Answer on the question #55386 - Chemistry - General chemistry

Question:

The standard electromotive force (emf) of the cell Pt(s) | H2(g) | HBr(aq) | AgBr(s) | Ag(s) was measured over a range of temperatures, and the data were fitted to the following polynomial:
E0 (V) = 0.07131- 4.99 ×10-4 (T(K)− 298K)−3.45×10−6 (T(K)− 298K)2
1. Evaluate the standard reaction Gibbs energy (ΔrG°) 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:

$$H_2 + 2AgBr = 2Ag + 2HBr,$$

Comprising the half-reactions:

$$H_2 - 2\,e^- = 2H^+$$

$$2AgBr + 2e^- = 2Ag + 2Br^-$$

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

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

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