Electro #1

Balance the following redox reactions and then use your standard reduction potentials in aqueous solution sheet to find the E_{tot} and then predict the sign of $\Delta G^{o}_{rxn,}$ and value of Keq.

Formulas:

$E^{\circ}_{cell} = E^{\circ}_{red} + E^{\circ}_{ox}$

$\Delta G^{\circ} = - nFE^{\circ}_{cell}$ (F= Faraday's Constant = 96500 J/mol*V and n=moles of electrons transferred)

For cells at nonstandard conditions: **$E_{cell} = E^{\circ}_{cell} - \frac{RT}{nF}(\ln Q)$** (R=8.31 J/mol*K)

OR when T = 298K, use NERNST EQUATION: **$E_{cell} = E^{\circ}_{cell} - \frac{0.0257}{n}(\ln Q)$**

1. Cr^{3+} (aq) + Mn (s) → Cr (s) + Mn^{2+} (aq)

2. Ni (s) + Ag^{+} (aq) → Ni^{2+} (aq) + Ag (cr)

3. Fe^{+3} (aq) + Cu (s) → Cu^{+2} (aq) + Fe^{+2} (aq)

4. Sn^{+2} (aq) + Cd (s) → Sn (s) + Cd^{+2} (aq)

5. Mg^{+2} (aq) + Na (s) → Mg (s) + Na^{+1} (aq)