Electrochemistry

Electrochemistry

- ~The study interactions of chemical and electrical energy.
- Electrochemistry deals with 2 types of processes
- 1. The production of an electric current from an oxidation reduction reaction, **galvanic**.
- 2. The use of an electric current to produce a chemical reaction, **electrolytic**.

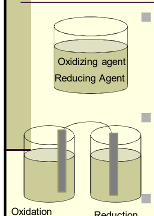
Terminology

- You may have noticed oxygen never gets oxidized, it normally gets reduced.
- The reason for this is because oxygen is an **<u>oxidizing agent</u>**.
- An oxidizing agent is something that causes something else to be oxidized.
- An oxidizing agent readily accepts (or takes) electrons from something else.
- In the process, the oxidizing agent gets reduced.
- A **<u>reducing agent</u>** is something that causes something else to be reduced.

Production of Current

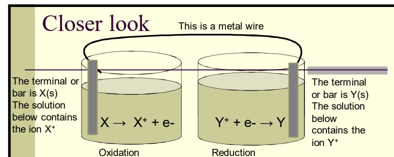
- Oxidation Reactions involve a transfer of electrons.
- Electric current is a <u>movement of electrons</u>.
- In order to produce a <u>usable</u> current, the electrons must be forced across a set path (circuit).
- In order to accomplish this, an oxidizing agent and something to oxidize must be separated from a reducing agent with something to reduce.

Pictures



- An Oxidation Reduction reaction in the same container will have electrons transferring, but we can't harness them.
- Separating the oxidation from the reduction, but connecting them by a wire would allow only electrons to flow.
- The solid metal bars in the diagram are called **terminals.**

Closer look



- On the left, the solution would contain X^+, the terminal would be X(s), the same is true for Y on the right
- We now have **excess electrons** being formed in the oxidizing reaction and **a need for electrons** in the reducing reaction with a **path** for them to flow through.
- However, if electrons did flow through the wire it would cause a negative and positive solution to form.

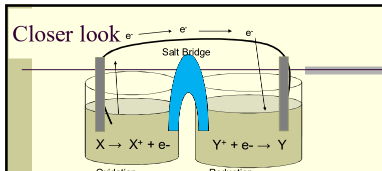
That's not possible

- Or at least it would require a lot of energy.
- A negative solution would theoretically be formed by adding electrons, and a positive one by removing electrons.
- The negative solution would then repel the electrons and stop them from flowing in, and a positive solution would attract the electrons pulling them back where they came from.
- Making it so the charged solutions wouldn't form.
- In order for this to work, I would need a way for ions to flow back and forth but keeping the solutions <u>mostly</u> separated.

Salt Bridge

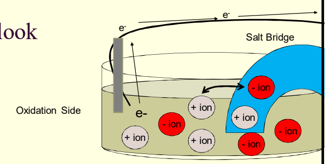
- Salt Bridge- a connector for two solutions previously discussed that allows ions to pass back and forth.
- This can be accomplished by a tube filled with an electrolyte (positive and negative ions) or a porous disc connecting the two solutions.

Closer look



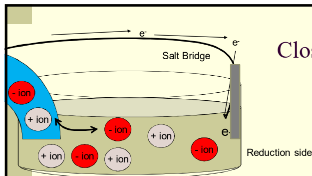
- Now electrons can flow across the wire from the oxidation reaction to the reduction reaction.
- As the oxidation reaction becomes positive, it removes negative ions and adds positive ions to the salt bridge.
- The reduction reaction does the reverse.

Closer look



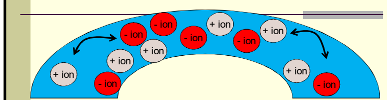
- Zooming in on the oxidizing side
- This would make the salt bridge positive…

Closer look



- (Zooming in on the reducing side)
- if the reverse wasn't happening on this side.

Close up of salt bridge

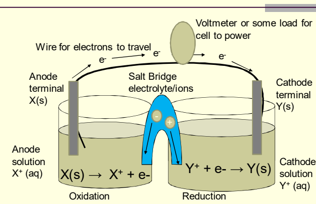


- The ions keep flowing in the salt bridge to keep everything neutral.
- Electrons do also travel across the salt bridge.
- This decreases the cell's effectiveness.

Electrochemical cell

- This is the basic unit of a battery.
- It is also called a galvanic cell, batteries have several galvanic cells linked together.
- Batteries always have two terminals.
- The terminal where oxidation occurs is called the **anode**. An Ox
- The terminal where reduction occurs is called the **cathode**. Red Cat

The total galvanic cell



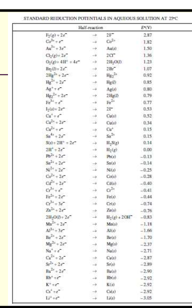
Cell Potential (E_{cell})

- **Cell potential (electromotive force, emf)** is the driving force in a galvanic cell that pulls electrons from the oxidation in one compartment to the reduction in the other.
- The **volt (V)** is the unit of electrical potential.
- Electrical charge is measured in **coulombs (C).**
- A volt is 1 joule of work per coulomb of charge transferred: 1 V = 1 J/C.
- A **voltmeter** is a device which measures cell potential.

How much voltage?

- Voltage of a cell depends on the half reactions.
- You will have a chart of several half reactions reduction potentials for the test.
- Obviously you cannot have two reductions. One will need to be turned into an oxidation.
- To do that flip the half reaction and flip the sign of the half reaction.

The Chart



STANDARD REDUCTION POTENTIALS IN AQUEOUS SOLUTION AT 25°C

Half-reaction			$E°$(V)
$F_2(g) + 2e^-$	→	$2F^-$	2.87
$Co^{3+} + e^-$	→	Co^{2+}	1.82
$Au^{3+} + 3e^-$	→	$Au(s)$	1.50
$Cl_2(g) + 2e^-$	→	$2Cl^-$	1.36
$O_2(g) + 4H^+ + 4e^-$	→	$2H_2O(l)$	1.23
$Br_2(l) + 2e^-$	→	$2Br^-$	1.07
$2Hg^{2+} + 2e^-$	→	Hg_2^{2+}	0.92
$Hg^{2+} + 2e^-$	→	$Hg(l)$	0.85
$Ag^+ + e^-$	→	$Ag(s)$	0.80
$Hg_2^{2+} + 2e^-$	→	$2Hg(l)$	0.79
$Fe^{3+} + e^-$	→	Fe^{2+}	0.77
$I_2(s) + 2e^-$	→	$2I^-$	0.53
$Cu^+ + e^-$	→	$Cu(s)$	0.52
$Cu^{2+} + 2e^-$	→	$Cu(s)$	0.34
$Cu^{2+} + e^-$	→	Cu^+	0.15
$Sn^{4+} + 2e^-$	→	Sn^{2+}	0.15
$S(s) + 2H^+ + 2e^-$	→	$H_2S(g)$	0.14
$2H^+ + 2e^-$	→	$H_2(g)$	0.00
$Pb^{2+} + 2e^-$	→	$Pb(s)$	−0.13
$Sn^{2+} + 2e^-$	→	$Sn(s)$	−0.14
$Ni^{2+} + 2e^-$	→	$Ni(s)$	−0.25
$Co^{2+} + 2e^-$	→	$Co(s)$	−0.28
$Cd^{2+} + 2e^-$	→	$Cd(s)$	−0.40
$Cr^{3+} + e^-$	→	Cr^{2+}	−0.41
$Fe^{2+} + 2e^-$	→	$Fe(s)$	−0.44
$Cr^{3+} + 3e^-$	→	$Cr(s)$	−0.74
$Zn^{2+} + 2e^-$	→	$Zn(s)$	−0.76
$2H_2O(l) + 2e^-$	→	$H_2(g) + 2OH^-$	−0.83
$Mn^{2+} + 2e^-$	→	$Mn(s)$	−1.18
$Al^{3+} + 3e^-$	→	$Al(s)$	−1.66
$Be^{2+} + 2e^-$	→	$Be(s)$	−1.70
$Mg^{2+} + 2e^-$	→	$Mg(s)$	−2.37
$Na^+ + e^-$	→	$Na(s)$	−2.71
$Ca^{2+} + 2e^-$	→	$Ca(s)$	−2.87
$Sr^{2+} + 2e^-$	→	$Sr(s)$	−2.89
$Ba^{2+} + 2e^-$	→	$Ba(s)$	−2.90
$Rb^+ + e^-$	→	$Rb(s)$	−2.92
$K^+ + e^-$	→	$K(s)$	−2.92
$Cs^+ + e^-$	→	$Cs(s)$	−2.92
$Li^+ + e^-$	→	$Li(s)$	−3.05

Which reaction to flip

- $E°_{cell}$ = $E°_{reduction}$ + $E°_{oxidation}$
- The E_{cell} (voltage of the cell) will always be positive. If is negative the cell won't happen on it's own.
- So if you have a reaction of Zn^{2+}/Zn and Cu^{2+}/Cu.
- $Zn^{2+} + 2e^- \rightarrow Zn$ E = -.76 V
- $Cu^{2+} + 2e^- \rightarrow Cu$ E = .34 V

Zinc will need to be flipped to an oxidation to make the cell positive

- $Zn \rightarrow Zn^{2+} + 2e^-$ E = .76 V
- $Cu^{2+} + 2e^- \rightarrow Cu$ E = .34 V
- E_{cell} = 1.10 V
- The overall reaction of the cell is
- $Zn + Cu^{2+} \rightarrow Cu + Zn^{2+}$

Write the equation for and figure out the electric potential of a cell based on…

- Sn^{4+}/Sn^{2+} & Pb^{2+}/Pb
- Zn^{2+}/Zn & Cr^{3+}/Cr
- Fe^{3+}/Fe^{2+} & Co^{3+}/Co^{2+}

Write the equation for and figure out the electric potential of a cell based on…

- Sn^{4+}/Sn^{2+} & Pb^{2+}/Pb
  - $Sn^{4+} + Pb \rightarrow Sn^{2+} + Pb^{2+}$
  - E^o_{cell} = .15+(.13) = .28 V
- Zn^{2+}/Zn & Cr^{3+}/Cr
  - $2\,Cr^{3+} + 3\,Zn \rightarrow 3\,Zn^{2+} + 2\,Cr$
  - E^o_{cell} = .76+(-.74) = .02 V (coefficients don't change this value)
- Fe^{3+}/Fe^{2+} & Co^{3+}/Co^{2+}
  - $Fe^{2+} + Co^{3+} \rightarrow Co^{2+} + Fe^{3+}$
  - E^o_{cell} = 1.82+(-.77) = 1.05 V