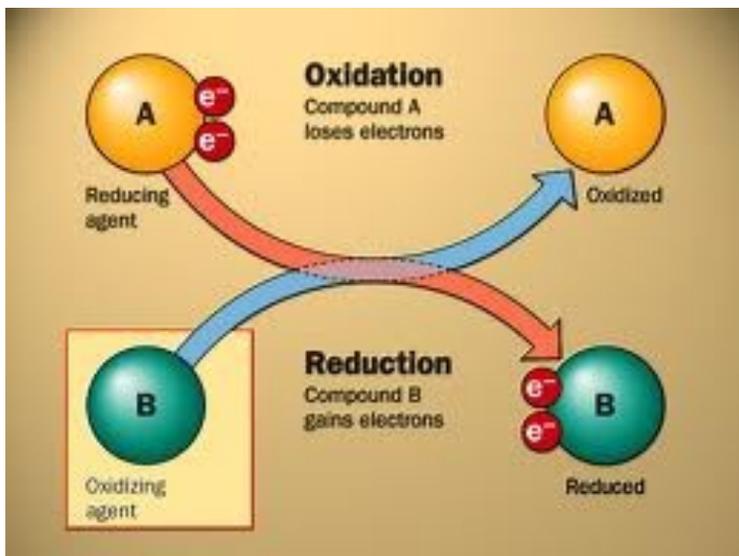
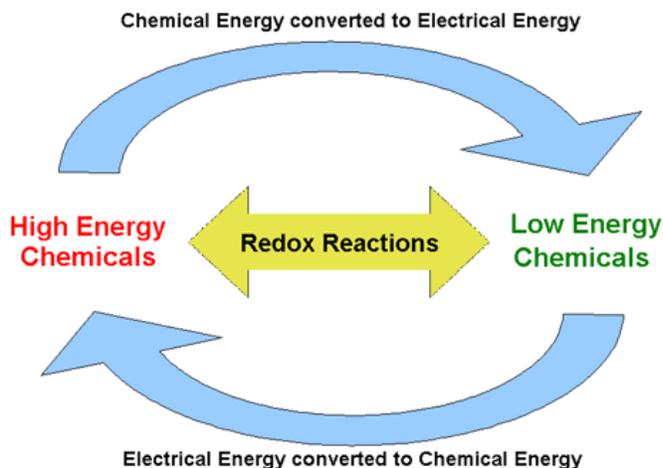


## Chapter 16 Oxidation & Reduction Diary

An **Oxidation-Reduction (redox) reaction** involves the transfer of electrons. So far we've learned already of many kinds of reactions in chemistry; Synthesis, decomposition, single and double replacement and acid-base neutralizations.

**Redox reactions** occur whenever there is a **transfer of electrons**.

Many of the reactions above are also Redox reactions. If you look closely at the balanced chemical equations we've seen in the past, and examine the oxidation numbers of the atoms & ions— if they change from one side of the yield arrow to the other it's a redox reaction!



**Oxidation is the loss of electrons**

**Reduction is the gain of electrons**

**"LEO the Lion goes GER"**

### Half Reactions

A half-reaction can be written to represent oxidation.



This is a common reaction for us, a synthesis that we even did in lab. What's going on with the electrons here? Let's take a closer look.

A half-reaction can be written to represent **oxidation**.



The magnesium atoms become +2 ions by **losing** 2 electrons each.

A half-reaction can be written to represent **reduction**.



At the same time the magnesium atoms become ions, the oxygen atoms become -2 anions by **gaining** one electron each.

**By combining these two reactions, we have an oxidation reaction and a reduction reaction, that are perfectly balanced. Mg loses two electrons and the oxygen picks them both up.**

**Redox has to be balanced. If you can control the flow of these electrons, they can do some work for you in between being oxidized and then reducing something else.**

**Make sure that you balance your oxidation & reductions.**

For every single electron that is oxidized off, it has to be picked up by some other atom or ion and be reduced. No left over electrons ever. Not even one.

### ***Oxidation numbers (states)***

Oxidation numbers (states) can be assigned to atoms and ions. Changes in oxidation numbers indicate that oxidation and reduction have occurred.

Oxidation numbers were used earlier in the year when we put together various ionic and molecular compounds (remember the five different nitrogen and oxygen compounds). Oxidation numbers are listed on our periodic tables. Atoms always have oxidation numbers of ZERO. Ions have oxidation numbers equal to their ionic charge. Atoms in molecular compounds can have a variety of oxidation numbers provided that all the oxidation numbers in a molecule sum to zero.



If you cannot make a particular polyatomic ion “work” as above, remember there are many more oxidation numbers listed outside your “selected” ones.

### **Example reactions**

One of our example reactions from earlier is this one. Let’s look at the oxidation numbers of all the species involved.



Mg atoms are  $\text{Mg}^\circ$   
Oxygen molecules (a pair of atoms) are also  $\text{O}_2^\circ$   
In MgO there is  $\text{Mg}^{+2}$  ion. The oxygen ion is  $\text{O}^{-2}$

The sum of the oxidation numbers of each formula unit of MgO is

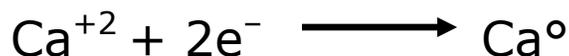
+2 plus -2 = zero (as expected and required)

In another reaction, for example:



This is a single replacement reaction ( table J shows K higher than Ca) so the reaction goes forward as potassium has a higher activity and it will go into solution and bump out the calcium. To do this, the potassium must oxidize (or lose electrons). When this happens, the calcium ions already in solution must pick up these electrons, therefore the  $\text{Ca}^{+2}$  ions are reduced.

The redox half reactions would be:



Since each half reaction is perfectly balanced we can rewrite these pair of reactions together, omitting the electrons—since they balance out on each side of the arrow.

We’ll write the **NET IONIC EQUATION**



It shows that potassium ions become potassium +1 ions, and the calcium +2 cations become atoms. It shows only the NET ion transfer inside the redox reaction.

Let's look at one more reaction now...



The oxidation (loss of electrons) half reaction is:



The reduction (gain of electrons) half reaction is:



The net ionic equation is:



The hydrogen ions, which do not change during this reaction are merely spectators to the redox.

They are called SPECTATOR IONS for this reason.

Single replacement reactions are also redox.

Things get one step more involved when the number of ions oxidized by one part does not match the number gained by the other. Then, balancing reactions comes into play first. For example...

$\text{Al}_{(S)} + \text{CaCl}_{2(AQ)}$  is a single replacement reaction

Since aluminum will oxidize here, forcing the calcium to become reduced, let's look at these half reactions, after we balance and complete this reaction..

Note: each Al loses  $3e^-$  but each Ca only gains  $2e^-$ , we must balance FIRST.



Oxidation is loss of electrons,



Reduction is gain of electrons,



### **Electrochemical cells**

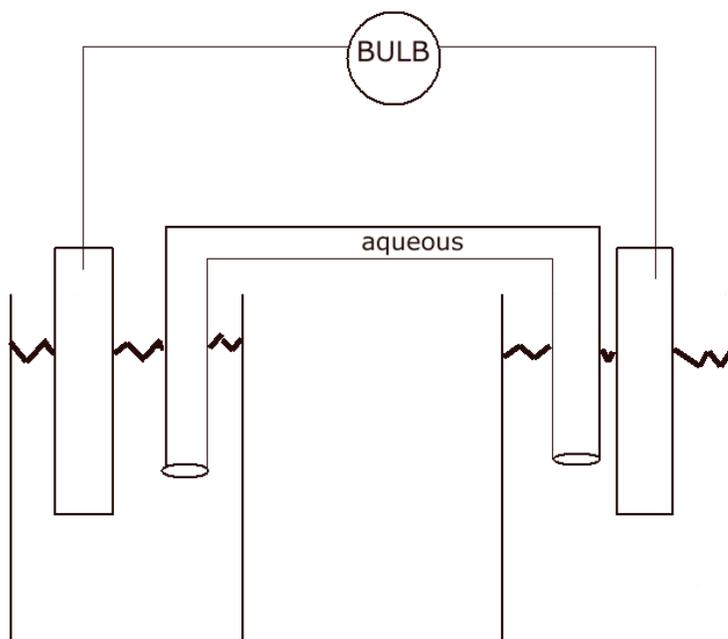
An electrochemical cell can either be **voltaic** or **electrolytic**. In an electrochemical cell, oxidation occurs at the anode and reduction at the cathode (Leo's a RED-CAT)

A cell is an electrochemistry setup that allows for redox reactions to occur. Voltaic cells spontaneously transfer electrons (as long as they can flow), so that the chemistry produces electricity.

Batteries are voltaic cells.

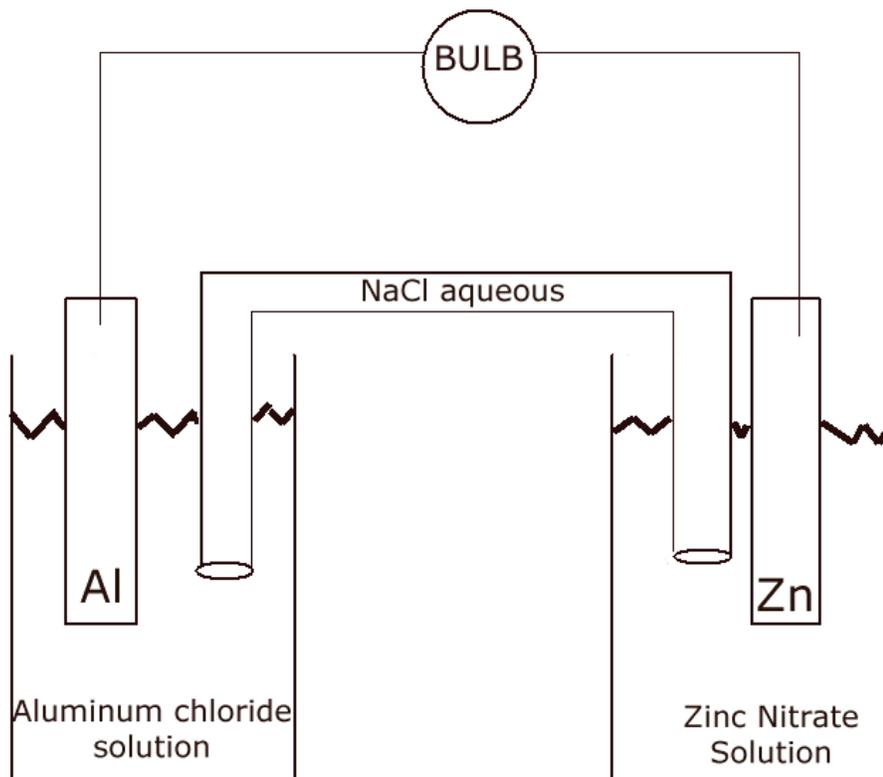
**This happens**  
**SPONTANEOUSLY, without**  
**the addition of energy.**

When we add in the metals, the solutions, and the salt bridge, it then becomes a battery.



**To the left is an aluminum zinc voltaic cell.**

Each metal must be in a solution which has itself as an ion. Aluminum metal is in the aluminum chloride solution. Zinc is in a zinc nitrate solution. There is a NaCl aqueous solution in the salt bridge. The metals are connected by wire, through a bulb.



Both Al and Zn would like to oxidize into ions, losing electrons. Only one can, table J decides for us.

**It's aluminum that will oxidize, not zinc. The more active metal will give away its electrons.**

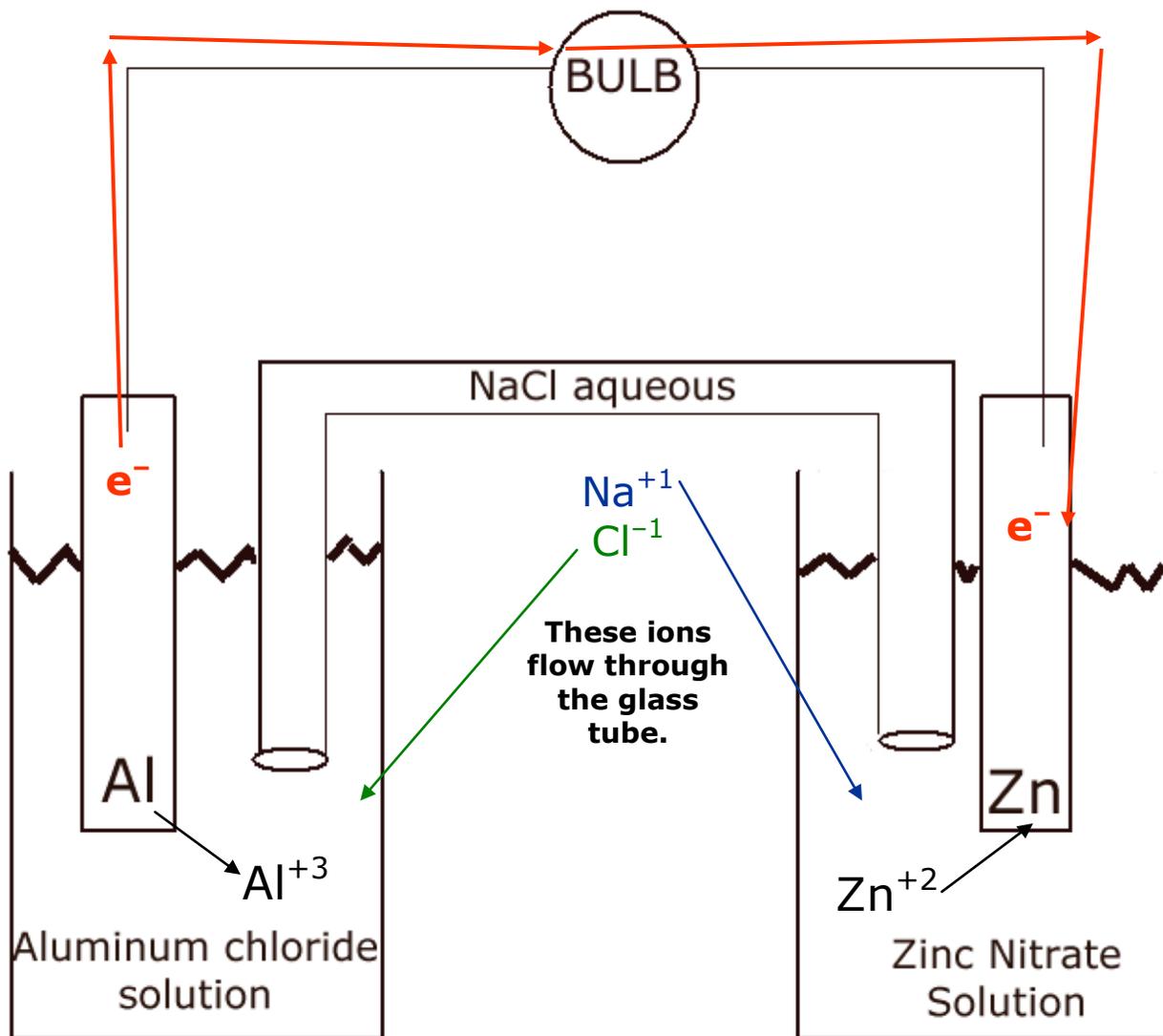
If Al oxidizes into  $\text{Al}^{+3}$  cations, the ions jump into solution, the electrons flow up through the wire to the zinc bar.

**Zinc wants to oxidize but it cannot do so if aluminum does, so it is forced to be the reduction side of this redox. It is forced to accept the electrons.**

The  $\text{Zn}^{+2}$  ions in solution are attracted to all the electrons on the zinc bar, so they jump onto the bar, picking up electrons, and get reduced into zinc atoms.

The addition of  $\text{Al}^{+3}$  cations to the left solution creates a + charge in the solution. The removal of  $\text{Zn}^{+2}$  ions from the other solution makes a - charge in that solution. This is a PROBLEM, big enough to stop the electricity flow immediately.

That's why we have an aqueous salt bridge. This one has  $\text{Na}^{+1}$  and  $\text{Cl}^{-1}$  ions. As the solution at left becomes more + charged, the chloride anions flow into it, which offsets that electric potential. As the solution at right gets more negatively charged, the sodium cations flow to it, offsetting that tendency. It's this ion flow that allows the electrons to keep transferring, or allows this REDOX to continue.



Each metal bar is an electrode. The names are cathode and anode. The way to keep them straight is to remember Leo the lion. He's a RED CAT. That reminds you that reduction happens on the cathode. Oxidation happens on the anode.

So, aluminum is the anode, because the zinc is the cathode.

### 3 reasons that batteries die...

1. Run out of anode. Sooner or later the aluminum atoms will all give up their electrons and jump into solution, and there will be none left.
2. Run out of salt bridge ions, sooner or later the ions will all move to opposite sides due to the change in electrical potential due to electron transfer.
3. Run out of cathode side cations. If enough electrons arrive on the zinc bar, all the zinc cations will become zinc atoms. At that point no more reduction can occur.

Let's quickly look at the REDOX of this battery.



Here it seems that the number of electrons is not balanced, that's because the half reactions set up is not balanced. The actual balanced half reactions are



That's the complete redox reaction. Always make sure your electrons balance, and if they don't, then balance your half reactions properly (like you'd balance your chemical equations).

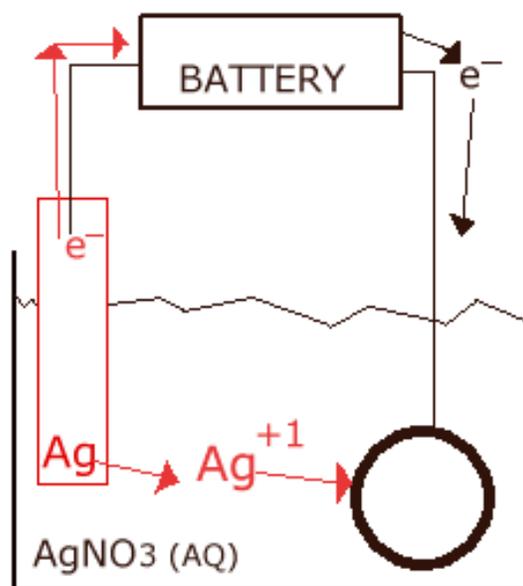
## ***Electrolytic cells***

In these cells, we use electricity to force redox, and we can use this to electroplate valuable metals onto less valuable metals. This covers our cheap filler with precious metals, so rings look great even though they might be inexpensive. We can use copper just coated with silver, for example.

The electrons provided by the battery jump onto the copper ring. The silver ions in solution are attracted and get reduced to silver atoms, which plate the copper.

To complete the circuit silver atoms in the bar oxidize and run up the wire to the other side of the battery.

***We are forcing the more active metal, which prefers to donate its electrons to reach a stable electron configuration, to accept electrons from the less active metal.***



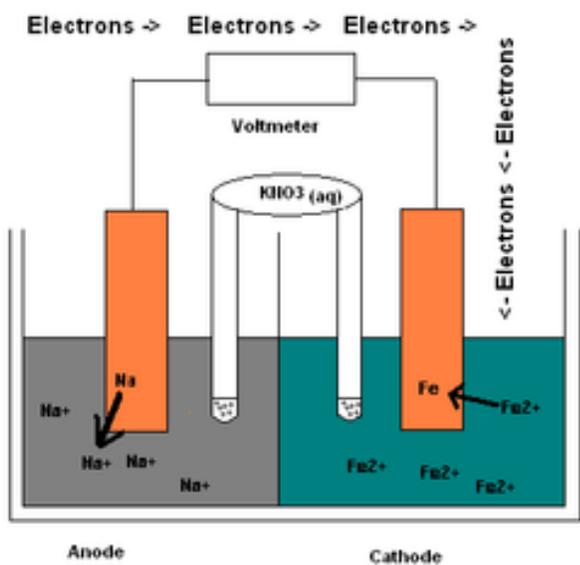
The oxidation half reaction is  $\text{Ag}^{\circ} \longrightarrow \text{Ag}^{+1} + 1\text{e}^{-}$

The reduction half reaction is  $\text{Ag}^{+1} + 1\text{e}^{-} \longrightarrow \text{Ag}^{\circ}$

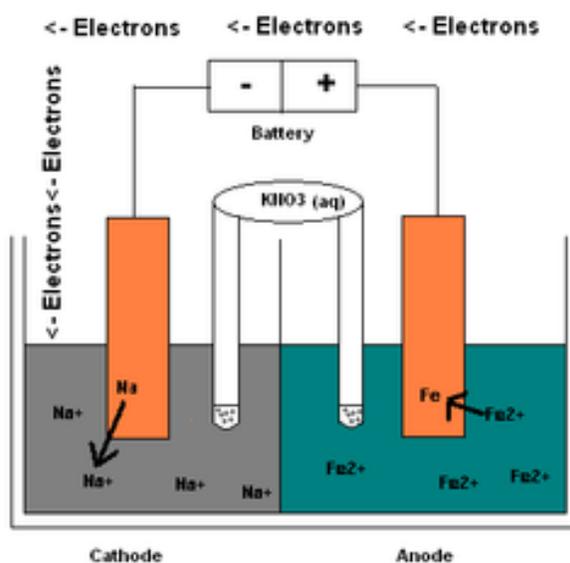
Electroplating is a billion dollar a year business. It is possible due to chemistry.

This is not a spontaneous redox, it requires energy, and that energy makes chemistry happen.

### **Voltaic**



### **Electrolytic**



[Voltaic cell Animation](#)