

oxidation state vs formal charge

redox reactions (oxidizer, reducer, oxidized, reduced)

balancing redox equations (half-reaction)

Ampere (A) - fundamental unit of current

(how much charge is moving over time)

volt (V) - electric potential

$1V = 1J / C$ coulomb - unit of charge

not on quiz

$1J = 1V \cdot C$

(F) Farad -

$\Delta G =$ potential energy 1 mol of e^-

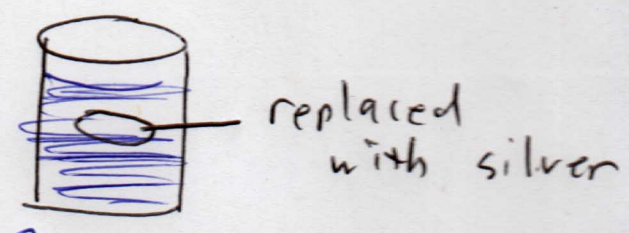
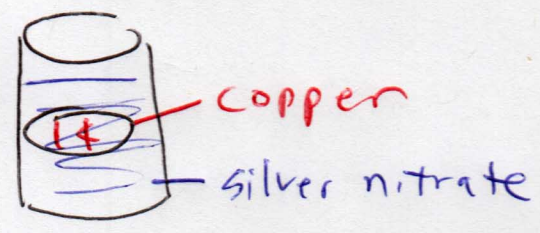
$= -\epsilon_{cell} \cdot \text{charge}$

conversion factor from farads to coulombs

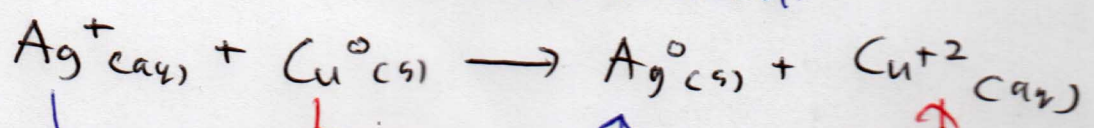
$\Delta G = -\epsilon_{cell} \cdot n \cdot F$

moles of e^- transferred in rxn

Standard Cell

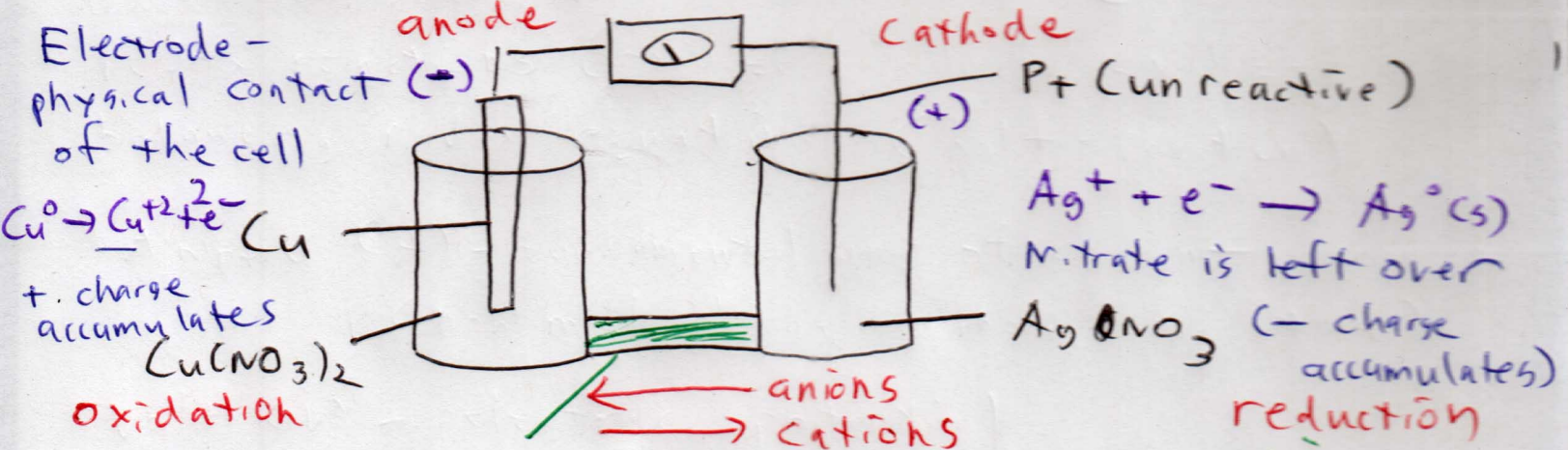
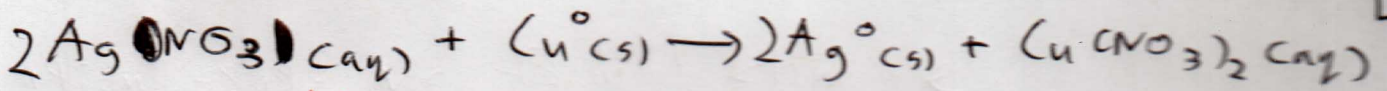


sol'n turns blue from copper nitrate formed.

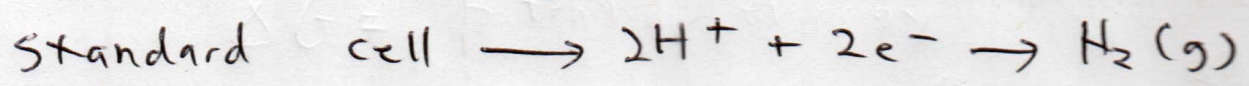


gains $1e^- \rightarrow$ reduced loses $2e^- \rightarrow$ oxidized

Silver(I) is a better oxidizing agent than Cu(II), which is why silver(I) reacts with copper(s).

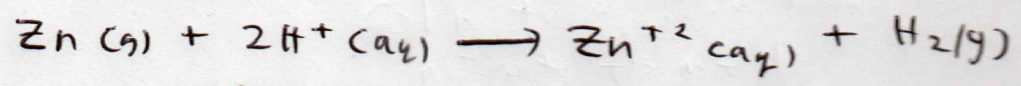
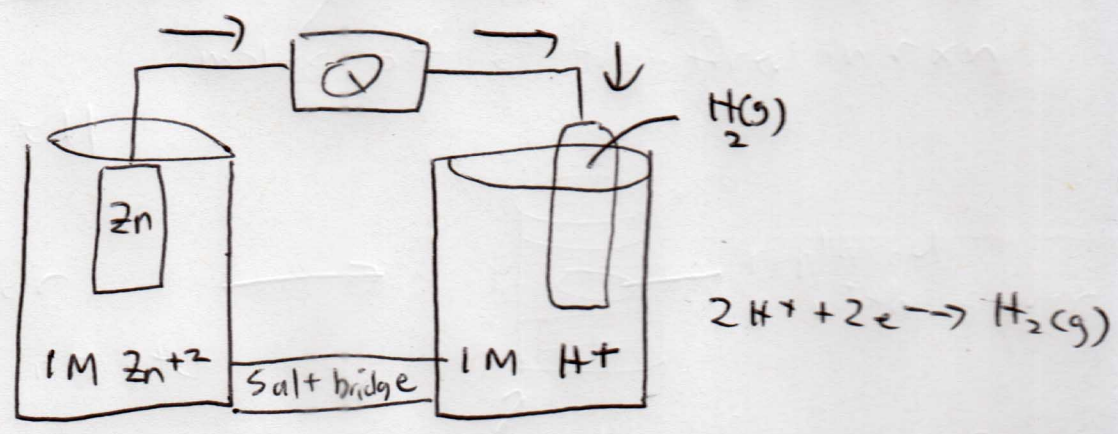


Salt bridge - contains a non-reactive salt (one that will not be oxidized or reduced) that acts as a source of cations and anions to balance the charges generated by the cell



Standard state: \square - standard state

$T = 298\text{K} (25^\circ\text{C})$ $P = 1 \text{ bar}$ $[] = 1.0\text{M}$



$$E^\circ = 0.76\text{V}$$

H^+/H_2 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).