1. 'A' will have negative polarity and 'B' will have positive polarity.

2. $W = \frac{itE}{96500} = \frac{1 \times 10 \times 60 \times 31.75}{96500}$

   Mass of silver will be different because the equivalent mass of Ag is different. Therefore masses of Cu and Ag deposit is different.

3. No. We can't store AgCl solution in Zinc pot because standard electrode potential of Zinc is less than silver.

4. Thus, quantity of electricity required for oxidation of 1 mol of H_2O to O_2

   Hence, charge required Q = 2 mol $\times$ 96500 C mol^{-1} = 193000 C.

5. Kohlrausch's law of independent migration of ions. The law states that limiting molar conductivity of an electrolyte can be represented as the sum of the individual contributions of the anion and cation of the electrolyte.

6. Standared hydrogen electrode (SHE) is used as reference electrode in determining the standard electrode potential because its reduction and oxidation potential value is 0.0 volt.

7. Aluminium metal cannot be produced by the electrolysis of aqueous solution of aluminium salt because Al^{3+} ions have greater discharge potential than H^+ ions present in solution.

8. Corrosion is an electrochemical phenomenon in which metal gets decomposed in the presence of air and water and forms compounds like oxides, sulphates, carbonates, sulphides etc.

9. The pH of the dil.H_2SO_4 solution will not be affected because there is no change in the concentration of H^+ ions.

   At anode: 2 $H_2O(l) \longrightarrow O_2(g) + 4H^+ + 4e^-$

   At cathode: $4H^+ + 4e^- \longrightarrow 2H_2 \uparrow$
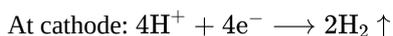
10. When the concentration of all the species involved in a half-cell is unity, then the electrode potential is called standard electrode potential.

11. Zn is oxidized and Ag_2O is reduced (as Ag^+ions change to Ag)

    $E_{cell}^0 = E^0[Ag_2O/Ag](red) + E^0[Zn/Zn^{2+}](ox)$

    = 0.344 + 0.76
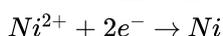
    = 1.104 V

    $\Delta_r G^0 = -nFE^0 cell = -2 \times 96500 \times 1.104 J$

    $= -2.13 \times 10^5 J$

12. given that Quantity of electricity passed $= 5A \times 20 \times 60s$

    = 6000 C time = 20 min

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

    Thus, 2 F, i.e. $2 \times 96500 C$ deposit Ni = 1 mole i.e. 58.7 g

    (at mass of Ni = 58.7)

    Thus 2 F i.e. $2 \times 96500 C$ deposit Ni = 1 mole
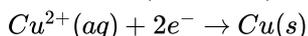
    6000 C will deposit Ni

    $= \frac{58.7}{2 \times 96500} \times 6000 g$ = 1.825 g

13. We have

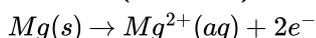 $Mg(s)|Mg^{+2}(0.1M)||Cu^{+2}(1 \times 10^{-3}M|Cu(s))$

    Half cell reactions of this cell are:

    At Cathode (Reduction):

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

    At Anode (Oxidation):

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

    For this cell, we have, n=2 moles of electrons.

    $(\text{Given}, E_{Cu^{2+}/Cu}^{\ominus} = +0.34\,V, E_{Mg^{+2}/Mg}^{\ominus} = -2.37V)$

    $E_{cell}^{\ominus} = E_{reduction}^{\ominus} - E_{oxidation}^{\ominus}$

    $E_{cell}^{\ominus} = E_{(Cu^{2+}/Cu)}^{\ominus} - E_{(Mg^{2+}/Mg)}^{\ominus}$

    = 0.34 - (-2.37)V

    = 2.71V

According to Nernst equation

$$E_{cell} = E_{cell}^{\ominus} - \frac{0.059}{n} \log \frac{[Mg^{2+}]}{[Cu^{2+}]}$$

$$E_{cell} = E_{cell}^{\ominus} - \frac{0.059}{2} \log \left[\frac{0.1}{10^{-3}} \right]$$
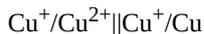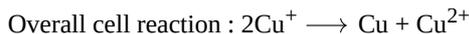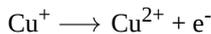
= 2.71 - 0.0295 log10^2
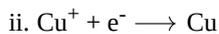
$= 2.71 - 0.0295 \times 2$

= 2.651

$\therefore E_{cell} = 2.651 \ V$

14. i. **Weak electrolyte:** The substance which partially ionized in solution is known as a weak electrolyte. Example: NH_4OH.

   Strong electrolyte: The substance which completely ionized in solution is known as a strong electrolyte. Example: NaCl.

   ii. $Cu^+ + e^- \longrightarrow Cu$

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

   Overall cell reaction : $2Cu^+ \longrightarrow Cu + Cu^{2+}$

   $Cu^+/Cu^{2+} \| Cu^+/Cu$
   E_{cell}^{θ} = 0.52 - 0.16 = 0.36 V
   $\Delta G^o = - nFE_{cell}^{\theta}$
   = -1 × 96500 × 0.36
   = -34740 J mol^{-1}

15. Here, n = 6, T = 273 + 25°C = 298 K

   $E_{cell}^{\circ} = E_{cathode}^{\circ} - E_{anode}^{\circ}$
   = -0.40- (- 0.74) = 0.34 V
   $\Delta_r G^{\circ} = -nFE_{cell}^{0}$
   $\Delta_r G^{\circ} = -6 \times 96500 \times 0.34$ = 196860J mol^{-1}
   $\Delta_r G^{\circ} = -2.303RT \log K_c$
   log K$_c$ = $-\frac{\Delta_r G^{\circ}}{2.303RT} = \frac{196860}{2.303 \times 8.314 \times 298}$
   log K$_c$ = 34.5014

   K$_c$ = antilog 34.5014

   K$_c$ = 3.173×10^{34}

16. i. **Weak electrolytes:** An electrolyte that ionizes partially in solution is called a weak electrolyte. The solution formed contains ions which are in equilibrium with un-ionised molecules, e.g., acetic acid dissolves in water to form H_3O^+ and CH_3COO^- ion.

   The solution contains H_3O^+ (hydronium ion),CH_3COO^- (acetate ion) and unionised CH_3COOH molecule.

   $CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$
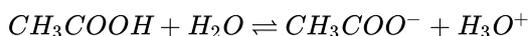
   The degree of ionisation of a weak electrolyte is much less than 1. These have low values of molar conductivities at high concentration. Degree of ionisation and molar conductivity both increases with dilution.

   ii. **Strong electrolyte:** An electrolyte which is almost completely ionised in solution is called a strong electrolyte. The degree of ionisation of a strong electrolyte is 1 or 100% (or nearly so). The solution formed contains ions which are in equilibrium with solid form of strong electrolyte.$NaCl + H_2O \rightarrow Na^+(aq) + Cl^+(aq)$

Strong electrolyte	Weak electrolyte
1. These have higher molar conductivities at all concentrations.	1. These have much lower conductivities at high concentration.
2. λ_m^o values increase very slightly with dilution.	2. λ_m^o values increase sharply with dilution.
3. Degree of ionisation is very high at all concentration i.e., almost fully ionized.	3. Degree of ionisation is very low at high concentration and increases with dilution.
4. Most of the salts like NaCl, Kcl, NaNO$_3$, BaCl$_2$ and mineral acids like HCl, H$_2$SO$_4$, HNO$_3$ and NaOH, KOH etc are common examples of strong electrolytes.	4. Salts like ammonium acetate, acetic acid, aq NH$_4$OH, aqueous CO$_2$ and organic acids and bases are common examples of weak electrolytes.

17. Strong electrolysis: Those electrolytes which dissociate into ions completely into aqueous solution are called strong electrolytes. for example:

KCl, NaOH, H_2SO_4
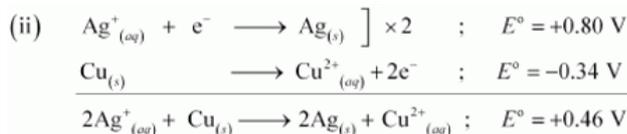
Weak electrolytes: Those electrolytes which do not dissociate into ions completely into aqueous solution are weak electrolytes. for example: CH_3COOH, NH_4Cl
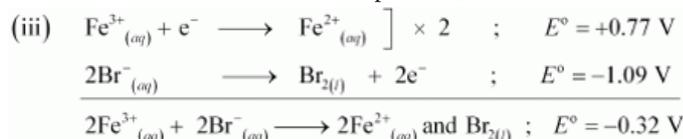
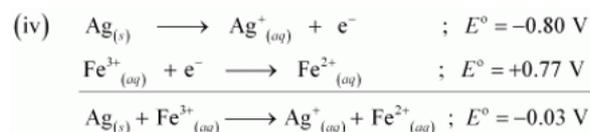18. A reaction is feasible if EMF of the cell is positive.

(i)
$$Fe^{3+}_{(aq)} + e^- \longrightarrow Fe^{2+}_{(aq)} \Big] \times 2; \qquad E^\circ = +0.77 \text{ V}$$
$$\underline{2I^-_{(aq)} \longrightarrow I_{2(s)} + 2e^-; \qquad E^\circ = -0.54 \text{ V}}$$
$$2Fe^{3+}_{(aq)} + 2I^-_{(aq)} \longrightarrow 2Fe^{2+}_{(aq)} + I_{2(s)}; \qquad E^\circ = +0.23 \text{ V}$$
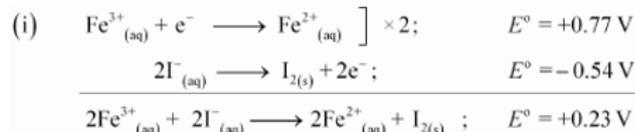
Since E^0 for the overall reaction is positive, the reaction between Fe^{3+}(aq) and I^-(aq) is feasible.

(ii)
$$Ag^+_{(aq)} + e^- \longrightarrow Ag_{(s)} \Big] \times 2 \quad ; \quad E^\circ = +0.80 \text{ V}$$
$$\underline{Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^- \quad ; \quad E^\circ = -0.34 \text{ V}}$$
$$2Ag^+_{(aq)} + Cu_{(s)} \longrightarrow 2Ag_{(s)} + Cu^{2+}_{(aq)} \quad ; \quad E^\circ = +0.46 \text{ V}$$

Since E^0 for the overall reaction is positive, the reaction between Ag^+(aq) and Cu(s) is feasible.

(iii)
$$Fe^{3+}_{(aq)} + e^- \longrightarrow Fe^{2+}_{(aq)} \Big] \times 2 \quad ; \quad E^\circ = +0.77 \text{ V}$$
$$\underline{2Br^-_{(aq)} \longrightarrow Br_{2(l)} + 2e^- \quad ; \quad E^\circ = -1.09 \text{ V}}$$
$$2Fe^{3+}_{(aq)} + 2Br^-_{(aq)} \longrightarrow 2Fe^{2+}_{(aq)} \text{ and } Br_{2(l)} \quad ; \quad E^\circ = -0.32 \text{ V}$$

Since E^0 for the overall reaction is negative, the reaction between Fe^{3+}(aq) and Br^-(aq) is not feasible.

(iv)
$$Ag_{(s)} \longrightarrow Ag^+_{(aq)} + e^- \quad ; \quad E^\circ = -0.80 \text{ V}$$
$$\underline{Fe^{3+}_{(aq)} + e^- \longrightarrow Fe^{2+}_{(aq)} \quad ; \quad E^\circ = +0.77 \text{ V}}$$
$$Ag_{(s)} + Fe^{3+}_{(aq)} \longrightarrow Ag^+_{(aq)} + Fe^{2+}_{(aq)} \quad ; \quad E^\circ = -0.03 \text{ V}$$

Since E^0 for the overall reaction is negative, the reaction between Ag(s) and Fe^{3+}(aq) is not feasible.

(v)
$$Br_{2(aq)} + 2e^- \longrightarrow 2Br^-_{(aq)} \quad ; \quad E^\circ = +1.09 \text{ V}$$
$$\underline{Fe^{2+}_{(aq)} \longrightarrow Fe^{3+}_{(aq)} + e^- \Big] \times 2 \quad ; \quad E^\circ = -0.77 \text{ V}}$$
$$Br_{2(aq)} + 2Fe^{2+}_{(aq)} \longrightarrow 2Br^-_{(aq)} + 2Fe^{3+}_{(aq)} \quad ; \quad E^\circ = +0.32 \text{ V}$$
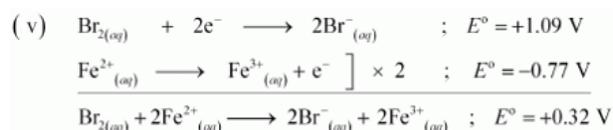
Since E^0 for the overall reaction is positive, the reaction between Br_2(aq) and Fe^{2+}(aq) is feasible.

19. Given that,

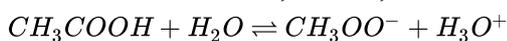Molarity (M) = 0.00241 M

Conductivity (κ) = 7.896 x 10^{-5} S cm^{-1}

For molar conductivity, we have

$$\Lambda_m (S\ cm^2\ mol^{-1}) = \frac{\kappa\ (S\ cm^{-1})\ \times\ 1000\ (cm^3/L)}{molariity\ (mol\ L^{-1})}$$

$$= \frac{7.896 \times 10^{-5} \times 1000}{0.0024}$$

= 32.9 S cm^2 mol^{-1}

For Dissociation Constant, we have,

$$CH_3COOH + H_2O \rightleftharpoons CH_3OO^- + H_3O^+$$

M0	0	0	at t=0
M-Mα	+Mα	+Mα	at t=t

Dissociation constant for this equation is:

$$K_a = \frac{(M\alpha)(M\alpha)}{(M - M\alpha)} = \frac{M\alpha^2}{1-\alpha} \quad ....(A)$$

Also, degree of dissociation is given by

$$\alpha = \frac{\Lambda_m}{\Lambda_m^\circ} \quad ......(B)$$

from (A) and (B), finally we have;

$$K_a = \frac{M\Lambda_m^2}{\Lambda_m^\circ (\Lambda_m^\circ - \Lambda_m)}$$

After putting corresponding values, we get

$$K_a = \frac{0.00241 \times 32.9}{390.5(390.5 - 32.9)}$$

$$\therefore K_a = 5.678 \times 10^{-7}$$