Chem 311 Name:
Fall 2007
Quiz 7

An indicator electrode was constructed and used to measure [Cd^{2+}] in contaminated process water from an industrial plant. A silver/silver chloride reference electrode was used as the cathode and the Cd/Cd^{2+} half cell was used as the anode. Calculate the expected Nerstian cell voltage for a 0.200 mM solution of Cd^{2+}. Why might the actual cell voltage of this solution measured with this indicator electrode deviate from this value?

$E^0(Ag/AgCl) = 0.222$ V
$E^0(Ag/AgCl, \text{sat KCl}) = 0.197$ V
$E^0(Hg/Hg_2Cl_2) = 0.268$ V
$E^0(Hg/Hg_2Cl_2, \text{sat KCl}) = 0.241$ V
$E^0(Cd/Cd^{2+}) = -0.402$ V
$E^0(Fe^{2+}/Fe^{3+}) = 0.771$ V

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$E_{cell} = E^+ + E^-$

$E^+ = 0.197$ V

$E^- = -0.402 - 0.05916/2 \log(1/[Cd^{2+}]) = -0.511$ V

$E = 0.197 – (-0.511) = 0.708$ V

This cell will have junction potentials that have not been taken into account in the Nernst Equation. These junction potentials are likely to be on the order of ± 1-30 mV. Thus, the actual cell voltage reading could be expected to range from about 0.680-740 mV.