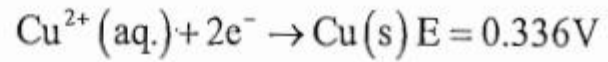
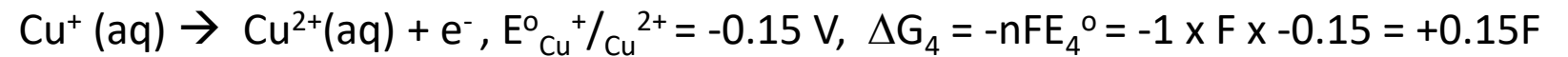


Q1: For the processes, $\text{Cu}^{2+}(\text{aq.}) + \text{e}^- \rightarrow \text{Cu}^+(\text{aq.})$ $E = 0.15\text{V}$

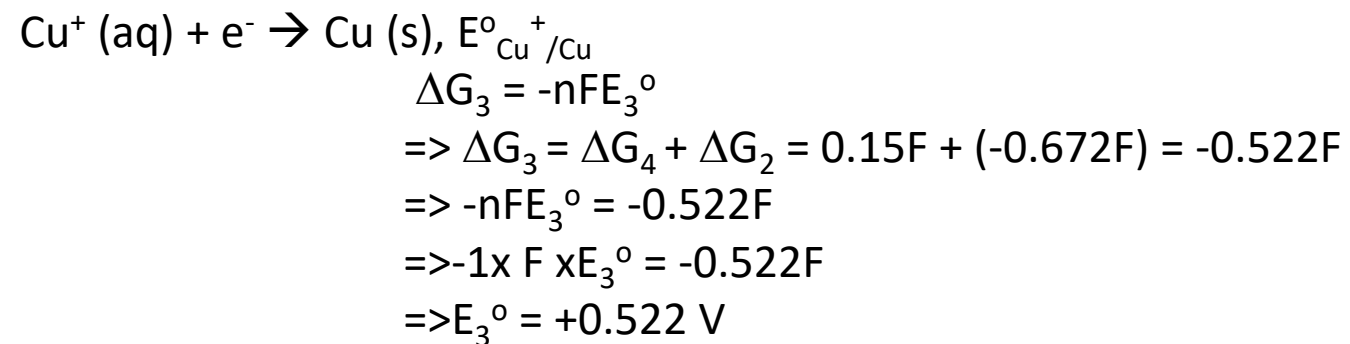


Therefore, E for the process, $\text{Cu}^+(\text{aq.}) + \text{e}^- \rightarrow \text{Cu}(\text{s})$

Ans: Eqn (1) can be re-written as



Adding eqn (4) and eqn (2),

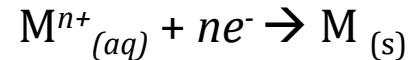


Electrode Potential (E°_{el}) and Nernst Equation

Nernst equation can also be applied for half cell reactions.

The electrode potential of half cells also deviate from the standard electrode potentials when the electrolyte concentration is not 1M or temperature is not 298K.

Taking the example of the reduction reaction:



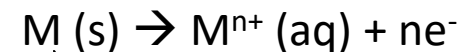
If $E^\circ_{M^{n+}/M}$ is the standard reduction potential, then Nernst equation can be applied as:

$$E_{M^{n+}/M} = E^\circ_{M^{n+}/M} - \frac{RT}{nF} \ln Q_c$$

$$\Rightarrow E_{M^{n+}/M} = E^\circ_{M^{n+}/M} - \frac{RT}{nF} \ln \frac{[M]}{[M^{n+}]}$$

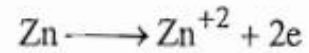
$$\Rightarrow E_{M^{n+}/M} = E^\circ_{M^{n+}/M} - \frac{RT}{nF} \ln \frac{1}{[M^{n+}]}$$

$$\Rightarrow E_{M^{n+}/M} = E^\circ_{M^{n+}/M} + \frac{RT}{nF} \ln [M^{n+}]$$

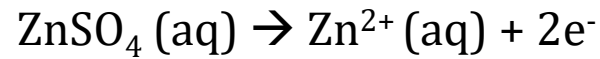


$$\begin{aligned} E_{M/M^{n+}} &= E^\circ_{M/M^{n+}} - (RT/nF) \ln [M^{n+}]/[M] = \\ &= E^\circ_{M/M^{n+}} - (RT/nF) \ln [M^{n+}] \end{aligned}$$

Q2. A zinc electrode is dipped in a 0.1 M solution of ZnSO_4 at 25°C . Assuming that salt is dissociated to 20% at this dilution. Calculate the electrode potential. $E^\circ_{\text{Zn}^{2+}/\text{Zn}} = -0.76\text{V}$



Ans:



At $t = 0$	0.1 M	0	0
	(100 %)	(0%)	(0%)

At $t = t$, 20% dissociated

80%	20%
-----	-----

Concentration of Zn^{2+} after 20% dissociation of $\text{ZnSO}_4 = (20 \times 0.1)/100 \text{ M} = 0.02 \text{ M}$

Applying the Nernst equation,

$$E_{\text{Zn}/\text{Zn}^{2+}} = E^\circ_{\text{Zn}/\text{Zn}^{2+}} - \frac{RT}{nF} \ln \frac{[\text{Zn}^{2+}]}{[\text{Zn}]}$$

Q3. The standard cell potential for the reaction, $\text{Zn(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+} + \text{Cu(s)}$, at 25°C is 0.5 V. The standard free energy (in kJ) is:

(a) -24.1 (b) -48.25 (c) -96.5 (d) -193.0

Q4. Consider the cell:

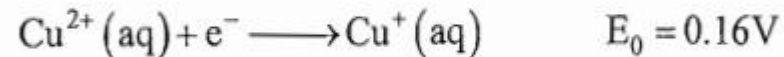
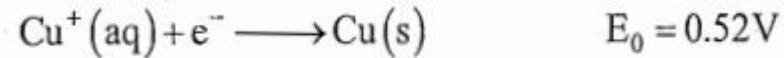
$$\text{Zn} | \text{Zn}^{2+} (a = 0.01) || \text{Fe}^{3+} (a = 0.001), \text{Fe}^{2+} (a = 0.01) | \text{Pt}$$

$E_{\text{cell}} = 1.71 \text{ V}$ at 25°C for the above cell. The equilibrium constant for the reaction:

$$\text{Zn} + 2\text{Fe}^{3+} \rightleftharpoons \text{Zn}^{2+} + 2\text{Fe}^{2+} \text{ at } 25^\circ\text{C} \text{ would be close to}$$

a = Activity coefficient

Q5. Given the standard potential for the following half-cell reaction at 298K,



Calculate the ΔG^0 (kJ) for the reaction,



At anode



At cathode



Net cell reaction



$$E_{\text{cell}}^0 = - \{RT/nF\} \ln K_c$$

$$\Rightarrow \ln K_c = \{-nFE_{\text{cell}}^0\}/RT$$

$$\Rightarrow 2.303 \log K_c = (-2 \times 96500 \times 1.71)/(8.314 \times 298)$$

$$\Rightarrow \log K_c = (-2 \times 96500 \times 1.71)/(8.314 \times 298 \times 2.303)$$