

Exam #3

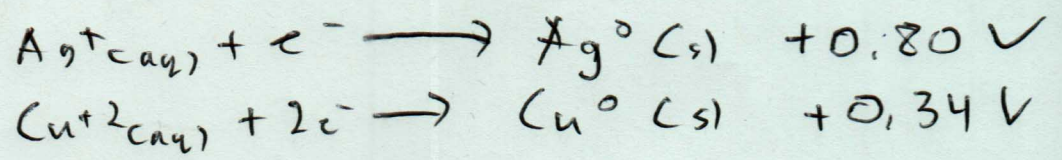
- Calculate oxidation states in redox rxns
- identify oxidizing + reducing agents and oxidizers + reducers
- Balancing redox rxns (acidic vs basic)
- Electrochemical cells
 - Design
 - Direction of e⁻ flow
 - (cell voltage (Nernst)) (Concentration cell)
- Standard state

~~Electrolysis~~

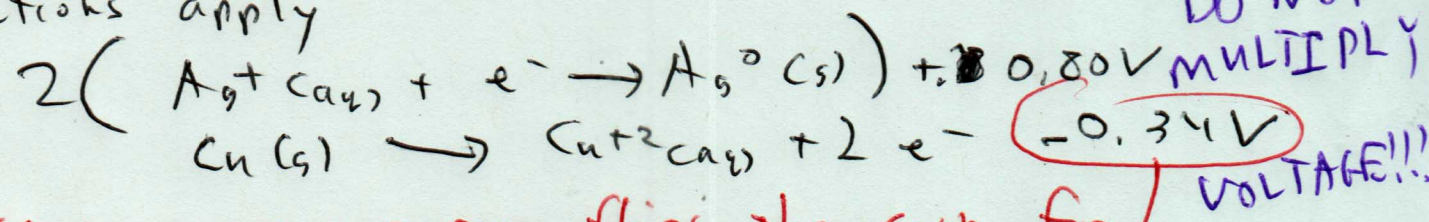
which is the better oxidizing agent, Cu²⁺ or Ag⁺?

Side note: Cu²⁺ and Ag⁺ would be very poor reducing agents since to be a reducing agent means the agent itself gets oxidized and Ag⁺ and Cu²⁺ are already at their normal maximum oxidation states,

since oxidizing agents get reduced, this is the same question as: which is better at being reduced?



Since the reduction potential of Ag⁺ is greater than Cu²⁺, Ag⁺ is the better oxidizing agent, since Ag⁺ would force Cu(s) to oxidize, the following half reactions apply



DO NOT MULTIPLY VOLTAGE!!!

Since reversing a rxn flips the sign for ΔG, it also causes a sign flip for voltage

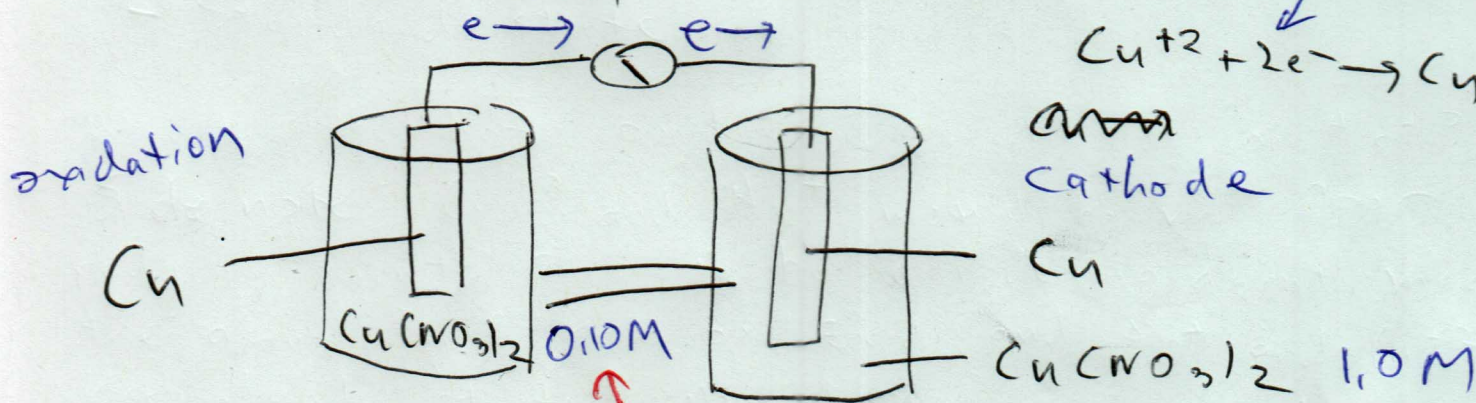
Since voltage is a potential per electron, changing the # of electrons does not change voltage, therefore, when calculating cell voltage, if one half rxn must be multiplied by a stoichiometric factor when balancing the equation, the half-cell voltage is not multiplied by that same factor,

$$\Delta G = \Delta G^\circ + RT \ln Q$$

$$\Delta G^\circ = -RT \ln K$$

pulls Cu^{2+} out of sol'n, reducing concentration

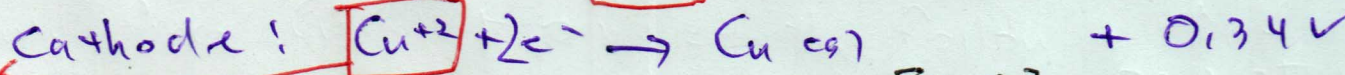
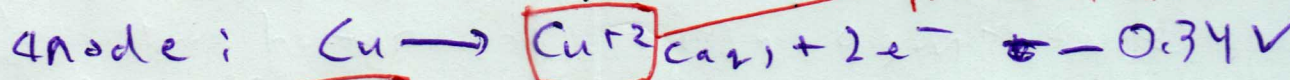
Concentration Cell



$Cu(s) \rightarrow Cu^{2+} + 2e^-$ dilute (product)
 Generates Cu^{2+} , increases concentration

Concentrated (reactant)

In a concentration cell, the standard voltage is always zero, since the oxidation + reduction half-reactions are identical + opposite.



Cu^{2+} forms as a product in dilute sol'n

Concentrated sol'n provides Cu^{2+} as the reactant

$$E = E^\circ - \frac{RT}{nF} \ln Q$$

$$Q = \frac{[Cu^{2+}]_{dil}}{[Cu^{2+}]_{con}} = 0V$$

$$= 0 - \frac{RT}{nF} \ln \frac{0.1}{1} \quad \text{If } Q < 1, E > 0$$

$$> 0$$