NCERT Solutions for Class-XII Chemistry

CHAPTER -2

NCERT Chemistry Class 12

1. Arrange the following metals in the order in which they displace each other from the solution of their salts. Al, Cu, Fe, Mg and Zn

1. The following is the order in which the given metals displace each other from the solution of their salts.
Mg, Al, Zn, Fe, Cu

2. Given the standard electrode potentials,
$K^+/K = -2.93V$, $Ag^+/Ag = 0.80V$,
$Hg^{2+}/Hg = 0.79V$
$Mg^{2+}/Mg = -2.37\ V$, $Cr^{3+}/Cr = -0.74V$
Arrange these metals in their increasing order of reducing power.

2. Reducing power of metals increase with the decrease of reduction potential. Hence, the increasing order of reducing power will be as,
Ag < Hg < Cr < Mg < K
Explanation -
When the reduction potential is lower, the element has more tendency to get oxidized and thus more will be reducing power. The metal that has more negative electrode potential will be the one with more reducing power. Thus, here potassium(K) has the highest reducing power among the given elements.

3. Depict the galvanic cell in which the reaction $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$ takes place. Further show:
(i) Which of the electrode is negatively charged?
(ii) The carriers of the current in the cell.
(iii) Individual reaction at each electrode.

3. The galvanic cell in which the given reaction takes place is depicted as:
$|Zn_{(s)}|\ |Zn^{2+}_{(aq)}|\ Ag^+_{(aq)}|\ Ag_{(s)}$
(i) Zn electrode (anode) is negatively charged.
(ii) Ions are carriers of current in the cell and in the external circuit, current will flow from silver to zinc.
(iii) The reaction taking place at the anode is given by,

$$Zn_{(s)} \longrightarrow Zn^{2+}_{(aq)} + 2e^-$$

The reaction taking place at the cathode is given by,

$Ag^{+}_{(aq)} + e^{-} \longrightarrow Ag_{(s)}$

4. Calculate the standard cell potentials of galvanic cells in which the following reactions take place:

(i) $2Cr(s) + 3Cd^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 3Cd$

(ii) $Fe^{2+}(aq) + Ag^{+}(aq) \rightarrow Fe^{3+}(aq) + Ag(s)$

Calculate the $\Delta rG\theta$ and equilibrium constant of the reactions.

4. (i) Solution -

Known - $E^0_{Cr^{3+}/Cr} = -0.74$ V

$E^0_{Cd^{2+}/Cd} = -0.40$ V

Δ_rG^0=?

K=?

The galvanic cell of the given reaction is written as -

$Cr_{(s)}|Cr^{3+}_{(aq)}||Cd^{2+}_{(aq)}|Cd_{(s)}$ → Reaction 1

Hence, the standard cell potential is given as,

$E^0 = E^0_R - E^0_L$

$= -0.40 - (-0.74)$

$\therefore E^0 = +0.34$ V

To calculate the standard Gibb's free energy, Δ_rG^0, we use,

$\Delta_rG^0 = -nE^0F$ → Equation 1

where nF is the amount of charge passed and E^0 is the standard reduction electrode potential.

Substituting n=6 (no. of e^- involved in the reaction 1), F=96487 C mol^{-1},

$E^0 = +0.34$ V in Equation 1, we get,

$\Delta_rG^0 = -6 \times 0.34V \times 96487$ C mol^{-1}

$= -196833.48$ CV mol^{-1}

$= -196833.48$ J mol^{-1}

$\therefore \Delta_rG^0 = -196.83348$ kJ mol^{-1}

To find out the equilibrium constant, K, we use the formula,

$$\log K = \frac{nE\infty}{0.0591}$$

$$= \frac{6 ¥ 0.34\ V}{0.0591}$$

$\log K = 34.5177$

K= antilog 34.5177

$\therefore K = 3.294 \times 10^{34}$

Answer - The standard Gibb's free energy, Δ_rG^0 is - 196.83348 kJ mol^{-1} and equilibrium constant, K is 3.294×10^{34}

(ii) Solution -

Known -

$E^0_{Fe^{3+}/Fe^{2+}} = 0.77V$

$E^0_{Ag^{+}/Ag} = 0.80$ V

Δ_rG^0=?

K=?

The galvanic cell of the given reaction is written as -

$Fe^{2+}{}_{(aq)}|Fe^{3+}{}_{(aq)}||\ Ag^{+}{}_{(aq)}|Ag_{(s)} \rightarrow$ Reaction 1

Hence, the standard cell potential is given as,

$E^0=E^0{}_R - E^0{}_L$

$= 0.80 - (0.77)$

$\therefore E^0 = + 0.03$ V

To calculate the standard Gibb's free energy, Δ_rG^0, we use,

$\Delta_rG^0= - nE^0F \rightarrow$ Equation 1

where nF is the amount of charge passed and E^0 is the standard reduction electrode potential.

Substituting n=1 (no. of e^- involved in the reaction 1), F=96487 C mol^{-1}, $E^0 = +$ 0.03V in Equation 1, we get,

$\Delta_rG^0= - 1\times0.03V\times96487$ C mol^{-1}

$= - 2894.61$ CV mol^{-1}

$= - 2894.61$ J mol^{-1}

$\therefore \Delta_rG^0 = - 2.89461$ kJ mol^{-1}

To find out the equilibrium constant, K, we use the formula,

$$\log K = \frac{nE\infty}{0.0591}$$

$$= \frac{1 ¥ 0.03\ V}{0.0591}$$

$\log K= 0.5076$

$K=$ Antilog 0.5076

$\therefore K=3.218$

Answer - The standard Gibb's free energy, Δ_rG^0 is - 2.89461 kJ mol^{-1} and equilibrium constant, K is 3.218

5. Write the Nernst equation and emf of the following cells at 298 K:

(i) Mg(s) | Mg^{2+}(0.001M) | | Cu^{2+}(0.0001 M) | Cu(s)

(ii) Fe(s) | Fe^{2+}(0.001M) | | H^+(1M)|H_2(g)(1bar) | Pt(s)

(iii) Sn(s) | Sn^{2+}(0.050 M) | | H^+(0.020 M) | H_2(g) (1 bar) | Pt(s)

(iv) Pt(s) | $Br_2(l)$ | Br^-(0.010 M) | | H^+(0.030 M) | H_2(g) (1 bar) | Pt(s).

5. (i) For the given reaction, the Nernst equation can be given as:

$$E_{cell} + E^{\Theta}{}_{cell} - \frac{0.0591}{n}\log\frac{[Mg^{2+}]}{[Cu^{2+}]}$$

$$= \{0.34 - (-2.36)\} - \frac{0.0591}{2}\log\frac{.001}{.0001}$$

$$= 2.7 - \frac{0.0591}{2}\log 10$$

$= 2.7 - 0.02955$

$= 2.67$ V (approximately)

(ii) For the given reaction, the Nernst equation can be given as:

$$E_{cell} = E^{\Theta}_{cell} - \frac{0.0591}{n}\log\frac{[Fe^{2+}]}{[H^+]^2}$$

$$= \{0 - (-0.44)\} - \frac{0.0591}{2}\log\frac{0.001}{1^2}$$

$= 0.44 - 0.02955\ (-3)$

$= 0.52865$ V

$= 0.53$ V (approximately)

(iii) For the given reaction, the Nernst equation can be given as:

$$E_{cell} = E^{\Theta}_{cell} - \frac{0.0591}{n}\log\frac{[Sn^{2+}]}{[H^+]^2}$$

$$= \{0 - (-0.14)\} - \frac{0.0591}{2}\log\frac{0.050}{(0.020)^2}$$

$= 0.14 - 0.0295 \times \log 125$

$= 0.14 - 0.062$

$= 0.078$ V

$= 0.08$ V (approximately)

(iv) For the given reaction, the Nernst equation can be given as:

$$E_{cell} = E^{\Theta}_{cell} - \frac{0.0591}{n}\log\frac{1}{[Br^-]^2[H^+]^2}$$

$$= (0 - 1.09) - \frac{0.0591}{2}\log\frac{1}{(0.010)^2(0.030)^2}$$

$$= -1.09 - 0.02955 \times \log\frac{1}{0.00000009}$$

$$= -1.09 - 0.02955 \times \log\frac{1}{9 \times 10^{-8}}$$

$= 1.09 - 0.02955 \times \log(1.11 \times 10^7)$

$= -1.09 - 0.02955(0.0453 + 7)$

$= -1.09 - 0.208$

$= -1.298$ V

6. In the button cells widely used in watches and other devices the following reaction takes place:

$$Zn(s) \longrightarrow Zn^{2+}_{(aq)} + 2e^-\ ;\ E^{\Theta} = 0.76V$$

$$Ag_2O_{(s)} + H_2O_{(l)} + 2e^- \longrightarrow 2Ag_{(s)} + 2OH^-_{(aq)};\ E^{\Theta} = 0.344V$$

$$Zn_{(s)} + Ag_2O_{(s)} + H_2O_{(l)} \longrightarrow Zn^{2+}_{(aq)} + 2Ag_{(s)} + 2OH^-_{(aq)};\ E^{\Theta} = 1.104\ V$$

Zn(s) +

$Ag_2O(s)$ +

$H_2O(l)$ $\rightarrow$
$Zn^{2+}(aq)$ +
$2Ag(s)$ +
$2OH^-(aq)$
Determine
$\therefore$ $E^\ominus$
$D_rG^\ominus$ and $E^\ominus$ for the reaction.

6. = 1.104 V
We know that,
$D_rG^\ominus = nFE^\ominus$
$= -2 \times 96487 \times 1.04$
$= -213043.296$ J
$= -213.04$ kJ

7. Define conductivity and molar conductivity for the solution of an electrolyte. Discuss their variation with concentration.

7. Conductivity of a solution is defined as the conductance of a solution of 1 cm in length and area of cross-section 1 sq. cm. The inverse of resistivity is called conductivity or specific conductance. It is represented by the symbolκ. If ρ is resistivity, then we can write:

$$k = \frac{1}{r}$$

The conductivity of a solution at any given concentration is the conductance (G) of one unit volume of solution kept between two platinum electrodes with the unit area of crosssection and at a distance of unit length.

i.e., $G = k\frac{a}{l} = k.1 = k$

Conductivity always decreases with a decrease in concentration, both for weak and strong electrolytes. This is because the number of ions per unit volume that carry the current in a solution decreases with a decrease in concentration.

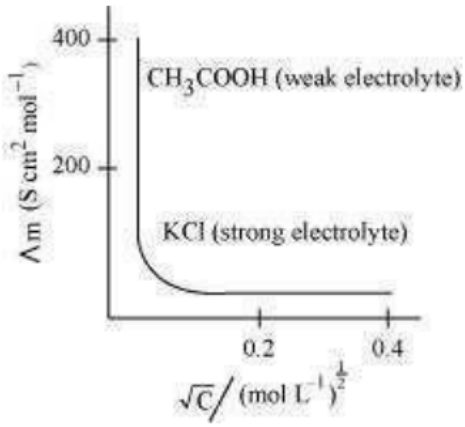
Molar conductivity:

Molar conductivity of a solution at a given concentration is the conductance of volume V of a solution containing 1 mole of the electrolyte kept between two electrodes with the area of cross-section A and distance of unit length.

$$L_m = k\frac{A}{l}$$

Now, l = 1 and A = V (volume containing 1 mole of the electrolyte).
\ L = kV

Molar conductivity increases with a decrease in concentration. This is because the total volume V of the solution containing one mole of the electrolyte increases on dilution. The variation of L_m with $\sqrt{c}$ for strong and weak electrolytes is shown in the following plot:



8. The conductivity of 0.20 M solution of KCl at 298 K is 0.0248 Scm^{-1}. Calculate its molar conductivity.

8. Given, κ
= 0.0248 S cm^{-1} c
= 0.20 M

$$L_m = \frac{k ¥ 1000}{c}$$

∴ Molar, conductivity,

$$= \frac{0.0248 ¥ 1000}{0.2}$$

= 124 Scm^{2}mol^{-1}

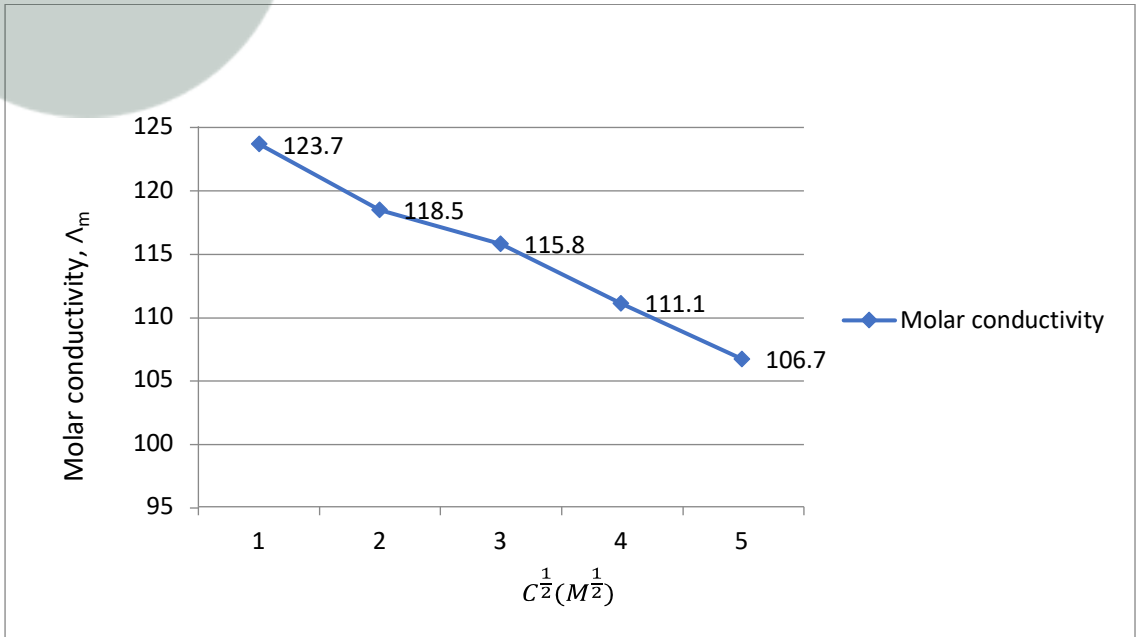
9. The resistance of a conductivity cell containing 0.001M KCl solution at 298 K is 1500 Ω. What is the cell constant if conductivity of 0.001M KCl solution at 298 K is 0.146 × 10^{-3} S cm^{-1}.

9. Given,
Conductivity, κ = 0.146 × 10^{-3} S cm^{-1}
Resistance, R = 1500 Ω
∴ Cell constant = κ × R
= 0.146 × 10^{-3} × 1500
= 0.219 cm^{-1}

10. The conductivity of sodium chloride at 298 K has been determined at different concentrations and the results are given below:
Concentration/M 0.001 0.010 0.020 0.050 0.100
$10^2 \times \kappa/S\ m^{-1}$ 1.237 11.85 23.15 55.53 106.74
for all concentrations and draw a plot between L_m a nd $c^{1/2}$. Find the value calculate L_m of
L^0_m

10. $\frac{S\,cm^{-1}}{100\,S\,m^{-1}} = 1$ (unit conversion factor)

Concentration (M)	κ (S m^{-1})	κ (S cm^{-1})	$L_m = \frac{k}{c} \times 1000\ S\ cm^2\ mol^{-1}$	$C^{\frac{1}{2}} (M^{\frac{1}{2}})$
10^{-3}	1.237×10^{-2}	1.237×10^{-4}	$\frac{1000\times1.237\times10^{-4}}{10^{-3}} = 123.7$	0.0316
10^{-2}	11.85×10^{-2}	11.85×10^{-4}	$\frac{1000\times11.85\times10^{-4}}{10^{-3}} = 118.5$	0.100
2×10^{-2}	23.15×10^{-2}	23.15×10^{-4}	$\frac{1000\times23.15\times10^{-4}}{10^{-3}} = 115.8$	0.141
5×10^{-2}	55.53×10^{-2}	55.53×10^{-4}	$\frac{1000\times55.53\times10^{-4}}{10^{-3}} = 111.1$	0.224
10^{-1}	106.74×10^{-2}	106.74×10^{-4}	$\frac{1000\times106.74\times10^{-4}}{10^{-3}} = 106.7$	0.316



Answer - L^0_m = Intercept on the L_m axis = 124.0 S cm^2 mol^{-1}, which is obtained by extrapolation to zero concentration.

11. Conductivity of 0.00241 M acetic acid is 7.896×10^{-5} S cm^{-1}. Calculate its molar conductivity and if Λ_m^0 for acetic acid is 390.5 S cm^2 mol^{-1}, what is its dissociation constant?

11. Given, $\kappa = 7.896 \times 10^{-5}$ S m^{-1} c
$= 0.00241$ mol L^{-1}

Then molar conductivity, $\Lambda_m = \frac{k}{c}$

$$\frac{7.896 \times 10^{-5}\,S\,cm^{-1}}{0.00241\,mol\,L^{-1}} \times \frac{1000\,cm^3}{L}$$

$= 32.76$S cm^2 mol^{-1}

$\Lambda_m^0 = 390.5\,S\,cm^2\ Mol^{-1}$

Again,

$$a = \frac{\Lambda_m}{\Lambda_m^0} = \frac{32.76\,S\,cm^2\ mol^{-1}}{390.5\,S\,cm^2\ mol^{-1}}$$

Now,

$= 0.084$

$$K_a = \frac{ca^2}{(1-a)}$$

$\therefore$ Dissociation constant,

$$\frac{(0.00241\,mol\,L^{-1})(0.084)^2}{(1-0.084)}$$

$= 1.86 \times 10^{-5}$ mol L^{-1}

12. How much charge is required for the following reductions:
(i) 1 mol of Al^{3+} to Al.
(ii) 1 mol of Cu^{2+} to Cu
(iii) 1 mol of MnO_4^- to Mn^{2+}.

12. (i)The electrode reaction is given as,

$Al^{3+}_{(aq)} + 3e^- \rightarrow Al_{(s)}$

$\therefore$ The quantity of charge required for the reduction of 1 mol of Al^{3+} = 3F
= 3 × 96487 C
= 289461 C

(ii) The electrode reaction is given as,

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

$\therefore$ The quantity of charge required for the reduction of 1 mol of Cu^{2+} = 2F
= 2 × 96487 C

= 192974 C

(iii) The electrode reaction is given as,

$MnO_4 \rightarrow Mn^{2+}$

i.e., $Mn^{7+} + 5e^- \rightarrow Mn^{2+}$

$\therefore$ The quantity of charge required for the reduction of 1 mol of Mn^{7+} = 5F

= 5 × 96487 C

= 482435 C

13. How much electricity in terms of Faraday is required to produce

(i) 20.0 g of Ca from molten $CaCl_2$.

(ii) 40.0 g of Al from molten Al_2O_3.

13. (i) According to the question,

$Ca^{2+} + 2e^{-1}$ æ æÆ Ca

40 g

Electricity required to produce 40 g of calcium = 2 F

Therefore, electricity required to produce 20 g of calcium $= \frac{2 ¥ 20}{40} F$

= 1 F

(ii) According to the question,

$Al^{3+} + 3e^-$ æ æÆ Al

27 g

Electricity required to produce 27 g of Al = 3 F

Therefore, electricity required to produce 40 g of Al = $\frac{3 ¥ 40}{27} F$

= 4.44 F

14. How much electricity is required in coulomb for the oxidation of

(i) 1 mol of H_2O to O_2.

(ii) 1 mol of FeO to Fe_2O_3.

14. (i)The electrode reaction for 1 mole of H_2O is given as,

$H_2O \rightarrow H_2 + \frac{1}{2}O_2$

i.e., O^{2-} Æ $\frac{1}{2}O_2 + 2e^-$

$\therefore$ The quantity of electricity required= 2F

= 2 × 96487 C

= 192974 C

15. A solution of $Ni(NO_3)_2$ is electrolysed between platinum electrodes using a current of 5 amperes for 20 minutes. What mass of Ni is deposited at the cathode?

15. Given,

Current = 5A

Time = 20 × 60 = 1200 s

∴ Charge = current × time

= 5 × 1200

= 6000 C

According to the reaction,

$$Ni^{2+}_{(aq)} + 2e^- \longrightarrow Ni_{(s)}$$

$$58.7\ g$$

Nickel deposited by 2 × 96487 C = 58.71 g

Therefore, nickel deposited by 6000 C = $\frac{58.71 ¥ 6000}{2 ¥ 96487}$ g

= 1.825 g

Hence, 1.825 g of nickel will be deposited at the cathode.

16. Three electrolytic cells A,B,C containing solutions of $ZnSO_4$, $AgNO_3$ and $CuSO_4$, respectively are connected in series. A steady current of 1.5 amperes was passed through them until 1.45 g of silver deposited at the cathode of cell B. How long did the current flow? What mass of copper and zinc were deposited?

16. Equivalent weight is Ag, $E_{Ag} = \frac{180}{1} = 180$

Equivalent weight is Cu, $E_{Cu} = \frac{63.5}{2} = 31.75$

Equivalent weight is Zn, $E_{Zn} = \frac{65}{2} = 32.5$

Using Faraday's second law of electrolysis, to find the mass of Cu and Zn, we use Equation

$$\frac{W_{Cu}}{E_{Cu}} = \frac{W_{Ag}}{E_{Ag}} = \frac{W_{Zn}}{E_{Zn}}$$ Æ Equation 1

fi $W_{Cu} = \frac{W_{Ag} ¥ E_{Cu}}{E_{Ag}}$

$W_{Cu} = \frac{1.45 ¥ 31.75}{108}$

∴ $W_{Cu} = 0.426$ g

fi $W_{Zn} = \frac{W_{Ag} ¥ E_{Zn}}{E_{Ag}}$

$W_{Zn} = \frac{1.45 ¥ 32.5}{108}$

∴ $W_{Zn} = 0.436$ g

To find the time of current flow, using Faraday's first law of electrolysis we get,

$M = Z \times I \times t \Rightarrow$ Equation 2

$\because Z = \frac{\text{Equivalent weight}}{96487}$, Equation 2 becomes,

$M = \frac{108}{96487} ¥ 1.5 ¥ t$

$t = \frac{1.45 ¥ 96487}{108 ¥ 1.5}$

t = 864 seconds.

Answer = The time of current flow, t = 864 seconds, the mass of Cu is 0.426 g and mass of Zn is 0.436 g

17. Using the standard electrode potentials given in Table 3.1, predict if the reaction between the following is feasible:

(i) $Fe^{3+}(aq)$ and $I^{-}(aq)$

(ii) Ag^{+} (aq) and Cu(s)

(iii) Fe^{3+} (aq) and Br^{-} (aq)

(iv) Ag(s) and Fe^{3+} (aq)

(v) Br_2 (aq) and Fe^{2+} (aq).

17. (i)

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

$$2I^{-}_{(aq)} \longrightarrow I_{2(s)} + 2e^{-}; \quad E° = -0.54\ V$$

$$2Fe^{3+}_{(aq)} + 2I^{-}_{(aq)} \longrightarrow 2Fe^{2+}_{(aq)} + I_{2(s)} \; ; E^{o} = +0.23V$$

Since $3_{+(aq)}$ and $I_{-(aq)}$ is feasible.

(ii)

$$Ag^{+}_{(aq)} + e^{-} \longrightarrow Ag_{(s)} \Big] \times 2; \quad E° = +0.80\ V\ Fe$$

$$Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^{-}; \; E° = -0.34\ V$$

$$2Ag^{+}_{(aq)} + Cu_{(s)} \longrightarrow 2Ag_{(s)} + Cu^{2+}_{(aq)} \; ; E^{o} = +0.46V$$

E° for the overall is positive, the reaction between Ag since (aq) and Cu(s) is feasible.

(iii)

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

$$2Br^{-}_{(aq)} \longrightarrow Br_{2(l)} + 2e^{-}; \; E° = -1.09\ V$$

$$2Fe^{3+}_{(aq)} + 2Br^{-}_{(aq)} \longrightarrow 2Fe^{2+}_{(aq)} \text{and } Br_{2(l)} \; ; E^{o} = -0.32V$$

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

(iv)

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

Since E° E for the overall reaction is negative, the reaction between $Ag_{(s)}$ and $Fe_{3+(aq)}$ is not feasible.

$$Br_{2(aq)} + 2e^{-} \longrightarrow 2Br^{-}_{(aq)} \quad ; \quad E° = +1.09\ V$$
$$Fe^{2+}_{(aq)} \longrightarrow Fe^{3+}_{(aq)} + e^{-}\ ]\times 2 \ ; \quad E° = -0.77\ V$$
$$Br_{2(aq)} + 2Fe^{2+}_{(aq)} \longrightarrow 2Br^{-}_{(aq)} + 2Fe^{3+}_{(aq)} ;\ E° = +0.32\ V$$

E° for the overall reaction is positive, the reaction between Br since $2_{(aq)}$ and $Fe_{2+(aq)}$ is feasible.

18. Predict the products of electrolysis in each of the following:
(i) An aqueous solution of $AgNO_3$ with silver electrodes.
(ii) An aqueous solution of $AgNO_3$with platinum electrodes.
(iii) A dilute solution of H_2SO_4 with platinum electrodes.
(iv) An aqueous solution of $CuCl_2$ with platinum electrodes.

18. Given -
All the ions are in aqueous state.
(i) Reaction in solution:
$AgNO_{3(s)} + aq \rightarrow Ag^{+} + NO^{3-}$

$H_2O \Leftrightarrow H^{+} + OH^{-}$
At cathode:
$Ag^{+}_{(aq)} + e^{-} \rightarrow Ag_{(s)}$
Ag^{+} ions have lower discharge potential than H^{+} ions. Hence, Ag^{+} ions get deposited as Ag in preference to H^{+} ions.
At anode:
$Ag_{(s)} \rightarrow Ag^{+}_{(aq)} + e^{-}$
As Ag anode is attacked by NO^{3-} ions, Ag of the anode will dissolve to form Ag^{+} ions in the aqueous solution.
(ii) Reaction in solution:
$AgNO_{3(s)} + aq \rightarrow Ag^{+} + NO^{3-}$

$H_2O \Leftrightarrow H^{+} + OH^{-}$
At cathode:
$2Ag^{+}_{(aq)} + 2e^{-} \rightarrow 2Ag_{(s)}$
Ag^{+} ions have lower discharge potential than H^{+} ions. Hence, Ag^{+} ions get deposited as Ag in preference to H^{+} ions.
At anode:
$2OH^{-}_{(aq)} \rightarrow O_{2(g)} + 2H^{+}_{(aq)} + 4e^{-}$

As anode is not attackable, out of OH^- and NO^{3-} ions, OH^- having lower discharge potential, will be discharged in preference to NO^{3-} ions. These then decompose to give out O_2.

(iii) Reaction in solution:

$H_2SO_{4(aq)} \rightarrow 2H^+_{(aq)} + SO_4^{2-}{}_{(aq)}$

At cathode:

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

At anode:

$2OH^-_{(aq)} \rightarrow O_{2(g)} + 2H^+_{(aq)} + 4e^-$

Answer - $\therefore$ H_2 gas is evolved at cathode and $O_{2(g)}$ is evolved at anode.

(iv) Reaction in solution:

$CuCl_{2(s)} + aq \rightarrow Cu^{2+}_{(aq)} + Cl^-_{(aq)}$

$H_2O \Leftrightarrow H^+ + OH^-$

At cathode:

$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(g)}$

At anode:

$2Cl^-_{(aq)} - 2e^- \rightarrow Cl_{2(g)}$

Answer - $\therefore$ Cu will be deposited at cathode and Cl_2 gas will be liberated at anode.