Formulas and Information

$E^{O}_{cell} = E^{O}_{red} - E^{O}_{ox}$
Both values are standard reduction potentials taken directly from the table.

$E^{O} > 0$ for a spontaneous reaction

Nernst eqn: $E = E^{O} - \left(\frac{0.0592}{n}\right)\log(Q)$

moles of electrons = $\frac{\text{total charge}}{\text{Faradays constant}} = \frac{(\text{current} \times \text{time})}{\left(9.65 \times 10^{4}\,\frac{\text{C}}{\text{mole e}^{-1}}\right)}$

$\ln(K) = \frac{nFE^{O}}{RT}$
At T = 298 K: $\log(K) = \frac{nE^{O}}{0.0592}$ and $K = 10^{\frac{nE}{0.0592}}$

For a redox reaction: $\Delta G^{o} = -\,nFE^{O}$, where n = # of e^{-1}'s in the reaction, and F is Faraday's constant: 9.65×10^{4} coulombs per mole of e^{-1}'s.

$\Delta G^{o} = -\,RT\ln(K)$ and $\Delta G^{o} < 0$ for a spontaneous reaction

$\Delta G = \Delta H^{o} - T\Delta S^{o}$ and $\Delta G^{o} = \Delta H^{o} - (298\text{ K}) \times \Delta S^{o}$
If ΔH^{o} and ΔS^{o} are either both positive or both negative, then the reaction becomes either spontaneous or non-spontaneous when $\Delta G = 0$

$\Delta G^{o}_{(rxn)} = \Sigma n\Delta G_{f}^{o}{}_{(products)} - \Sigma m\Delta G_{f}^{o}{}_{(reactants)}$

$w = -\,Fx$ $w = -\,mgh$ $w = -\,P(\Delta V)$ if P is constant

$\Delta U = q + w$

$\Delta H = \Delta U + P(\Delta V)$ at constant pressure

$\Delta G = \Delta G^{o} + RT\ln(Q)$ R = 8.314 J/K

Atomic Weight of Cu is 63.55 g/mole

1 Ampere = 1 C/s