Kantipur Engineering College

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Solution of Tutorial I, Engineering Chemistry, Electrochemical Cells

1. Write cell notation, electrode reaction and E_{cell} for a cell at 25°C having following electrodes: $E^o_{Zn^{++}/Zn}$ = -0.76 V, $[Zn^{++}]$ = 1.2M, $E^o_{Cu/Cu^{++}}$ = -0.34 V, $[Cu^{++}]$ = 0.01M.

Solution: $E^o_{Cu^{++}/Cu}$ = 0.34 V

Cell Notation Zn / Zn^{++} (1.2 M) // Cu^{++} (0.01M) / Cu

Cell Reaction   Anode: $Zn \longrightarrow Zn^{++} + 2e$ (oxidation)

Cathode: $Cu^{++} + 2e \longrightarrow Cu$ (reduction)

Net Reaction: $Zn + Cu^{++} \longrightarrow Zn^{++} + Cu$ (redox reaction)

Now, $Ecell = E°cell - 2.303\ \frac{RT}{nF} \log Q$

Since, $E°cell = E°cathode - E°anode = 0.34 - (-0.76) = \ 1.1\,Volt$

And, $Q = \frac{[Zn^{++}][Cu]}{[Cu^{++}][Zn]} = \frac{[Zn^{++}]}{[Cu^{++}]} = \frac{1.2}{0.01}$

Therefore, $Ecell = 1.1 - 2.303\ \frac{8.314 \times 298}{2 \times 96500} \log \frac{1.2}{0.01}$ = **1.038 V Answer**

2. From the given electrode couple $E°_{Fe/Fe^{++}}$= 0.44 volt, $[Fe^{++}]$ = 0.5M and $E°_{Ag/Ag^{+}}$= -0.80volt, $[Ag^{+}]$ = 0.2M. Write the (i) electrode reaction (ii) net cell reaction (iii) cell notation (iv) E_{cell} and (v) spontaneity of the cell.

Solution: Cell Notation Fe / Fe^{++} (0.5 M) // Ag^{+} (0.2M) / Ag

Cell Reaction   Anode: $Fe \longrightarrow Fe^{++} + 2e$ (oxidation)

Cathode: $2Ag^{+} + 2e \longrightarrow 2Ag$ (reduction)

Net Reaction: $Fe + 2Ag^{+} \longrightarrow Fe^{++} + 2Ag$ (redox reaction)

Now, $Ecell = E°cell - 2.303\ \frac{RT}{nF} \log Q$

Since, $E°cell = E°cathode - E°anode = 0.80 — (-0.44) = \ 1.24\,Volt$

And, $Q = \frac{[Fe^{++}][Ag]^2}{[Ag^{+}]^2[Fe]} = \frac{[Fe^{++}]}{[Ag^{+}]^2} = \frac{0.5}{0.2^2}$

Therefore, $Ecell = 1.24 - 2.303\ \frac{8.314 \times 298}{2 \times 96500} \log \frac{0.5}{0.2^2}$ = **1.207V Answer**

3. Find the Emf of the cell at 25°C in which silver electrodes are dipped in 0.1 M and 0.5 M silver nitrate solution respectively. $E°_{Ag/Ag^{+}}$= -0.80volt

Solution: $E°_{Ag^{+}/Ag}$= 0.80volt (standard reduction potential)

Reduction rxn: $Ag^{+} + e \longrightarrow Ag$

$\therefore\ Q = \frac{[Ag]}{[Ag^{+}]} = \frac{1}{[Ag^{+}]}$

For first electrode (silver dipped in 0.1M silver nitrate) at 25°C;

$E_1red = E^o red - \frac{0.0591}{n} \log \frac{1}{[Ag^{+}]} = 0.80 - \frac{0.0591}{1} \log \frac{1}{0.1} = 0.7409$ V

For second electrode (silver dipped in 0.5M silver nitrate) at 25°C;

$E_2red = E^o red - \frac{0.0591}{n} \log \frac{1}{[Ag^+]} = 0.80 - \frac{0.0591}{1} \log \frac{1}{0.5} = 0.7822$ V

Since the reduction potential of second electrode is greater it functions as cathode.

$\therefore$ Ecell = E_2red - E_1red = 0.7822 – 0.7409 = **0.413 V Answer**

4. The Emf of a cell consisting of standard AgCl/Ag, Cl^- electrode and copper electrode dipped in $CuSO_4$ is found to be 0.06V. What is the molar concentration of Cu^{++} ions in the cell?

   $AgCl + e \rightarrow Ag + Cl^-$ $\quad E^o = +0.22V$

   $Cu^{++} + 2e \rightarrow Cu$ $\quad E^o = +0.34V$

   Solution: Looking at the reduction potential value, the system of Cu function as cathode and AgCl/Ag, Cl^- function as anode.

   $\therefore$ Ecell = Ecathode – Eanode

   Or, Ecathode = Ecell – Eanode = 0.06 – 0.22 = -0.16 V [anode is in standard state]

   Now, Ecathode = E^ocathode - $\frac{0.0591}{n} \log \frac{1}{[Cu^{++}]}$

   Or, $-0.16 = 0.34 - \frac{0.0591}{2} \log \frac{1}{[Cu^{++}]}$

   $\therefore$ $[Cu^{++}]$ = **0.01 M Answer**

5. The Emf of a following cell of the following cell at 25°C is 1.12 volt. Find the concentration of $CuSO_4$?

   Zn/ Zn^{++} (0.1M) // Cu^{++} (xM)/ Cu. Standard electrode potential provided in the previous questions.

   Solution: Cell reaction from question 1 (same)

   $Zn + Cu^{++} \longrightarrow Zn^{++} + Cu$

   $\therefore Q \frac{[Zn^{++}]}{[Cu^{++}]} = \frac{0.1}{[Cu^{++}]}$

   Since, $E°cell = E°cathode - E°anode = 0.34 - (-0.76) = 1.1\, Volt$

   And, $Q = \frac{[Zn^{++}][Cu]}{[Cu^{++}][Zn]} = \frac{[Zn^{++}]}{[Cu^{++}]} = \frac{0.1}{x}$

   Therefore, $Ecell = 1.1 - 2.303\ \frac{8.314 \times 298}{2 \times 96500} \log \frac{0.1}{x}$

   Or, $1.12 = 1.1 - 2.303\ \frac{8.314 \times 298}{2 \times 96500} \log \frac{0.1}{x}$

   Therefore, x = $[Cu^{++}]$ = **0.5 M Answer**

6. Formulate a cell with the following cell reaction: (standard electrode potential provided in question 1 and 2)

   $2Ag^+ + Cu \rightarrow 2Ag + Cu^{++}$

   Solution: Here copper is oxidized and silver is reduced. Hence, copper functions as anode and silver as cathode. The cell notation for the cell is:

   Cu/Cu^{++} (M) // Ag^+ (M)/ Ag ***Answer***

7. Predict which one of the following reactions is feasible. Given $E^o_{Zn++/Zn} = -0.76$ volt and $E^o_{Cd++/Cd} = -0.40$ volt.

   a) $Zn^{++} + Cd \rightarrow Zn + Cd^{++}$

      Solution: This reaction is not possible because zinc has tendency to oxidize and cadmium has tendency to get reduced. Zinc already is oxidized form and cadmium in reduced form.

   b) $Zn + Cd^{++} \rightarrow Zn^{++} + Cd$

      Solution: This reaction takes place because cadmium has higher reduction potential then zinc, hence Cd^{++} gets reduced to Cd by oxidizing Zn to Zn^{++}.