Answer on Question #56849 - Chemistry - General Chemistry

Question:

For each of the following electrochemical cells, calculate the Eo, write the balanced cell reaction equation and identify what is being oxidized and reduced, and calculate the Q and Ecell. Note that the substances with M given are in solution. All other substances are solids.

a. Cd (s)| Cd(NO3)2 (1.24 M) || FeCl2 (0.500 M), FeCl3 (0.750 M) | Pt (s)

b. Ag, Ag2CrO4 (s)| Na2CrO4 (2.00 M) || Mn(NO3)2 (0.100 M), HNO3 (2.00 M) | MnO2, Pt(s)

Answer:

Standard potential of the cell is the difference between right half-cell potential, where the reduction occurs, and left half-cell potential, where the oxidation occurs. Let's find the potentials of the half-cells for each case and then, subtracting the one for the left half-cell from the one for the right half-cell, find the E_o:

a. $Fe^{3+} + e^- \rightarrow Fe^{2+}$, ϕ_o = +0.77 V
   $Cd^{2+} + 2e^- \rightarrow Cd(s)$, ϕ_o = -0.40 V
$$E_o = E_{right} - E_{left} = +0.77 - (-0.40) = +1.17\,V$$
b. $MnO_2(s) + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$, ϕ_o=+1.23 V
   $Ag_2CrO_4(s) + 2e^- \rightarrow 2Ag(s) + CrO_4^{2-}$ (aq), ϕ_o = +0.45 V
$$E_o = E_{right} - E_{left} = +1.23 - 0.45 = 0.78\,V$$

Balanced cell reaction equations are:

a. $2Fe^{3+} + Cd(s) \rightarrow 2Fe^{2+} + Cd^{2+}$, Fe^{3+} is reduced to Fe^{2+}, Cd(s) is oxidized to Cd^{2+}.
b. $MnO_2(s) + 4H^+ + 2e^- + 2Ag(s) + CrO_4^{2-}(aq) \rightarrow Mn^{2+} + 2H_2O + Ag_2CrO_4(s)$, Mn is reduced from +4 to +2, and Ag is oxidized from 0 to +1.

Now, let's calculate the reaction quotient:

a. $Q = \frac{c(Fe^{2+})^2 \cdot c(Cd^{2+})}{c(Fe^{3+})^2} = \frac{0.5^2 \cdot 1.24}{0.75^2} = 0.55$
b. $Q = \frac{c(Mn^{2+})}{c(CrO_4^{2-}) \cdot c(H^+)^4} = \frac{0.1}{2 \cdot 2^4} = 3.13 \cdot 10^{-3}$

And finally, we can calculate E_{cell}:

a. $E_{cell} = E_o - \frac{RT}{nF} lnQ = 1.17 - \frac{0.05916}{2} \cdot \lg(0.55) = 1.18\,V$
b. $E_{cell} = E_o - \frac{RT}{nF} lnQ = 0.78 - \frac{0.05916}{2} \cdot \lg(3.13 \cdot 10^{-3}) = 0.85\,V$