Answer on Question #53064 – Chemistry – General chemistry

Question:

By use of the following data at 25°C, solubility product(Ksp) of AgBr 4.88*10^-13 standard electrode potential: E(Ag/Ag)=0 .799 E(Br2(liquid)/Br) =1.065 V(1) Calculate the standard electrode potential for AgBr(s)/Ag at 25°C
(2) Design a cell for calculation of the standard Gibbs Energy formation of AgBr(s).

Answer (1):

According to the conditions the cell should look like:

Ag|AgBr(s), HBr|Br_2 (p_{Br2}= 1 atm), Pt

The electrochemical process on the left side: $Ag \rightarrow Ag^+ + e^-$

The electrochemical process on the right side: $1/2Br_2 + e^- \rightarrow Br^-$,

it means that the total reaction is: $Ag + 1/2Br_2 \rightarrow Ag^+ + Br^- \rightarrow AgBr$

Using the Nernst equation for this process:

$E_{AgBr/Ag,\ Br^-} = E^0_{AgBr/Ag,\ Br^-} - (RT/F) \times \ln((C_{Ag+}C_{Br-})/p_{Br2})$

where C- the concentrations of species and p_{Br2} – the pressures of bromine gas.

$E_{AgBr/Ag,\ Br^-} = E^0_{AgBr/Ag,\ Br^-} - (RT/F) \times \ln(K_{sp}/p_{Br2})$

$E^0_{AgBr/Ag,\ Br^-} = E^0_{Ag+/Ag} + E^0_{Br2/\ 2Br^-} = -0.799\ V + 1.065\ V = 0.266\ V$

The standard potential at 25 C^0 and the pressure 1 atm:

$E_{AgBr/Ag,\ Br^-} = 0.266\ V - 0.059 \times \lg K_{sp} = 0.266\ V + 0.7264\ V = 0.9924\ V$

Answer (2):

Scheme is the same:

Ag|AgBr(s), HBr|Br_2 (p_{Br2}= 1 atm), Pt

The standard Gibbs Energy formation of AgBr(s) at 25 C^0 is calculated according to the equation:

$\Delta G^0_f = -F\Delta E$, where F – the Faraday number.

$\Delta G^0_f = -0.9924 \times 96493$ J/mol = - 95 759.65 J/mol = - 95,760 kJ/mol