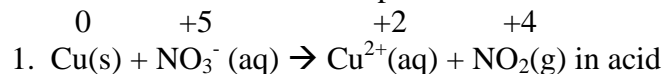


SOHe method of balancing redox equations

*If you balance half reactions using this method, you will always have a balanced overall reaction

Practice using SOHe method

*Know difference between species and element



Elements whose oxidation numbers are changing

Cu: 0 \rightarrow +2 (oxidation)

N: +5 \rightarrow +4 (reduction)

Species on reactant side

Cu (s) (oxidized)

NO₃⁻ (reduced)

Species on product side

Cu²⁺ (aq)

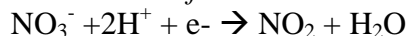
NO₂ (g)

Oxidation half reaction



*add 2e⁻ to product side to make charge on reactant side = charge on product side

Reduction half reaction



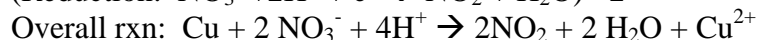
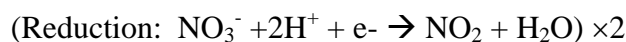
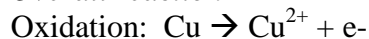
*Add oxygens needed by adding H₂O

*Add H⁺ to side that needs hydrogens (do this if the redox reaction is in acid solution)

if the redox reaction were happening in basic solution, you would add OH⁻ to one side and H₂O to the other side, to balance the hydrogens, because the side that gets H₂O is adding one more H

*Add e⁻ to balance net charge

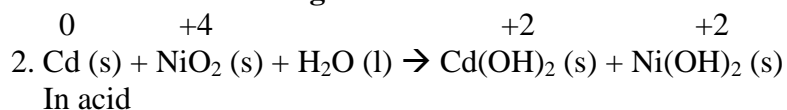
Overall reaction



*Multiply reduction rxn by 2 to make e⁻ cancel

note that the least common multiple of electrons (which canceled) was 2 electrons

Another Practice using SOHe method



Elements whose oxidation numbers are changing

Cd: 0 \rightarrow +2 (oxidized)

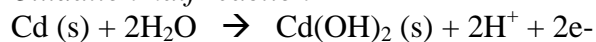
Ni: +4 \rightarrow +2 (reduced)

Species on reactant side

Cd (s) is getting oxidized, so it is the reducing agent

NiO₂ is getting reduced so it is the oxidizing agent

Oxidation half reaction



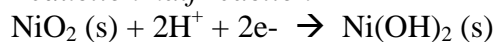
net charge = 0

net charge = +2

2e⁻ added to neutralize charge

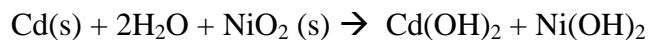
*When adding e⁻ check with oxidation numbers (oxidation side) to ensure correct number of e⁻ were added.

Reduction half reaction



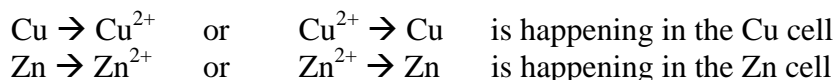
*Check with oxidation numbers (reduction side) that correct number of e⁻ were added

Overall reaction



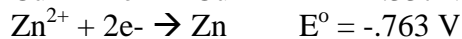
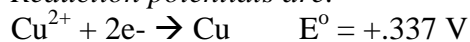
least common multiple of electrons that canceled when you were adding up the half reactions is
n = 2 (need this for Nernst equation – save for later)

Electrochemical Cell Animation shown in class



*Use table of standard reduction potentials to figure out which one of cells is oxidation (anode) and which one is reduction (cathode)

Reduction potentials are:

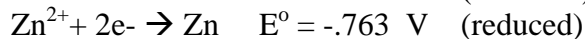
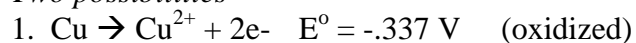


note that both of these half-reactions are written as reductions

for a redox reaction, one of the half-reactions stays a reduction, and the other one turns around and becomes an oxidation

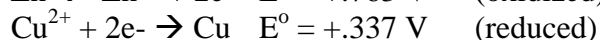
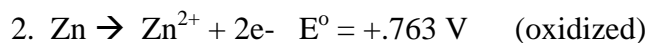
to get E° for an oxidation, just flip the sign of E° for the reduction when you turn the reaction backwards to get the oxidation half-reaction

Two possibilities



$$E^\circ \text{ cell} = (-.337) + (-.763)$$

$$E^\circ \text{ cell} = -1.10 \text{ V}$$



$$E^\circ \text{ cell} = (+.763) + (+.337)$$

$$E^\circ \text{ cell} = +1.10$$

***This one is spontaneous because E° cell is positive**

Parts of an electrochemical cell and cell notation

*Standard cell notation is not in book....only in class – will be on final exam