oxidation state vs formal charge
redox reactions (oxidizer, reducer, oxidized, reduced)
balancing redox equations (half-reaction)

Ampère (A) - fundamental unit of current
  (how much charge is moving over time)

Volt (V) - electric potential
  \[ 1 \text{V} = 1 \text{J} / \text{C} \]

Coulomb - unit of charge
  \[ 1 \text{J} = 1 \text{V} \cdot \text{C} \]

\[ \Delta G = \text{potential energy} = -\mathcal{E} \cdot \text{charge} \]

\[ \Delta G = -\mathcal{E}_{\text{cell}} \cdot n \cdot F \]

(F) Farad - conversion factor from farads to coulombs

\[ \text{moles of } e^- \text{ transferred in } \text{rxn} \]

Standard Cell

1. Copper
2. Silver nitrate

\[ \text{Silver(I) is a better oxidizing agent than } \text{Cu(II), which is why Silver(I) reacts with Copper (s).} \]
2Ag(NO₃)₂(s) + Cu(s) → 2Ag(s) + Cu(NO₃)₂(s)

Electrode:
physical contact of the cell

Anode:
Cu → Cu²⁺ + 2e⁻

Cathode:

Pt (unreactive)

Ag⁺ + e⁻ → Ag(s)

nitrate is left over

Ag(NO₃)₂

(+) cations

(−) anions

Salt bridge — contains a non-reactive salt (one that will not be oxidized or reduced) that acts as a source of cations and anions to balance the charges generated by the cell.

Standard cell → 2H⁺ + 2e⁻ → H₂(g)

Standard state:

E° = 0.76 V

H⁺/H₂ is used as a standard half-cell when measuring the voltage of redox reactions. When the voltage is measured in this way, it is referred to the standard oxidation or reduction potential (depending on whether hydrogen is reduced or oxidized by the other half of the cell).