+\,(aq) + Fe^{2+}\,(aq)$$
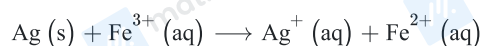
$$E^o_{cell} = E^o_{\left(Fe^{3+}/Fe^{2+}\right)} - E^o_{\left(Ag^+/Ag\right)} = 0.77 - 0.80 = -0.03 \text{ V}$$

The reaction is not feasible.

$$Fe^{2+} + Ag^+ \longrightarrow Ag\,(s) + Fe^{3+}\,(aq)$$

Q7: (B) 212 kJ/mol
Calculation for $E^-_{OH^-|Ag_2O|Ag}$
Reduction $\left[2Ag^+ + 2e^- \rightarrow 2Ag\,(s), \Delta G^o_1 = -2 \times F \times E^o_{Ag^+|Ag}\right.$
Ionisation $\left[Ag_2O + H_2O\,(l) \rightarrow 2Ag^+ + 2OH^- \;\Delta G^o_2 = -2.303RT\log k_{sp^2}\right.$
Net $Ag_2O + H_2O\,(l) + 2e^- \rightarrow 2Ag\,(s) + 2OH^-,\; \Delta G^o_3 = -2 \times F \times E^o_{OH^-|Ag_2O|Ag}$
$\Delta G^o_3 = \Delta G^o_1 + \Delta G^o_2$
$+2 \times F \times E^o_{OH^-|Ag_2O|Ag} = +2 \times F \times E^o_{Ag^+|Ag} + \frac{2.303RT\log k_{sp^2}}{2F}$
$E^o_{OH^-|Ag_2O|Ag} = E^o_{Ag^+|Ag} + \frac{0.06}{2} \times 2\log k_{sp}$
$= 0.8 + 0.06\log\left(2 \times 10^{-8}\right)$
$= 0.8 + 0.06\left[-8 + 0.3\right]$
$= 0.8 + 0.06 \times (-7.7)$
$= 0.8 - 0.462$
$= 0.338 \text{Volt} \cong 0.34\text{V} \;\Delta G^o = -nFE^o_{ceu}$
$= -nF(E^o_c - E_A)_{Red}$
$= -2 \times \left(E^o_{OH^-|Ag_2O|Ag} - E^o_{Zn^{+2}|Zn}\right)$
$= -2 \times 96500 \times (0.34 + 0.76)$
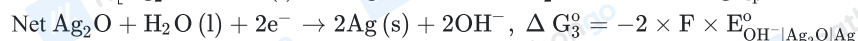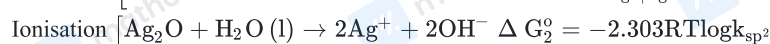
$= -2 \times 96500 \times 1.1 \text{volt}$

$= 212300 \text{J}$

$= 212.3 \text{ kJ}$

Q8: (A) $400 \text{ mho cm}^2 \text{ g eq}^{-1}$

$\Lambda_{eq} = k \times \frac{100}{N} = \frac{1}{R} \times \frac{1}{a} \times \frac{1000}{N}$

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

$= \frac{1}{220} \times 0.88 \times \frac{1000}{0.01}$
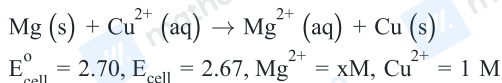
$= 400 \text{ mho cm}^2 \text{g eq}^{-1}$

Q9: (B) $m/3$

Total current passed in 30 sec $=$ Area of current$-$time curve

$= \left(\frac{1}{2} \times 100 \times 10\right) + (100 \times 10) + (100 \times 10) + \left(\frac{1}{2} \times 100 \times 10\right)$

$= 3 \times 100 \times 10 \text{ mA} \times \text{s}$

$= 3000 \times 10^{-3} \text{ amp s} = 3 \text{ C}$

$W = \frac{EQ}{F}$ or $\frac{W}{Q} = \frac{E}{F} = Z = \frac{m}{3}$

Q10: 10

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

$E_{cell}^{o} = 2.70, E_{cell} = 2.67, Mg^{2+} = xM, Cu^{2+} = 1 \text{ M}$

$E_{cell} = E_{cell}^{o} - \frac{RT}{nF} \ln x$

$2.67 = 2.70 - \frac{RT}{2F} \ln x$
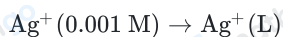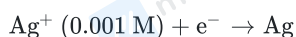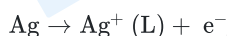
$-0.03 = - \frac{R \times 300}{2F} \times \ln x$

$\ln x = \frac{0.03 \times 2}{300} \times \frac{F}{R}$

$= \frac{0.03 \times 2 \times 11500}{300 \times 1}$

$\ln x = 2.30 = \ln (10)$

$x = 10$

Q11: 15.00

$Ag \rightarrow Ag^+\ (L) + e^-$

$Ag^+\ (0.001 \text{ M}) + e^- \rightarrow Ag$

$Ag^+ (0.001 \text{ M}) \rightarrow Ag^+ (L)$

$0.6 = 0 - \frac{0.06}{1} \log \frac{K_{sp}}{10^{-2} \times 10^{-3}}$

$10^{-10} = \frac{K_{sp}}{10^{-5}} \Rightarrow K_{sp} = 10^{-15}$

Q12: (A) 0.38 M

No. of Faraday $= \frac{1.93 \times 50 \times 60}{96500} = 0.06$

$\therefore$ Moles of Zn^{2+} deposited $= \frac{80}{100} \times \frac{0.06}{2} = 0.024$

$\therefore \left[Zn^{2+} \right] = \frac{0.5 \times 0.2 - 0.024}{0.2} = 0.38$ M

Q13: (D) -2.25 V

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

$E = E^o - \frac{0.059}{3} \log \frac{1}{[Al^{3+}]}$

at pH $= 12 \Rightarrow$ pOH $= 14 - 12 = 2 \Rightarrow \left[OH^- \right] = 0.01$ M

$\left[Al^{+3} \right] = \frac{10^{-36}}{[OH^-]^3} = \frac{10^{-36}}{(0.01)^3} = 10^{-30}$

$\therefore E = -1.66 - \frac{0.059}{3} \log \left(\frac{1}{10^{-30}} \right)$

$= -1.66 - \frac{0.059}{3} \times 30 = -1.66 - 0.59 = -2.25$ V

Q14: (B) 96500 sec

$2H_2O \rightarrow 2H_2 + O_2$
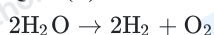
One mole of H_2O need 2 moles of e^- s for electrolysis.

So, Q $= I \times t$

$\Rightarrow t = \frac{Q}{I}$

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

$\therefore t = 96500$ sec

Q15: 3

Ag^+, Au^{+3}, Cu^{+2}

SRP of H_2O is more than SRP of lithium.

$\therefore$ H_2O will be reduced at cathode in case of Li^+ aq. Solution

Q16: (D) 20.16

1 F $\rightarrow$ 11.2 L Cl_2 at STP

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

$\therefore$ Volume of Cl_2 at STP $= 1.8 \times 11.2$ L

$= 20.16$ L

Q17: 37

$E_{cell} = 0$; when cell is completely discharged

$E_{cell} = E_{cell}^0 - \frac{0.059}{2} \log \left[\frac{[Zn^{2+}]}{[Cu^{2+}]} \right]$

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

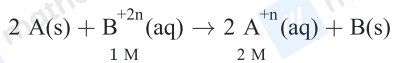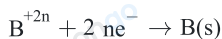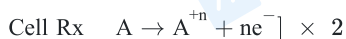$\log \left[\frac{[Zn^{2+}]}{[Cu^{2+}]} \right] = \frac{2 \times 1.1}{0.059} = 37.3$

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

$\therefore x = 37.3$

Q18: -11.60

$A(s) \mid A^{+n}(aq.\ 2\ M) \parallel B^{+2n}(aq.\ 1\ M) \mid B(s)$

$\Delta H^o = 2\Delta G^o E_{cell} = 0$

Cell Rx $\quad A \to A^{+n} + ne^-] \ \times \ 2$

$B^{+2n} + 2\ ne^- \to B(s)$

$2\ A(s) + \underset{1\ M}{B^{+2n}}(aq) \to 2\ \underset{2\ M}{A^{+n}}(aq) + B(s)$

$\Delta G = \Delta G^o + RT\ \ln \dfrac{\left[A^{+n}\right]^2}{\left[B^{+2n}\right]}$

$\Delta G^o = -RT\ \ln \dfrac{\left[A^{+n}\right]^2}{\left[B^{+2n}\right]} = -RT \times \ln \dfrac{2^2}{1} = -RT\ln 4$

$\Delta G^o = \Delta H^o - T\Delta S^o$

$\Delta G^o = 2\Delta G^o - T\Delta S^o$

$\Delta S^o = \dfrac{\Delta G^o}{T} = -\dfrac{RT\ \ln 4}{T}$

$= -8.3 \times 2 \times 0.7 = -11.6\ \text{J/K mol}^{-1}$

Q19: 4

$\wedge_M^o (HA) = \wedge_M^\infty (HCl) + \wedge_M^\infty (NaA) - \wedge_M^\infty (NaCl)$

$= 425 + 100 - 125 = 400\ Scm^2 mol^{-1}$

$pH = 4,\ [H^+] = 10^{-4} = \alpha C$

$\alpha = \dfrac{\wedge m}{\wedge_m^\infty} = \dfrac{200}{400} = 0.5$

$K_a = \dfrac{(C\alpha)\alpha}{(1-\alpha)} = \dfrac{10^{-4}(0.5)}{(1-\alpha)} = 10^{-4},\ pK_a = 4$

Q20: (B) $0.219\ cm^{-1}$

Cell constant $= \dfrac{\text{conductivity}}{\text{conduc}\tan\text{ce}}$

$= \text{Conductivity} \times \text{Resistance}$

$= 0.146 \times 10^{-3}\ Scm^{-1} \times 1500\ \Omega$

$= 0.219\ cm^{-1}$

Q21: (D) $-125.5\ JK^{-1}$

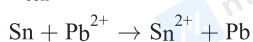$\Delta S = +nF\dfrac{dE}{dt}$

$= 2 \times 96500 \times (-0.00065)$

$= -125.5\ JK^{-1}$

Q22: (D) 2.14

At equilibrium

$E_{cell} = 0$

$E_{cell}^0 = 0.01\ V$

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

$E_{cell} = E_{cell}^0 - \dfrac{0.06}{n} \log Q$

$0 = 0.01 - \dfrac{0.06}{2} \log \dfrac{\left[Sn^{2+}\right]}{\left[Pb^{2+}\right]}$

$0.01 = \dfrac{0.06}{2} \log \dfrac{\left[Sn^{2+}\right]}{\left[Pb^{2+}\right]}$

$$\frac{1}{3} = \log \frac{\left[Sn^{2+}\right]}{\left[Pb^{2+}\right]}$$

$$\frac{\left[Sn^{2+}\right]}{\left[Pb^{2+}\right]} = 10^{\frac{1}{3}} = 2.14$$

Q23: 4

At Cathode :
$$2H^+ (1M) + 2e^- \rightarrow H_2 (g)$$
At Anode :
$$H_2 (g) \rightarrow 2H^+ (xM) + 2e^-$$

Now:

$$E_{Cell} = E^o_{Cell} - \frac{0.0591}{2} \log \left(\frac{\left([H^+]_{anode}\right)^2}{\left([H^+]_{cathode}\right)^2} \right)$$

$$E_{Cell} = E^o_{Cell} - \frac{0.0591}{2} \log \left(\frac{([H^+]_{anode})}{([H^+]_{cathode})} \right)^2$$

$$E_{Cell} = E^o_{Cell} - \frac{0.0591}{2} \log \left(\frac{x}{1} \right)^2$$

$$0.2364 = 0 - \frac{0.0591}{2} \log \left(\frac{x}{1} \right)^2$$

$$0.2364 \times 2 = -0.0591 \log \left(\frac{x}{1} \right)^2$$

$$\frac{0.2364 \times 2}{0.0591} = -\log \left(\frac{x}{1} \right)^2$$

$$-\log \left(\frac{x}{1} \right)^2 = \frac{0.2364 \times 2}{0.0591} = 8$$

$$-2 \log \left(\frac{x}{1} \right) = 8$$

$$\log \left(\frac{x}{1} \right) = -4$$
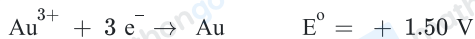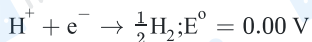
$$x = 10^{-4}$$

$$pH = -\log(x)$$

$$pH = 4$$

Q24: (D) Cr

According to the given data standard reduction potential of Cr is lowest. Lower the standard reduction potential, stronger the reducing agent. Therfore Cr is the strongest reducing agent.
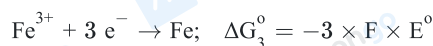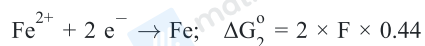
Q25: (A) Cu, Ag, Au

The metals whose reduction potentials are greater than the reduction potential of Hydrogen do not liberate H_2 gas.

$$H^+ + e^- \rightarrow \frac{1}{2}H_2; E^o = 0.00 \text{ V}$$

$$Au^{3+} + 3 e^- \rightarrow Au \qquad E^o = + 1.50 \text{ V}$$

$$Ag^+ + e^- \rightarrow Ag \qquad E^o = + 0.799 \text{ V}$$

$$Cu^{2+} + 2 e^- \rightarrow Cu \qquad E^o = + 0.34 \text{ V}$$

Q26: (C) $P > S > Q > R$

Reducing power, i.e. the tendency to lose electrons increases as the reduction potential decreases

Q27: (C) -0.036 V

$$Fe^{3+} + e^- \rightarrow Fe^{2+}; \quad \Delta G^o_1 = -1 \times F \times 0.77$$

$Fe^{2+} + 2\ e^- \to Fe; \quad \Delta G_2^o = 2 \times F \times 0.44$

$Fe^{3+} + 3\ e^- \to Fe; \quad \Delta G_3^o = -3 \times F \times E^o$

$=0.11$ V

$\therefore\ E^o = \frac{-0.11}{3} = -0.036$ V

Q28: (B) $Ni < Zn < Mg < Ca$

Reducing power $\propto$ Tendency to undergo oxidation

$\propto \frac{1}{\text{reduction potential values}}$

So, reducing power decreases as $Ca > Mg > Zn > Ni$

Q29: 8

The number of faraday's passed $= \frac{2 \times 10^{-3} \times 16 \times 60}{96500}$

$$= 1.99 \times 10^{-5}$$

$\Rightarrow$ number of gram equivalent of Cu^{2+} deposited as $Cu(s)$

$$= 1.99 \times 10^{-5}$$

$\Rightarrow$ number of moles of Cu^{2+} deposited

$$= \frac{1.99}{2} \times 10^{-5} \approx 10^{-5}$$

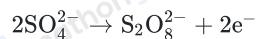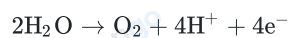Absorbance is directly proportional to $\left[Cu^{2+} \right]$.

No. of moles of Cu^{2+} initially present $= 2 \times 10^{-5}$ moles

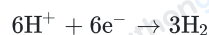$\left[Cu^{2+} \right]_{\text{Initial}} = 2 \times 10^{-5} \times \frac{1000}{250}$

$\qquad = 8 \times 10^{-5} M$

Q30: (A) Thrice that of O_2

$2H_2O \to O_2 + 4H^+ + 4e^-$

$2SO_4^{2-} \to S_2O_8^{2-} + 2e^-$

For 1 mole each of O_2 and $S_2O_8^{2-}$, no of mole of $e^- = 6$ mole e^-

So 6 mole e^-

$6H^+ + 6e^- \to 3H_2$

Hence it produces 3 moles H_2

Q31: 3

$I = 5.0$ A, $t = 4.0$ h

Moles of metal deposited $= \frac{1}{4}$

$Q = 5.0 \times 4 \times 60 \times 60$ coulombs

Charge in faradays

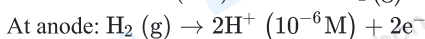$Q = \frac{5 \times 4 \times 60 \times 60}{96000} = 0.75$ F

$\frac{1}{4}$ moles of metal will be deposited by $\frac{3}{4}$ F charge required

So, for 1 mole, $\frac{3}{4} \times 4 = 3F$

Oxidation $= +3$

Q32: 4

Cell reaction would be like, assume the concentration of hydrogen ion in positive electrode is 'x'

$Pt\ (H_2)\ 1\ atm\ |\ H^+\ (10^{-6}M)\ ||\ H^+\ (x\ M)\ |\ Pt\ (H_2)\ 1\ atm$, if $E_{cell} = 0.118$ V.

At cathode: $2H^+\ (xM) + 2e^- \rightarrow H_2\ (g)$

At anode: $H_2\ (g) \rightarrow 2H^+\ (10^{-6}M) + 2e^-$

$E_{Cell} = E^o_{Cell} - \frac{0.591}{2}\log\left(\frac{([H^+]_{anode})^2}{([H^+]_{cathode})^2}\right)$

$E_{Cell} = E^o_{Cell} - \frac{0.591}{2}\log\left(\frac{([H^+]_{anode})}{([H^+]_{cathode})}\right)^2$

$E_{Cell} = E^o_{Cell} - \frac{0.0591}{2}\log\left(\frac{10^{-6}}{x}\right)^2$

$0.118 = 0 - \frac{0.0591}{2}\log\left(\frac{10^{-6}}{x}\right)^2$

$0.118 \times 2 = -0.0591\log\left(\frac{10^{-6}}{x}\right)^2$

$\frac{0.118 \times 2}{0.0591} = -\log\left(\frac{10^{-6}}{x}\right)^2$

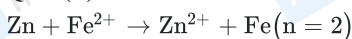$-\log\left(\frac{10^{-6}}{x}\right)^2 = \frac{0.118 \times 2}{0.0591}$

$-2\log\left(\frac{10^{-6}}{x}\right) = 4$

$\log\left(\frac{10^{-6}}{x}\right) = -2$

$\left(\frac{10^{-6}}{x}\right) = 10^{-2}$

$x = 10^{-4}$

Q33: (B) $10^{0.32/0.0295}$

$Zn + Fe^{2+} \rightarrow Zn^{2+} + Fe(n = 2)$

$E = E^o - \frac{0.0591}{n}\log Q$

$0.02905 = E^o - \frac{0.0591}{2}\log\frac{0.01}{0.001}$

$E^o = 0.2905 + 0.0295 = 0.32$ volt

$E^o = \frac{0.0591}{n}\log K_{eq}$

$0.32 = \frac{0.0591}{2}\log K_{eq} = 0.02945\ \log K_{eq}$
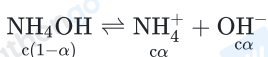
$K_{eq} = 10^{0.32/0.0295}$

Q34: 11.3

$\Lambda^o_{eq}Ba(OH)_2 = \lambda^o Ba^{2+} + \lambda^o_{eq}OH^-$ .....(I)

$\lambda^o_{eq}\ BaCl_2 = \lambda^o_{eq}\ Ba^2 + \lambda^o_{eq}Cl^-$ .....(II)

$\lambda^o_{eq}\ NH_4Cl = \lambda^o_{eq}\ NH_4^+ + \lambda^o_{eq}Cl^-$ .....(III)

$\lambda^o_{eq}\ NH_4OH = \lambda^o_{eq}\ NH_4^+ + \lambda^o_{eq}OH^-$
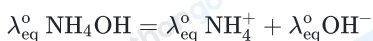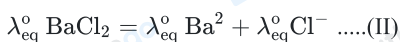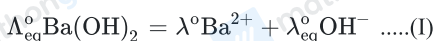
$I + III - II$

$\lambda^o_{eq}\ NH_4OH = (228.8 + 129.8) - 120.3 = 238.33\ cm^2eq^{-1}$

$\lambda_{eq}\ NH_4OH = \frac{4.766 \times 10^{-4} \times 1000}{0.2} = 2.383$

$\alpha = \frac{\lambda_{eq}NH_4OH}{\lambda^o_{eq_1}\ NH_4OH} = 10^{-2}$

$\underset{c(1-\alpha)}{NH_4OH} \rightleftharpoons \underset{c\alpha}{NH_4^+} + \underset{c\alpha}{OH^-}$

$[OH^-] = 0.2 \times 10^{-2} = 2 \times 10^{-3}$

$$pOH = 3 - \log 2 \Rightarrow pH = 14 - (3 - \log 2) = 11.3$$

Q35: 100
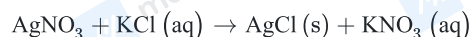
$\kappa = \frac{1}{R}\frac{1}{A}$

i.e. $\frac{1}{A} = R \times \kappa = 300 \times 1.33 \times 10^{-2} \simeq 4.0 \text{cm}^{-1}$

$\kappa_{NaCl} = \frac{1}{R}\frac{1}{A} = \frac{1}{400} \times 4.0$

$\Lambda_{m(NaCl)} = \kappa \times \frac{1000}{M} = \frac{4.0}{400} \times \frac{1000}{0.1} = 100$

Q36: (D) (S)

$$AgNO_3 + KCl\,(aq) \rightarrow AgCl\,(s) + KNO_3\,(aq)$$

Initially there is aq. KCl solution now as solution of $AgNO_3$ is added, $AgCl\,(s)$ is formed. Hence conductivity of solution is almost compensated (or slightly increase) by the formation of KNO_3. After end point conductivity increases more rapidly because addition of excess $AgNO_3$ solution.

Q37: (B) 1.57 V

$$2Fe\,(s) + O_2\,(g) + 4H^+\,(aq) \rightarrow 2Fe^{2+}\,(aq) + 2H_2O\,(l)$$
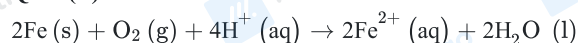
$N = 4$ (no. of moles of electrons involved)

From Nernst equation,
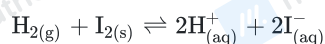
$$E_{cell} = E_{cell}^o - \frac{0.0591}{n}\log Q$$

$$= 1.67 - \frac{0.0591}{4}\log\ \log\frac{(10^{-3})^2}{0.1 \times (10^{-3})^4}\ \left[(H^+) = 10^{-pH}\right]$$

$$= 1.67 - 0.106 = 1.57\ V$$

Q38: 3

The cell reaction is

$H_{2(g)} + I_{2(s)} \rightleftharpoons 2H_{(aq)}^+ + 2I_{(aq)}^-$

Here n = 2 (n = number of electons involved)

$E = E^\circ - \frac{0.0591}{n}\log\frac{[H^+]^2[I^-]^2}{P_{H_2}}$

$0.7714 = 0.535 - \frac{0.0591}{2}\log\frac{[H^+]^2[0.1]^2}{1}$

On solving $(H^+) = 10^{-3}$

$pH = -\log(H^+) = -\log(10^{-3})$

So $pH = 3$

Q39: 50

$$\lambda^o\,(Ba(OH)_2) = \lambda^o\,(BaCl_2) + 2\lambda^o\,(NaOH) - 2\lambda^o\,(NaCl)$$

$$\lambda^o\,(Ba(OH)_2) = 300 + 2\,(250) - 2\,(125)$$

$$\lambda^o\,(Ba(OH)_2) = 300 + 500 - 250$$

$$\lambda^o\,(Ba(OH)_2) = 550\ \text{mho cm}^2\text{mol}^{-1}$$

Q40: 0.91

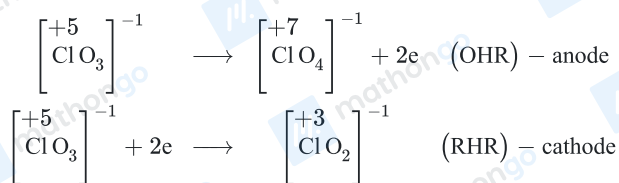From the given cell reaction and nernst equation.

$$E_{cell} = E^o_{cell} - \frac{0.0591}{n} \log \frac{\left[Ni^{2+}\right]}{\left[Ag^{+}\right]^2}$$

$$= 1.05 \text{ V} - \frac{0.0591}{2} \log \frac{[0.160]}{[0.002]^2}$$

$$= 1.05 - \frac{0.0591}{2} \log \left(4 \times 10^4\right)$$

$$= 1.05 - \frac{0.0591}{2} \left(4.6021\right)$$

$$= 1.05 - 0.14 = 0.91 \text{ V}$$

$$E_{cell} = 0.91 \text{ V}$$

Q41: (D) 0.191 M

The redox reaction can be split as

$$\begin{bmatrix} +5 \\ Cl\,O_3 \end{bmatrix}^{-1} \longrightarrow \begin{bmatrix} +7 \\ Cl\,O_4 \end{bmatrix}^{-1} + 2e \quad \left(OHR\right) - anode$$

$$\begin{bmatrix} +5 \\ Cl\,O_3 \end{bmatrix}^{-1} + 2e \longrightarrow \begin{bmatrix} +3 \\ Cl\,O_2 \end{bmatrix}^{-1} \quad \left(RHR\right) - cathode$$

$$2ClO_3^{-1} \underset{}{\overset{n=2e^-}{\rightleftarrows}} ClO_4^{-1} + ClO_2^{-1}$$

$$E^o_{cell} = E^o_{red\,(cath)} - E^o_{red\,(anode)}$$

$$= 0.33 - 0.36 = -0.03 \text{ V}$$

as $E^o_{red\,(Anode)} = E^o_{ClO_4^{-1}/ClO_3^{-1}} = +0.36 \text{ V}$

$$\left(= -E^o_{ClO_3^{-1}/ClO_4^{-1}}\right)$$

at equilibrium $E_{cell} = 0$

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

Writing concentration of speces at equilibrium

$$2ClO_3^{-1} \rightleftharpoons ClO_4^{-1} + ClO_2^{-1}$$
$$\left(1-2x\right) \qquad (x) \qquad (x)$$
$$\text{Molar} \qquad \text{Molar} \qquad \text{Molar}$$
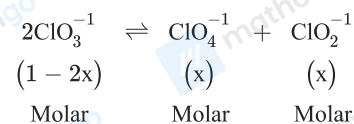
$$E^o_{cell} = -0.03 = \frac{0.059}{2} \log \frac{x^2}{\left(1-2x\right)^2} = 0.059 \log \frac{x}{1-2x}$$

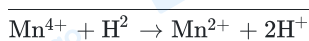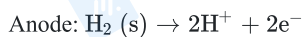$\frac{-0.03}{0.059} = \log\frac{x}{1-2x} \Rightarrow \log\frac{1-2x}{x} = \frac{0.03}{0.059}$

$\log\frac{1-2x}{x} = 0.509$

$\frac{1-2x}{x} = 3.229$

$1 - 2x = 3.229 \, x \Rightarrow 5.229x = 1$

$x = \frac{1}{5.229} = 0.191$ M.

Q42: (D) 2

Anode: $H_2 \, (s) \rightarrow 2H^+ + 2e^-$

Cathode: $Mn^{4+} + 2e^- \rightarrow Mn^{2+}$

$\overline{Mn^{4+} + H^2 \rightarrow Mn^{2+} + 2H^+}$

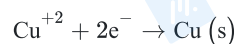$E = E^\circ - \frac{0.059}{2}\log_{10}\left(\frac{[Mn^{2+}][H^+]^2}{[Mn^{4+}]P_{H_2}}\right)$

$0.092 = 0.151 - \frac{0.059}{2}\log_{10}(10^x)$
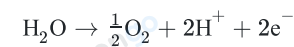
$0.092 = 0.151 - \frac{0.059}{2}x$

$x = 2$

Q43: (C) H_2SO_4

At cathode Cu^{+2} will get reduced due to greater reduction potential
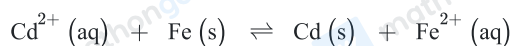
$Cu^{+2} + 2e^- \rightarrow Cu \, (s)$

At anode, H_2O will get preferentially discharged.

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

Hence the final solution will contain H^+ & SO_4^{-2} or H_2SO_4

Q44: 0.05

$$Cd^{2+} \, (aq) \; + \; Fe \, (s) \; \rightleftharpoons \; Cd \, (s) \; + \; Fe^{2+} \, (aq)$$
At eqm. $1 - x$ $-$ $-$ x

At equilibrium, $E^o = \frac{0.0591}{2} \log\left(\frac{Fe^{2+}}{Cd^{2+}}\right)$

$0.037 = \frac{0.0591}{2} \log\left(\frac{x}{1-x}\right)$

$x = \left[Fe^{2+}\right] \Rightarrow 0.947$ M

$$\therefore \left[Cd^{2+} \right] = 0.053 \ M$$

Q45: 3

$$Ag^+ + e^- \rightarrow \underset{\text{1 mol (1eq.)}}{Ag}$$
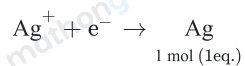
0.02 q. of Ag = 0.02 F = 0.02 × 96500 = 1930 C

1.314 Au is deposited by 0.02 F

197 is deposited by $\frac{0.02}{1.314} \times 197 = 3F$