$$Cu(s) + 2Ag^+(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$$

0   +1   +2   0

Cu oxidized
Ag reduced

Oxi half-rxn// $Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$ anode

Red half-rxn// $2Ag^+(aq) + 2e^- \longrightarrow 2Ag(s)$ cathode   2 mol e⁻ transferred

Cell Diagram:

$$\underset{anode}{Cu(s)|Cu^{2+}(aq)} \;||\; \underset{cathode}{Ag^+(aq)|Ag(s)}$$

$$Cu(s) + 2AgNO_3(aq) \longrightarrow Cu(NO_2)_2(aq) + Ag(s)$$

$$Cu(s)|Cu(NO_3)_2(aq) \;||\; AgNO_3(aq)|Ag(s)$$

$$Cu(s)|Cu(NO_3)_2(aq, 1M) \;||\; AgNO_3(aq, 1M)|Ag(s)$$

@25°C

$$E^\circ_{cell} = E^\circ_{red} - E^\circ_{ox}$$
$$= E^\circ_{cat} - E^\circ_{an}$$

$$= \underset{Ag}{0.7996\,V} - \underset{Cu}{0.3419\,V}$$

e⁻ flow from low to high potential

$$= 0.46\,V$$

$$\Delta G^\circ = -nFE^\circ_{cell}$$

$$= -(2\,mol\,e^-)\left(\frac{96.485\,kJ}{mol\,V}\right)(0.46\,V)$$

$$= -88.77\,kJ\,mol^{-1}$$

$$K = e^{-\Delta G^\circ/RT} = 3.57\times10^{15}$$

$$E^\circ_{cell} = \left(\frac{RT}{nF}\right)\ln K$$

$$= \frac{(8.314\,J\,mol^{-1}K^{-1})(298.15\,K)\left(\frac{kJ}{10^3J}\right)}{2\,mol\,e^- \left(\frac{96.485\,kJ}{mol\,V}\right)}\ln 3.57\times10^{15}$$

$$= 0.46\,V$$