1. Calculate the electrode potential of the following half-cells.
   
   (1) $\text{Ag}^+ (0.0261 \text{ M})/\text{Ag} \quad (E_{\text{Ag}^+/\text{Ag}}^0 = 0.799 \text{ V})$
   
   (2) $\text{Fe}^{3+} (6.72 \times 10^{-4} \text{ M}), \text{Fe}^{2+} (0.100 \text{ M})/\text{Pt} \quad (E_{\text{Fe}^{3+}/\text{Fe}^{2+}}^0 = 0.771 \text{ V})$
   
   (3) $\text{Ag}_2\text{CrO}_4$ (sat’d), $\text{CrO}_4^{2-} (0.050 \text{ M})/\text{Ag} \quad (E_{\text{Ag}_2\text{CrO}_4/\text{Ag}}^0 = 0.446 \text{ V})$
   
   (4) $\text{HCl} (1.76 \text{ M})/\text{H}_2 (0.987 \text{ am}), \text{Pt}$

2. Calculate the equilibrium constant for the reaction $2\text{Fe}^{3+} + 3\Gamma = 2\text{Fe}^{2+} + \Gamma^3$ $ [E_{\text{Fe}^{3+}/\text{Fe}^{2+}}^0 = 0.771, E_{\Gamma^3/\Gamma}^0 = 0.536 \text{ V}]$

   In form of half-cell reaction,
   (i) $3\Gamma \longrightarrow \Gamma^3 + 2e^- \quad \text{(At anode)} \quad E^\circ = 0.77 \text{ V}$
   (ii) $2 \text{Fe}^{3+} + 2e^- \longrightarrow 2\text{Fe}^{2+} \quad \text{(At cathode)} \quad E^\circ = 0.77 \text{ V}$

   Standard reduction potential of both are given, therefore

   
   $E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ$
   
   $= E_{\text{Fe}^{3+}/\text{Fe}^{2+}}^\circ + E_{\Gamma^3/\Gamma}^\circ$
   
   $= 0.77 - 0.54 = 0.23$

   With the help of Nernst equation at equilibrium

   
   $E_{\text{cell}}^\circ = \frac{0.0591}{n} \log_{10} K_{\text{eq}}$

   (because at equilibrium $(E_{\text{cell}} = 0)$

   
   $0.23 = \frac{0.0591}{2} \log_{10} K_{\text{eq}}$

   \[ \therefore K_{\text{eq}} = 6.07 \times 10^7 \]
The oxidation of Ag to Ag$^+$ occurs at the anode, which is the left half-cell. Because the solution contains a source of Cl$^-$, the anodic reaction is
\[
    \text{Ag}^+ (aq) + \text{Cl}^- = \text{AgCl} (s) + e^-
\]
The cathodic reaction, which is the right half-cell, is the reduction of Fe$^{3+}$ to Fe$^{2+}$.
\[
    \text{Fe}^{3+} (aq) + e^- = \text{Fe}^{2+} (aq)
\]
The overall cell reaction, therefore, is
\[
    \text{Ag} (s) + \text{Fe}^{3+} (aq) + \text{Cl}^- (aq) = \text{AgCl} (s) + \text{Fe}^{2+} (aq)
\]
The electrochemical cell’s shorthand notation is
\[
    \text{Ag} (s) | \text{HCl} (aq, a_{Cl^-} = 0.100), \text{AgCl (sat’ld)} \ || \ \text{FeCl}_2 (aq, a_{Fe^{2+}} = 0.0100), \text{FeCl}_3 (aq, a_{Fe^{3+}} = 0.0500) | \text{Pt} (s)
\]
Note that the Pt cathode is an inert electrode that carries electrons to the reduction half-reaction. The electrode itself does not undergo reduction.

4. Write the reactions occurring at the anode and the cathode for the galvanic electrochemical cell with the following shorthand notation.
\[
    \text{Pt} (s) | \text{H}_2 (g), \text{H}^+ (aq) \ || \text{Cu}^{2+} (aq) | \text{Cu} (s)
\]
**Solution:**
The oxidation of H$_2$ to H$^+$ occurs at the anode.
\[
    \text{H}_2 (g) = 2\text{H}^+ (aq) + 2e^-
\]
and the reduction of Cu$^{2+}$ to Cu occurs at the cathode.
\[
    \text{Cu}^{2+} (aq) + 2e^- = \text{Cu} (s)
\]
The overall cell reaction, therefore, is
\[
    \text{Cu}^{2+} (aq) + \text{H}_2 (g) = 2\text{H}^+ (aq) + \text{Cu} (s)
\]