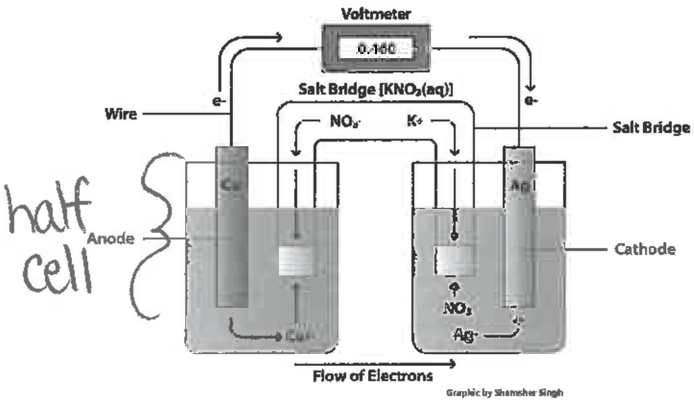
Electrochemical Cells (Batteries)

During redox reactions, electrons pass from the oxidized substance to the reduced substance. This flow of electrons is electricity.

Parts of a Simple Battery (Voltaic Cell)



Made of two half-cells, each having:

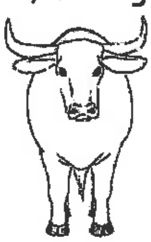
a) A metal electrode
b) A solution of ions

The half-cells are connected via:

c) External wire
d) Salt bridge

a) The Electrodes

- Anode – is where oxidation occurs, making it the negative post (source of e⁻
- Cathode – site of reduction; positive post



An Ox

Anode is where oxidation occurs

Red Cat

Reduction happens at the cathode

Cathode positive post

b) Solution of Ions

- Source of ions to keep cell electrically neutral

c) External Wire

- Allows e^- to flow from anode to cathode

d) Salt bridge

- Allows ions to flow between solutions
- Anions go to anode
- Cations go to cathode

Once connected, the reaction should occur SPONTANEOUSLY. As it proceeds, the concentrations of ions in solution will reach equilibrium and the cell will become DEAD.

Cell Example: Zn and Cu Electrodes

What is oxidized/reduced?

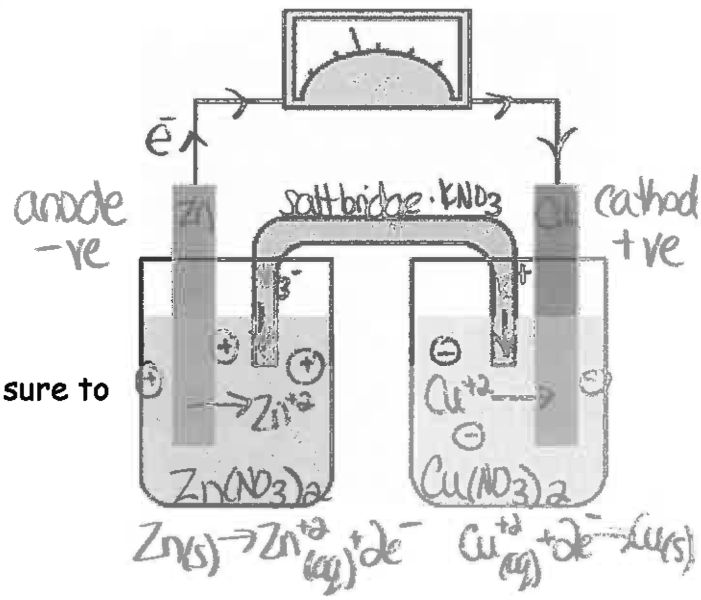
- Look at the Reduction Table. The metal nearest the BOTTOM will undergo Oxidation at the anode.

So… Zn is lower ∴ Zn undergoes oxidation @ anode
Cu is higher ∴ Cu undergoes reduction @ cathode

What are the half reactions? loses e^- gains e^-

- Go back to what is oxidized/reduced.

So… $Zn_{(s)} \rightarrow Zn^{+2}_{(aq)} + 2e^-$
$Cu^{+2}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$



What is the Net Equation?

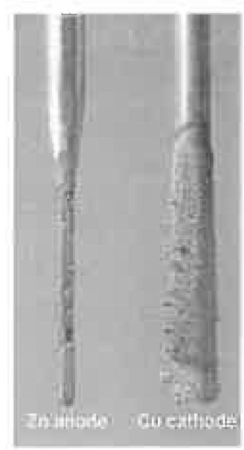
- Add the two half reactions together; making sure to balance e^- and atoms.

So… $Zn_{(s)} \rightarrow Zn^{+2}_{(aq)} + \cancel{2e^-}$
$Cu^{+2}_{(aq)} + \cancel{2e^-} \rightarrow Cu_{(s)}$

$Zn_{(s)} + Cu^{+2}_{(aq)} \rightarrow Zn^{+2}_{(aq)} + Cu_{(s)}$

Which electrode gains/loses weight?

- Look at half reactions! Which forms solid metal? Which dissolves into ions?



So… $Zn_{(s)}$ is dissolving ∴ it loses mass
$Cu_{(s)}$ is getting made ∴ it gains mass

Which way do the ions in the salt bridge flow?

- The ions move towards the solution of opposite charge.

So… K^+ moves toward the cathode
NO_3^- moves toward the anode

STANDARD REDUCTION POTENTIALS FOR HALF-REACTIONS

Ionic concentrations are a 1 M in water at 25 °C

	Half-reaction	E° (Volts)	
very strong oxidizing agents ↑	$F_{2(g)} + 2e^- \rightarrow 2F^-$	+2.87	very weak reducing agents ↓
	$H_2O_2 + 2\,H^+ + 2e^- \rightarrow 2\,H_2$	+1.77	
	$MnO_4^- + 8\,H^+ + 5e^- \rightarrow Mn^{2+} + 4\,H_2O$	+1.52	
	$Au^{3+} + 3e^- \rightarrow Au_{(s)}$	+1.50	
	$Cl_{2(g)} + 2\,e^- \rightarrow 2\,Cl^-$	+1.36	
	$Cr_2O_7^{2-} + 14\,H^+ + 6e^- \rightarrow 2\,Cr^{3+} + 7\,H_2O$	+1.33	
	$MnO_{2(s)} + 4H^+ + 2\,e^- \rightarrow Mn^{2+} + 2\,H_2O$	+1.28	
	$\frac{1}{2}\,O_{2(g)} + 2\,H^+ + 2\,e^- \rightarrow H_2O$	+1.23	
	$Br_2(l) + 2\,e^- \rightarrow 2Br^-$	+1.06	
	$NO_3^- + 4\,H^+ + 3e^- \rightarrow NO_{(g)} + 2\,H_2O$	+0.96	
	$\frac{1}{2}\,O_{2(g)} + 2\,H^+ + 2\,e^- \rightarrow H_2O$	+0.82	
	$Ag^+ + e^- \rightarrow Ag_{(s)}$	+0.80	
	$NO_3^- + 2\,H^+ + e^- \rightarrow NO_{2(g)} + H_2O$	+0.78	
	$Fe^{3+} + e^- \rightarrow Fe^{2+}$	+0.77	
	$O_{2(g)} + 2H^+ + 2\,e^- \rightarrow H_2O_2$	+0.68	
	$I_{2(s)} + 2e^- \rightarrow 2\,I^-$	+0.53	
	$Cu^{2+} + 2e^- \rightarrow Cu_{(s)}$	+0.34	
	$SO_4^{2-} + 4\,H^+ + 2e^- \rightarrow SO_{2(g)} + 2\,H_2O$	+0.17	
	$Sn^{4+} + 2\,e^- \rightarrow Sn^{2+}$	+0.15	
	$S_{(s)} + 2\,H^+ + 2\,e^- \rightarrow H_2S_{(g)}$	+0.14	
	$2H^+ + 2e^- \rightarrow H_{2(g)}$	0.00	
	$Fe^{3+} + 3e^- \rightarrow Fe_{(s)}$	-0.04	
	$Pb^{2+} + 2e^- \rightarrow Pb_{(s)}$	-0.13	
	$Sn^{2+} + 2e^- \rightarrow Sn_{(s)}$	-0.14	
	$Ni^{2+} + 2e^- \rightarrow Ni_{(s)}$	-0.25	
	$Cd^{2+} + 2e^- \rightarrow Cd_{(s)}$	-0.40	
	$Fe^{2+} + 2e^- \rightarrow Fe_{(s)}$	-0.44	
	$Cr^{3+} + 3e^- \rightarrow Cr_{(s)}$	-0.74	
	$Zn^{2+} + 2e^- \rightarrow Zn_{(s)}$	-0.76	
	$2H_2O + 2e^- \rightarrow 2OH^- + H_{2(g)}$	-0.83	
	$Mn^{2+} + 2e^- \rightarrow Mn_{(s)}$	-1.18	
	$Al^{3+} + 3e^- \rightarrow Al_{(s)}$	-1.66	
	$Mg^{2+} + 2e^- \rightarrow Mg_{(s)}$	-2.37	
	$Na^+ + e^- \rightarrow Na_{(s)}$	-2.71	
	$Ca^{2+} + 2e^- \rightarrow Ca_{(s)}$	-2.87	
	$Sr^{2+} + 2e^- \rightarrow Sr_{(s)}$	-2.89	
very weak oxidizing agents	$Ba^{2+} + 2e^- \rightarrow Ba_{(s)}$	-2.90	
	$Cs^+ + e^- \rightarrow Cs_{(s)}$	-2.92	very strong reducing agents
	$K^+ + e^- \rightarrow K_{(s)}$	-2.92	
	$Li^+ + e^- \rightarrow Li_{(s)}$	-3.00	

the strongest oxidizing agent always undergoes a reduction at the cathode

the strongest reducing agent always undergoes an oxidation at the anode

Calculating Cell Voltage aka Standard Cell Potential

- Subtract the E^0 for the oxidation reaction from the E^0 for the reduction reaction

$$E^0_{cell} = E^0_{red} - E^0_{oxid}$$

OR

$$E^0_{cell} = E^0_{cathode} - E^0_{anode}$$

For example: Cu is being reduced
Zn is being oxidized

$Cu^{+2}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ $E^\circ = +0.34V$

$Zn^{+2}_{(aq)} + 2e^- \rightarrow Zn_{(s)}$ $E^\circ = -0.76V$

↳ write as is from table

even if we balance rxn, we do NOT change E°

$E^\circ_{cell} = E^\circ_R - E^\circ_O$
$= +0.34 - (-0.76)$

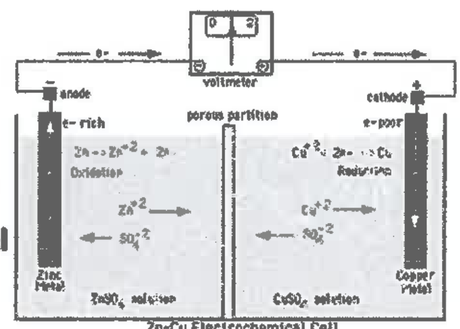
$E^\circ_{cell} = +1.10V$

The + sign of the cell potential tells us the redox reaction is <u>spontaneous</u>, meaning the cell does work.

Representing Cells: Line Notation

Line Notation

An abbreviated representation of an electrochemical cell



Line notation ⇒ $Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$

Anode material | Anode solution || Cathode solution | Cathode material