

Answer on the question #55386 - Chemistry - General chemistry

Question:

The standard electromotive force (emf) of the cell $\text{Pt(s)} \mid \text{H}_2(\text{g}) \mid \text{HBr}(\text{aq}) \mid \text{AgBr}(\text{s}) \mid \text{Ag}(\text{s})$ was measured over a range of temperatures, and the data were fitted to the following polynomial:
 $E^0 (\text{V}) = 0.07131 - 4.99 \times 10^{-4} (T(\text{K}) - 298\text{K}) - 3.45 \times 10^{-6} (T(\text{K}) - 298\text{K})^2$

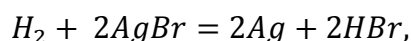
1. Evaluate the standard reaction Gibbs energy ($\Delta_r G^\circ$) at 298 K.

Solution:

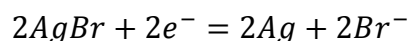
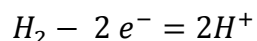
The electrode potential is linked to the Gibbs free energy through the relation:

$$\Delta G = -nFE^0$$

The reaction that occurs in the cell is:



Comprising the half-reactions:



Then, the number of electrons that take part in the reaction is 2. Let's calculate the change in Gibbs free energy:

$$\Delta G^{298} = -2 * 96485.33 * 0.07131 = 13.76 \text{ kJ mol}^{-1}$$

Answer: $\Delta G^{298} = 13.76 \text{ kJ mol}^{-1}$