

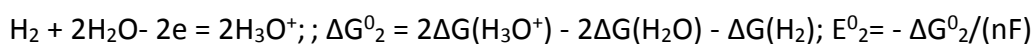
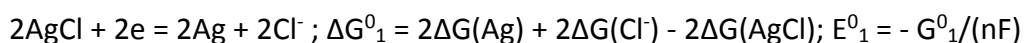
Answer on Question #55705 - Chemistry - Physical chemistry

Question:

For the reaction $\text{H}_2 + 2\text{AgCl} + 2\text{H}_2\text{O} = 2\text{Ag(s)} + 2\text{H}_3\text{O}^+ + 2\text{Cl}^-$ at 250C. The standard free energy of formation of AgCl (s), $\text{H}_2\text{O(l)}$, H_3O^+ and Cl^- are -110, -237, $-200-\alpha$, $(-168+\alpha)$ KJ/mol. Where α is not known. Calculate the cell voltage if this reaction is run at 250C and 0.8 atm in a cell in which $[\text{H}_3\text{O}^+]$ and $[\text{Cl}^-]$ are 0.006M and 0.02M respectively

Solution

The corresponding half-reactions are



For the overall process

$$\Delta G^0 = 2\Delta G^0(\text{H}_3\text{O}^+) + 2\Delta G^0(\text{Ag}) + 2\Delta G^0(\text{Cl}^-) - 2\Delta G^0(\text{AgCl}) - 2\Delta G^0(\text{H}_2\text{O}) - \Delta G^0(\text{H}_2)$$

The reference data for $\Delta G^0(\text{Ag}) = 0$ and $\Delta G^0(\text{H}_2) = 0$ because these substances are simple.

The standard free energy change is $\Delta G^0 = 2(-200 - \alpha) + 2(\alpha - 168) + 2(0) - 2(-110) - 2(-237) = -42 \text{ kJ}$

The non-standard free energy change is $\Delta G = \Delta G^0 + RT \ln([\text{Cl}^-]^2 [\text{H}_3\text{O}^+]^2 / p(\text{H}_2))$

$$\Delta G = -42 \times 10^3 + 8.31 \times 298 \times \ln(0.02^2 \times 0.006^2 / 0.8) = -8.611 \times 10^4 \text{ J}$$

The cell voltage is

$$E = -G/(nF) = 8.611 \times 10^4 / (2 \times 96500) = -0.446 \text{ V}$$

Answer: -0.446 V