

**Kantipur Engineering College**  
Dhapakhel, Lalitpur

**Solution of Tutorial I, Engineering Chemistry, Electrochemical Cells**

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

**Solution:**  $E_{\text{Cu}^{++}/\text{Cu}}^{\circ} = 0.34 \text{ V}$

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

Cell Reaction Anode:  $\text{Zn} \longrightarrow \text{Zn}^{++} + 2\text{e}^-$  (oxidation)  
Cathode:  $\text{Cu}^{++} + 2\text{e}^- \longrightarrow \text{Cu}$  (reduction)

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

Now,  $E_{\text{cell}} = E^{\circ}_{\text{cell}} - 2.303 \frac{RT}{nF} \log Q$

Since,  $E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} = 0.34 - (-0.76) = 1.1 \text{ Volt}$

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

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

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

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

Cell Reaction Anode:  $\text{Fe} \longrightarrow \text{Fe}^{++} + 2\text{e}^-$  (oxidation)  
Cathode:  $2\text{Ag}^{+} + 2\text{e}^- \longrightarrow 2\text{Ag}$  (reduction)

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

Now,  $E_{\text{cell}} = E^{\circ}_{\text{cell}} - 2.303 \frac{RT}{nF} \log Q$

Since,  $E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} = 0.80 - (-0.44) = 1.24 \text{ Volt}$

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

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

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.  $E_{\text{Ag}^{+}/\text{Ag}}^{\circ} = -0.80\text{volt}$

**Solution:**  $E_{\text{Ag}^{+}/\text{Ag}}^{\circ} = 0.80\text{volt}$  (standard reduction potential)

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

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

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

$E_{1\text{red}} = E^{\circ}_{\text{red}} - \frac{0.0591}{n} \log \frac{1}{[\text{Ag}^{+}]} = 0.80 - \frac{0.0591}{1} \log \frac{1}{0.1} = 0.7409 \text{ V}$

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

$$E_{2\text{red}} = E^{\circ}_{\text{red}} - \frac{0.0591}{n} \log \frac{1}{[Ag^+]} = 0.80 - \frac{0.0591}{1} \log \frac{1}{0.5} = 0.7822 \text{ V}$$

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

$$\therefore E_{\text{cell}} = E_{2\text{red}} - E_{1\text{red}} = 0.7822 - 0.7409 = \mathbf{0.413 \text{ V Answer}}$$

4. The Emf of a cell consisting of standard AgCl/Ag, Cl<sup>-</sup> electrode and copper electrode dipped in CuSO<sub>4</sub> is found to be 0.06V. What is the molar concentration of Cu<sup>++</sup> ions in the cell?



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

$$\therefore E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}$$

$$\text{Or, } E_{\text{cathode}} = E_{\text{cell}} + E_{\text{anode}} = 0.06 - 0.22 = -0.16 \text{ V [anode is in standard state]}$$

$$\text{Now, } E_{\text{cathode}} = E^{\circ}_{\text{cathode}} - \frac{0.0591}{n} \log \frac{1}{[Cu^{++}]}$$

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

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

5. The Emf of a following cell of the following cell at 25°C is 1.12 volt. Find the concentration of CuSO<sub>4</sub>? Zn/ Zn<sup>++</sup> (0.1M) // Cu<sup>++</sup> (xM)/ Cu. Standard electrode potential provided in the previous questions.

**Solution:** Cell reaction from question 1 (same)



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

$$\text{Since, } E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} = 0.34 - (-0.76) = 1.1 \text{ Volt}$$

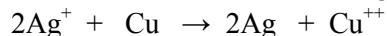
$$\text{And, } Q = \frac{[Zn^{++}][Cu]}{[Cu^{++}][Zn]} = \frac{[Zn^{++}]}{[Cu^{++}]} = \frac{0.1}{x}$$

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

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

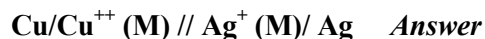
$$\text{Therefore, } x = [Cu^{++}] = \mathbf{0.5 \text{ M Answer}}$$

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



**Solution:** Here copper is oxidized and silver is reduced. Hence, copper functions as anode and silver as cathode.

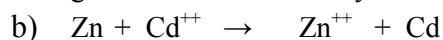
The cell notation for the cell is:



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



**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.



**Solution:** This reaction takes place because cadmium has higher reduction potential than zinc, hence Cd<sup>++</sup> gets reduced to Cd by oxidizing Zn to Zn<sup>++</sup>.