

## Electricity

<b>Word</b>	<b>Definition</b>
Alternating Current (AC)	The current produced by power plants, where the polarity of the electrical current alternates at the rate of 60 Hz (50 Hz in Europe).
Anode	The electrode at which oxidation occurs.
Cathode	The electrode at which reduction occurs.
Converter	A device that takes AC wall current at 120 volts and converts it to DC current at the voltage required by the electrical device the converter is plugged into.
Direct Current (DC)	The current produced by batteries, where the electricity flows only from anode to cathode. This is the current type used by electronic devices.
Electrolysis	Splitting apart the elements in a compound with the application of electricity.
Electrolytic cell	A cell that uses electricity from an outside source to force a nonspontaneous redox reaction to occur. Examples of these include recharging a battery, electrolytic decomposition of binary compounds and electroplating.
Electroplating	An electrolytic process that involved oxidizing a source metal into solution with the use of an external power source and then reducing the metal ion onto a metallic object that is to be plated.
Half-reaction	A reaction that describes the change in oxidation number and subsequent gain or loss of electrons that occurs during oxidation or reduction.
Load	The device that makes use of the electric current produced by an electrochemical cell.
Oxidation number	The charge of an ion or the apparent charge of a nonmetal atom in a covalent bond.
Oxidizing agent	The species that is reduced and therefore removes the electrons from the species that was oxidized.
Reducing agent	The species that is oxidized and therefore gives electrons to the species that was reduced.
Salt bridge	A semi-permeable barrier that allows the flow of ions from one half-cell to each other, but prevents the direct mixing of those ions.
Species	The symbol and charge of an element or ion in a redox reaction.
Transformer	An electrical device that steps voltage up or down.
Voltaic cell	An electrochemical cell that produces electric current as a result of a spontaneous redox reaction, used to make batteries. Consists of two half-cells connected by a salt bridge and two electrodes that connect to a load that uses the electricity produced by the cell.

## 1) Half-Reactions (HW: p. 27-29)

**Essential Question:** What is electricity and how is it formed?

In this unit, you will be introduced to the amazing world of electricity and how it can be both generated by and used by chemical reactions. You will also learn how solar, wind, tide, fossil fuel power plants and fuel cells work.

The human species is amazing in its ingenuity. We discovered how to make fire, how to control our environment despite our physical weaknesses. We developed civilizations, asked the big questions and worked to get answers to them. As time went on, our understanding of the universe and our place in it was refined, and with better science came better technology to help us shape our environment. Energy was the big obstacle to overcome. The Pyramids were built using basic machines powered by human slaves. Fields were plowed using machines pulled by oxen. People were carried from place to place using horses, camels or litters borne up by human power.

It was discovered in the 1800's that steam could be used to generate power that could be used for transportation, and the steam engine was born. Railways sprang up overnight. Then the internal combustion engine was invented, and cars multiplied like bacteria on the rut-strewn carriage roads. Then those roads were paved, and automobiles were joined by motorcycles, ATV's and scooters. The engine was used to power vehicles that fly, as well, as the Wright brothers pushed off from Kitty Hawk in their brand new airplane. War spurred innovation, and World War I was fought using air power. These developments continued, and our reliance on chemical power grew.

Houses used to be lighted by candles or oil lamps. These burned fuels to produce light, but they were a fire hazard and they produced smoky residues. Thomas Edison played with getting a thin filament of material to glow using the power of electricity, eventually leading to the commercialization of the light bulb, but to the very power grid and other electric appliances that filled homes from the 1920's onwards. Toasters, refrigerators, electric stoves, radios, televisions and even weed whackers...all made possible by this new source of power, electricity!

Eventually, we became consumed with electric-powered devices. Toys, which were once used by children to role-play and develop social and cognitive skills they would need as adults, began shipping with features that used electricity to bring the toy alive! People wanted to listen to their radio wherever they were, not just by the fireside. This, coupled with the invention of the transistor, spurred a revolution in portable electronics. Using chemical cells called "batteries", these devices could run off of a portable source of electricity no matter where you were! Flashlights gave greater freedom to move about at night, transistor radios were used by millions to listen to music, news and sports outside the home, and that was just the beginning! With the development of the computer microchip, new portable electronic devices were invented. One of the most important inventions in the history of portable electronics was the Sony Walkman. This was a tape cassette player, barely larger than the tape itself, which could play back your music anywhere you were. You could create mix tapes by recording whatever music you wanted from the radio or record album, and take it with you! The battery had important implications in the field of personal communication, as well. First the pager was invented, that could flash a pre-programmed message or phone number on what was basically a calculator-style screen. That led to the invention of the cell phone, the Smartphone and the iPhone. The Walkman gave way to the mp3 player, the Diamond Rio. Then Apple got in on the game and quickly rose to dominate the portable media market. Pocket calculators were invented in the 1970's, and have led to monster pocket computing powerhouses that do any kind of math you might want to throw at them. Home computers became very popular in the 1980's (the Commodore 64 paving the way for future devices), and they quickly left the desk and became the laptop computers of today. Even vehicles are starting to get in on the battery game...as the price of oil continues its long trend upwards (and its availability in the world trends downwards), vehicles will increasingly rely on electricity as their main power source. Electric motors are quiet, and release no pollution of their own. Within a few years, cars that simply recharge at night by plugging them into your wall socket at home will become commonplace, and will be eagerly purchased by commuters that can't afford the price of gasoline. And don't forget the huge market of portable power tools, drills, saws and other devices that operate on a battery pack!

So how does this little device, this battery that powers our mobile world, work? What magic makes it possible to get power from a little box or cylinder marked "AA"? What do new technologies, like the lithium-polymer cell phone battery, have over the old zinc/manganese based batteries that you throw out after each use? And what other amazing uses can we put electricity to? The answers to these and many other questions will be revealed in this unit. Prepare yourselves...for the portable power of ELECTROCHEMISTRY.

Before you can learn how batteries work, here's a refresher on how elements gain and lose electrons, and the types of reactions that we will be working with in this unit!

### Quick Periodic Table Review

Type of Element	Metal	Nonmetal
<b>Electronegativity (and what it means)</b>	Metals have LOW electronegativity, which means they do not attract electrons strongly in a chemical bond.	Nonmetals have HIGH electronegativity, which means they attract electrons strongly in a chemical bond.
<b>Ionization Energy (and what it means)</b>	Metals have LOW ionization energy, which means they can easily lose electrons when energy is added.	Nonmetals have HIGH ionization energy, which means they hold onto their valence electrons very strongly.
<b>How ions are formed</b>	Metal atoms LOSE their valence electrons (OXIDATION) to form POSITIVE charged ions (CATIONS).	Nonmetal atoms GAIN enough electrons (REDUCTION) to give themselves a stable octet of 8 valence electrons to form NEGATIVE charged ions (ANIONS).
<b>What happens to the atomic radius when ion is formed</b>	The radius of the metal atom DECREASES as the metal atom loses its valence electrons, meaning the ionic radius of a metal ion is SMALLER than the original atomic radius was.	The radius of the nonmetal atom INCREASES as the nonmetal atom gains valence electrons, meaning the ionic radius of a nonmetal ion is LARGER than the original atomic radius was.

### Quick Redox Reaction Type Review

Redox reaction type	General Formula	Example	Uses In This Unit
<b>Synthesis</b>	$A + B \rightarrow AB$	$2 H_2 + O_2 \rightarrow 2 H_2O$	This reaction can be used to create electricity in a FUEL CELL battery. The electrons lost by $H_2$ go through a wire into the device to be powered, then into to $O_2$ , where they are gained. It's called a fuel cell, because you recharge it by adding more $H_2$ to the cell.
<b>Decomposition</b>	$AB \rightarrow A + B$	$2 NaCl \rightarrow 2 Na + Cl_2$	This reaction is carried out by ADDING ELECTRICITY to the liquid form of NaCl. The $Na^{+1}$ is forced to gain electrons and reduce to $Na^0$ , and the $Cl^{-1}$ is forced to lose electrons and form $Cl_2^0$ . This is called ELECTROLYTIC DECOMPOSITION.
<b>Single Replacement</b>	$A + BC \rightarrow AC + B$	$Zn + Cu(NO_3)_2 \rightarrow Zn(NO_3)_2 + Cu$	This reaction is used to make VOLTAIC CELLS (which, when more than one are connected together, is called a "battery"), our source of portable electricity. This example is called a zinc/copper cell, and it generates 1.10 volts of electricity.

## Redox reactions used for electrochemistry are driven by a change in charge.

<b>OXIDATION</b>	Loss of electrons	Becomes more + charged
<b>REDUCTION</b>	Gain of electrons	Becomes more - charged

### Rules For Determining Oxidation Numbers

The first step in determining which species was oxidized and which was reduced is knowing what the charge of each species starts at and ends at, so you can determine whether electrons were lost or gained. Why is that important? Ever put a battery in backwards? It doesn't work! Circuits are designed to only let electrons through ONE way. It is possible to destroy an electronic device by putting the battery in backwards. We need to know which end is - (the end that loses electrons, oxidation) and which end is + (the end that gains electrons, reduction), so we need to know how to find the charges (oxidation numbers) of the species involved!

**The sum of all charges in a compound equals zero.**

**The sum of all charges in a polyatomic ion equals the charge of the polyatomic ion.**

#### **1) Uncombined elements have a charge of ZERO.**

Na has a charge of zero  
O<sub>2</sub> has a charge of zero  
C<sub>60</sub> has a charge of zero

#### **2) If an element has only one charge listed on the periodic table, then that is its oxidation number.**

Na has one charge, +1.  
Ca has one charge, +2  
Al has one charge, +3

#### **3) If a nonmetal atom is the negative ion in an ionic compound, then the top charge listed is its oxidation number.**

In NaCl, Cl is the negative ion, so its charge is -1.  
In Li<sub>3</sub>N, N is the negative ion, so its charge is -3.  
In ZnS, S is the negative ion, so its charge is -2.

#### **4) If an element has more than one charge listed, use the other charges to figure it out.**

ZnSO<sub>4</sub>: Zn has a charge of +2, S has three charges listed (-2, +4 and +6) and O has a charge of -2. What is S's charge?

There is one Zn with a charge of +2, so Zn's contribution to the compound's charge is +2.  
There are four O's each with a charge of -2, so O's contribution to the compound's charge is -8  
A compound has a net charge of zero, so (+2) + (X) + (-8) = 0, so X = +6.  
There is only one S, so its charge is +6.

Fe(NO<sub>3</sub>)<sub>2</sub>: Fe has two charges listed (2 and +3), N has eight charges listed (-3, -2, -1, +1, +2, +3, +4 and +5) and O has one charge listed (-2). We have to find the charges of Fe and N.

Use the fact that the nitrate ion (NO<sub>3</sub>) is -1, according to Reference Table E. There are 2 nitrate ions, for a total charge of -2, so Fe must be +2 to cancel it out.

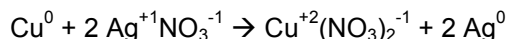
There is one Fe with a charge of +2, so Fe's total contribution to the compound's charge is +2.  
There are six O's, each with a charge of -2 for a total contribution of -12.  
A compound has a net charge of zero, so (+2) + (X) + (-12) = 0, so X = +10.  
There are two N's, so the +10 is distributed among them. Each N is +5.

## Identifying which species is oxidized and which is reduced

If the charge becomes more positive going from left to right, the species is oxidized and loses electrons.  
If the charge becomes more negative going from left to right, the species is reduced and gains electrons.

For the reaction  $\text{Cu} + 2 \text{AgNO}_3 \rightarrow \text{Cu}(\text{NO}_3)_2 + 2 \text{Ag}$ :

First determine the oxidation numbers for each species. If a polyatomic ion remains constant on both sides, just look up the charge of the polyatomic ion. Since we have  $\text{NO}_3$  on the left and the right, we will just look up the charge of  $\text{NO}_3$ .

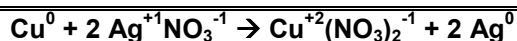


Cu goes from 0 to +2, so  **$\text{Cu}^0$  is oxidized** by losing two electrons.  
Ag goes from +1 to 0, so  **$\text{Ag}^{+1}$  is reduced** by gaining one electron.

**Agents:** The species that is oxidized is called the reducing agent, the species that is reduced is called the oxidizing agent.

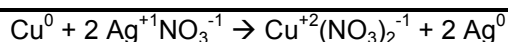
**Oxidizing agent:** the species that causes the oxidation. Oxidation is a loss of electrons, so the oxidizing agent is the species that removes the electrons from the oxidized species. The reduced species gains electrons from the oxidized species. The reduced species causes the oxidized species to be oxidized, so it is the agent of its oxidation.

**Reducing agent:** the species that causes the reduction. Reduction is a gain of electrons, so the reducing agent is the species that gives electrons to the reduced species. The oxidized species loses electrons to the reduced species. The oxidized species causes the reduced species to be reduced, so it is the agent of its reduction.



Since  **$\text{Cu}^0$  is oxidized, it is the reducing agent.** It causes  $\text{Ag}^{+1}$  to be reduced because  $\text{Cu}^0$  gives electrons to  $\text{Ag}^{+1}$ .  
Since  **$\text{Ag}^{+1}$  is reduced, it is the oxidizing agent.** It causes  $\text{Cu}^0$  to be oxidized because  $\text{Ag}^{+1}$  takes electrons from  $\text{Cu}^0$ .

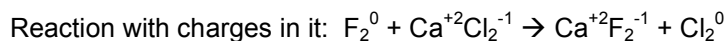
**Spectator Ions:** The ion that does not change its charge.



The  $\text{NO}_3^{-1}$  remains the same charge on both sides, and is therefore the spectator ion.

### PUTTING IT TOGETHER:

For the reaction  $\text{F}_2 + \text{CaCl}_2 \rightarrow \text{CaF}_2 + \text{Cl}_2$ :



Oxidized:  $\text{Cl}^{-1}$  (because it goes from  $-1$  to  $0$ )

Reduced:  $\text{F}_2^0$  (because it goes from  $0$  to  $-1$ )

Oxidizing Agent:  $\text{F}_2^0$

Reducing Agent:  $\text{Cl}^{-1}$

Spectator Ion:  $\text{Ca}^{+2}$

## Half-Reactions

A redox reaction may be split into 2 **half-reactions**, one signifying oxidation and the other signifying reduction.

**For the reaction  $\text{Cu} + 2 \text{AgNO}_3 \rightarrow \text{Cu}(\text{NO}_3)_2 + 2 \text{Ag}$ :**

1) Write the charges:  $\text{Cu}^0 + 2 \text{Ag}^{+1}\text{NO}_3^{-1} \rightarrow \text{Cu}^{+2}(\text{NO}_3)_2^{-1} + 2 \text{Ag}^0$

2) **Oxidation Half-Reaction:**  $\text{Cu}^0 \rightarrow \text{Cu}^{+2} + 2\text{e}^-$ . What this says is that Cu goes from 0 to +2 by losing 2 electrons. The electrons are put on the products side to a) show a loss and b) satisfy the Law of Conservation of Charge.

3) **Reduction Half-Reaction:**  $\text{Ag}^{+1} + 1\text{e}^- \rightarrow \text{Ag}^0$ . What this says is that Ag goes from +1 to 0 by gaining an electron. The electron is put on the reactants side to a) show a gain and b) satisfy the Law of Conservation of Charge.

**For the reaction  $\text{F}_2 + \text{CaCl}_2 \rightarrow \text{CaF}_2 + \text{Cl}_2$ :**

1) Write the charges:  $\text{F}_2^0 + \text{Ca}^{+2}\text{Cl}_2^{-1} \rightarrow \text{Ca}^{+2}\text{F}_2^{-1} + \text{Cl}_2^0$

2) **Oxidation Half-Reaction:**  $\text{Cl}^{-1} \rightarrow \text{Cl}^0 + \text{e}^-$ . Notice that  $\text{Cl}^0$  is actually diatomic,  $\text{Cl}_2^0$ . Writing half-reactions with diatomic species requires that it be written in diatomic form, with the rest of the half-reaction balanced:  
 $2 \text{Cl}^{-1} \rightarrow \text{Cl}_2^0 + 2\text{e}^-$ . What this means is that two chloride ions form diatomic chlorine by losing two electrons.

3) **Reduction Half-Reaction:** Note that  $\text{F}_2^0$  is diatomic. Therefore the half-reaction is  $\text{F}_2^0 + 2\text{e}^- \rightarrow 2 \text{F}^{-1}$ .

**For the reaction  $\text{N}_2 + 3 \text{H}_2 \rightarrow 2 \text{H}_3\text{N}$ :**

1) Write the charges:  $\text{N}_2^0 + 3 \text{H}_2^0 \rightarrow 2 \text{H}_3^{+1}\text{N}^{-3}$

2) Oxidation Half-Reaction:  $\text{H}_2^0 \rightarrow 2 \text{H}^{+1} + 2 \text{e}^-$

3) Reduction Half-Reaction:  $\text{N}_2^0 + 6 \text{e}^- \rightarrow 2 \text{N}^{-3}$

**For the reaction  $\text{Zn} + 2 \text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$ :**

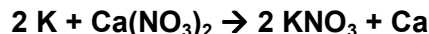
1) Write the charges:  $\text{Zn}^0 + 2 \text{H}^{+1}\text{Cl}^{-1} \rightarrow \text{Zn}^{+2}\text{Cl}_2^{-1} + \text{H}_2^0$

2) Oxidation Half-Reaction:  $\text{Zn}^0 \rightarrow \text{Zn}^{+2} + 2 \text{e}^-$

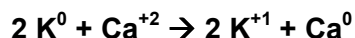
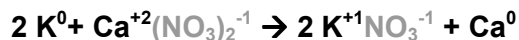
3) Reduction Half-Reaction:  $2 \text{H}^{+1} + 2 \text{e}^- \rightarrow \text{H}_2^0$

## Balancing Net Ionic Redox Reactions Using Half-Reactions

Because it does not matter what spectator ions are used in a redox reaction, they can be omitted to form the “net ionic reaction”. The spectator ions are still there, but are simply ignored when writing the reaction. This makes it easier to see exactly which species is being oxidized and which is being reduced.



Since the nitrate ion,  $\text{NO}_3^{-1}$ , is the spectator ion in this reaction, the reaction can be rewritten as a net ionic reaction:



Notice how not only mass is conserved (same number of K's and Ca's on both sides), but charge is also conserved (total of +2 on each side).

### **Balancing Net Ionic Reactions:**

The number of electrons lost by the species being oxidized must equal the number of electrons gained by the species being reduced.

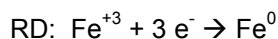
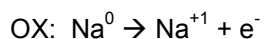
The difference in charge between both sides of the oxidized species will be the coefficient of the reduced species.

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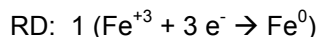
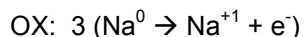
### **Balance the net ionic redox reaction $\text{Na}^0 + \text{Fe}^{+3} \rightarrow \text{Na}^{+1} + \text{Fe}^0$**

Note that the net charge on the left side is +3, but the net charge on the right side is only +1.

1) Write the half-reactions:



2) Multiply the half-reactions by the number of electrons in the other half-reaction:



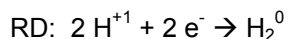
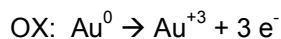
3) Multiply through and put the resulting coefficients into the net ionic reaction:



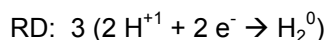
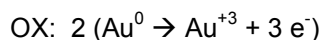
### **Balance the net ionic redox reaction $\text{Au}^0 + \text{H}^{+1} \rightarrow \text{Au}^{+3} + \text{H}_2^0$**

Note that the net charge on the left side is +1, but the net charge on the right side is +2.

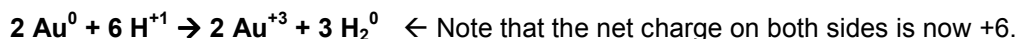
1) Write the half-reactions:



2) Multiply the half-reactions by the number of electrons in the other half-reaction:



3) Multiply through and put the resulting coefficients into the net ionic reaction:



## 2) Electrochemistry (HW: p. 30- 33)

**Essential Question:** What is electricity and how is it formed?

Electrochemistry is the science (and art) of creating portable power cells that are commonly known as batteries. In this section, you will learn the secrets of how batteries work, and how to design them for yourself!

### Redox Potentials and Metal Activity

#### Reference Table J: Activity Series

The higher up the table a substance is, the **MORE** reactive it is.

**Table J**  
**Activity Series\*\***

Most	Metals	Nonmetals	Most
	Li	F <sub>2</sub>	
	Rb	Cl <sub>2</sub>	
	K	Br <sub>2</sub>	
	Cs	I <sub>2</sub>	
	Ba		
	Sr		
	Ca		
	Na		
	Mg		
	Al		
	Ti		
	Mn		
	Zn		
	Cr		
	Fe		
	Co		
	Ni		
	Sn		
	Pb		
	**H <sub>2</sub>		
	Cu		
	Ag		
	Au		
Least			Least

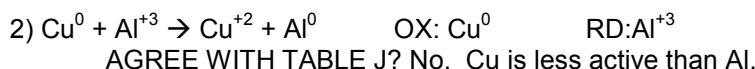
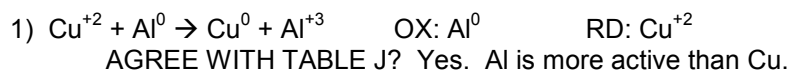
\*\*Activity Series based on hydrogen standard

Note: H<sub>2</sub> is not a metal

**For metals, more reactive means a higher potential to be oxidized.**

**Which metal has the greatest potential to be oxidized?** Li, lithium. It is the most active metal, which means it has the greatest potential to lose electrons. Batteries (cells) that contain lithium (lithium camera batteries, lithium coin cells for calculators, lithium ion batteries for your cell phone) pack a LOT of voltage in a compact, lightweight space. Lithium ion and lithium polymer batteries are also rechargeable, unlike disposable one-use zinc/manganese batteries!

**Which reaction is more likely to occur:**

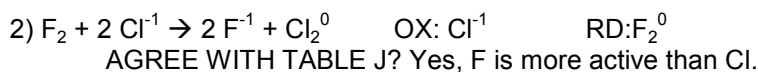
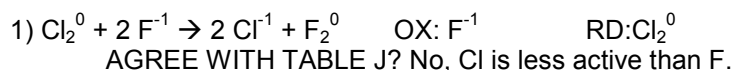


**Therefore, reaction 1 will be the one that takes place.**

**For nonmetals, more reactive means a higher reduction potential.**

**Which nonmetal has the greatest potential to be reduced?** F<sub>2</sub>, fluorine. It is the most active nonmetal, which means it has the greatest potential to gain electrons.

Which reaction is more likely to occur:



**Therefore, reaction 2 will be the one that takes place.**



## Determining Spontaneity of Redox Reactions (Voltage)

In order for a redox reaction to be spontaneous, there has to be a tendency for one species to gain electrons more than another. The tendency for a species to gain electrons in a redox reaction is measured in volts, the basic unit of electromotive force (emf). The voltage you will calculate is  $E^0$ , or the initial voltage of the cell. As the cell is used, the voltage will decrease as the reactants are used up.

Electronic devices are designed to run on a certain voltage. Batteries with a higher voltage than is needed are used, and a voltage regulator in the device limits the voltage allowed to pass through it. As the batteries are used up, they will eventually reach the minimum voltage needed to keep the device operating. We think that the battery is "dead", but in fact, it still has electromotive force...just not enough to keep the device running, so the device shuts down.

**AE Reference Table F (next page) can be used to determine the initial voltage ( $E^0$ ) of a redox reaction:**

### Steps:

1) Write the oxidation and reduction half-reactions.

**The voltages listed on Table F are REDUCTION POTENTIALS. Instead of having to come up with another table for oxidation potentials, all you have to do is reverse the sign on the reduction potential if you need the oxidation potential.**

2) Assign the reduction half-reaction the voltage listed on the chart. Species being reduced is on the left.

3) Assign the oxidation half-reaction the negative of the voltage listed on the chart. Species being oxidized is on the right.

From the chart:

Half-Reactions	Reduction Potential, $E^0$ (volts)
$\text{Cu}^{+2} + 2 \text{e}^- \rightarrow \text{Cu (s)}$	+0.34

**Therefore, for the oxidation half-reaction  $\text{Cu (s)} \rightarrow \text{Cu}^{+2} + 2\text{e}^-$ , the voltage would be -0.34 volts.**

4) Add the two voltages together. This gives the net voltage, or potential difference of the reaction.

5) If the voltage is positive, the reaction is spontaneous. If the voltage is negative, the reaction is nonspontaneous.

What is the potential difference of the reaction  $\text{Zn} + \text{Cu}(\text{NO}_3)_2 \rightarrow \text{Zn}(\text{NO}_3)_2 + \text{Cu}$ ?

With charges:  $\text{Zn}^0 + \text{Cu}^{+2}(\text{NO}_3)_2^{-1} \rightarrow \text{Zn}^{+2}(\text{NO}_3)_2^{-1} + \text{Cu}^0$

OX:  $\text{Zn}^0 \rightarrow \text{Zn}^{+2} + 2\text{e}^-$  (reduction potential is  $-0.76$  volts, so reverse it to get oxidation potential) =  $+0.76$  v

RD:  $\text{Cu}^{+2} + 2\text{e}^- \rightarrow \text{Cu}^0$  (reduction potential is listed on the table as) =  $+0.34$  v

$(+0.76 \text{ v}) + (+0.34 \text{ v}) = \mathbf{+1.10 \text{ volts}}$ . Positive voltage means the reaction is **SPONTANEOUS**.

What is the potential difference of the reaction  $2 \text{NaCl} \rightarrow 2 \text{Na} + \text{Cl}_2$ ?

With charges:  $2 \text{Na}^{+1}\text{Cl}^{-1} \rightarrow 2 \text{Na}^0 + \text{Cl}_2^0$

OX:  $2 \text{Cl}^{-1} \rightarrow \text{Cl}_2^0 + 2 \text{e}^-$  (red. potential is  $+1.36$  volts, so reverse it to get oxidation potential) =  $-1.36$  v

RD:  $\text{Na}^{+1} + 1 \text{e}^- \rightarrow \text{Na}^0$  (reduction potential is listed on the table as) =  $-2.71$  v

$(-1.36 \text{ v}) + (-2.71 \text{ v}) = \mathbf{-4.07 \text{ volts}}$ . Negative voltage means the reaction is **NONSPONTANEOUS**.

## AE Reference Table F

### Standard Reduction Potentials (Ionic Concentrations 1 M in Water at 298 K and 1 atm)

(Adapted from New York State Regents Chemistry Reference Table N, 1987 Revision)

Half-Reactions	Reduction Potential, $E^0$ (volts)
$F_2(g) + 2e^- \rightarrow 2F^{-1}$	+2.87
$Au^{+3} + 3e^- \rightarrow Au(s)$	+1.50
$Cl_2(g) + 2e^- \rightarrow 2Cl^{-1}$	+1.36
$Br_2(l) + 2e^- \rightarrow 2Br^{-1}$	+1.09
$Hg^{+2} + 2e^- \rightarrow Hg(l)$	+0.85
$Ag^{+1} + 1e^- \rightarrow Ag(s)$	+0.80
$Fe^{+3} + 1e^- \rightarrow Fe^{+2}$	+0.77
$I_2(s) + 2e^- \rightarrow 2I^{-1}$	+0.54
$Cu^{+1} + 1e^- \rightarrow Cu(s)$	+0.52
$Cu^{+2} + 2e^- \rightarrow Cu(s)$	+0.34
$Sn^{+4} + 2e^- \rightarrow Sn^{+2}$	+0.15
$2H^{+1} + 2e^- \rightarrow H_2(g)$	0.00
$Pb^{+2} + 2e^- \rightarrow Pb(s)$	-0.13
$Sn^{+2} + 2e^- \rightarrow Sn(s)$	-0.14
$Ni^{+2} + 2e^- \rightarrow Ni(s)$	-0.26
$Co^{+2} + 2e^- \rightarrow Co(s)$	-0.28
$Fe^{+2} + 2e^- \rightarrow Fe(s)$	-0.45
$Cr^{+3} + 3e^- \rightarrow Cr(s)$	-0.74
$Zn^{+2} + 2e^- \rightarrow Zn(s)$	-0.76
$Mn^{+2} + 2e^- \rightarrow Mn(s)$	-1.19
$Al^{+3} + 3e^- \rightarrow Al(s)$	-1.66
$Mg^{+2} + 2e^- \rightarrow Mg(s)$	-2.37
$Na^{+1} + 1e^- \rightarrow Na(s)$	-2.71
$Ca^{+2} + 2e^- \rightarrow Ca(s)$	-2.87
$Sr^{+2} + 2e^- \rightarrow Sr(s)$	-2.89
$Ba^{+2} + 2e^- \rightarrow Ba(s)$	-2.91
$Cs^{+1} + 1e^- \rightarrow Cs(s)$	-2.92
$K^{+1} + 1e^- \rightarrow K(s)$	-2.93
$Rb^{+1} + 1e^- \rightarrow Rb(s)$	-2.98
$Li^{+1} + 1e^- \rightarrow Li(s)$	-3.04

These are reduction potentials. To find oxidation potentials, reverse the half reaction and the sign for the voltage. For the reduction of  $Zn^{+2}$ , the potential is  $-0.76$ , so for the oxidation of  $Zn^0$ , the potential is  $+0.76$ .

## Using the Potential to Create Electricity

Voltaic Cell - a redox reaction whose two half-reactions are carried out separately, and the electrons given off by the oxidation half-reaction are used to power a device, and then given to the reduction half-reaction.  
ALSO CALLED A BATTERY (9v) OR CELL (AAA, AA, C, D)

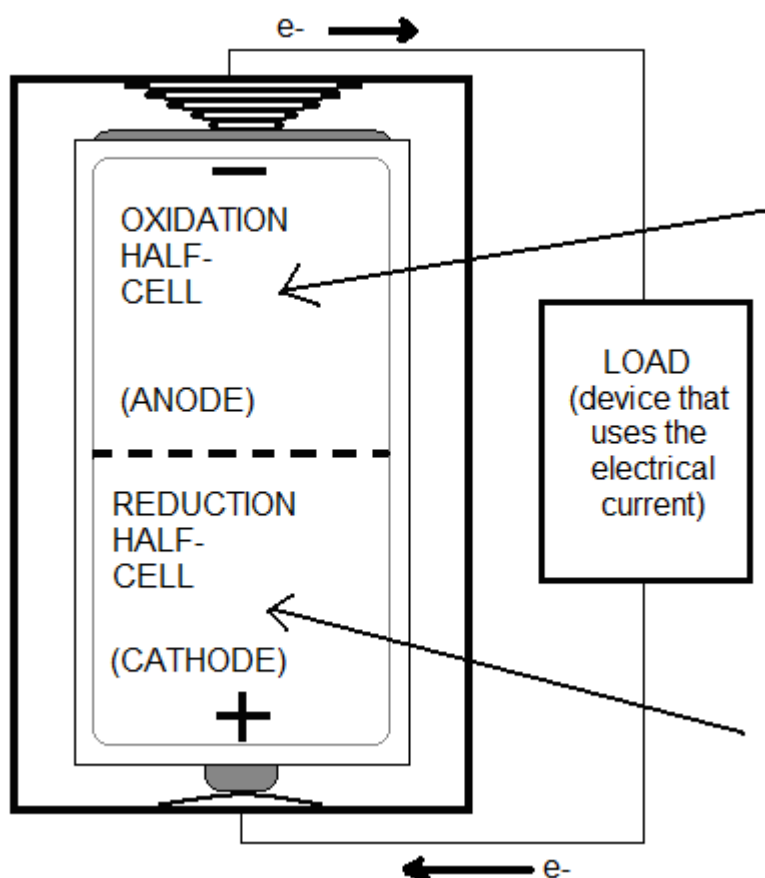
### CONSTRUCTION OF A VOLTAIC CELL

- 1) Each half-reaction is carried out in a half-cell. In these cells, the uncombined free element acts as the electrode which is immersed in a solution of the ion of that element.
- 2) These half-cells are connected at their electrodes by a wire that is used to transport the electrons from the oxidized electrode to the reduced electrode.
- 3) Hooked up to the wire between the two half-cells is the load, the device that is running off of the current produced by the difference in oxidation potential and reduction potential of the two electrodes.
- 4) Connecting the half-cells directly is a salt bridge that is used to complete the circuit. It is made of a porous substance that contains a salt that contains a cation different from the electrode cations.

### LABELING AN ELECTROCHEMICAL CELL

- 1) The Oxidation Half-Cell is called the anode (mnemonic: AN OX).
- 2) The Reduction Half-Cell is called the cathode (mnemonic: RED CAT).
- 3) Electrons travel from the anode to the cathode (from loss to gain, from oxidation to reduction) across the wire.
- 4) Anions travel across the salt bridge from the cathode to the anode.

### TYPICAL DRY CELL SIMPLIFIED SCHEMATIC (AAA, AA, C and D 1.50-volt cells)



The metal in the oxidation half-cell ( $Zn^0$  in your typical battery) LOSES electrons (oxidizes):



The electrons pass into the device out the - end of the cell. The amount of  $Zn^0$  in the cell decreases over time, eventually getting used up. When all of the  $Zn^0$  is used up, the battery dies.

The electrons pass through the device that is using the electron flow to power it. This device is referred to in electrical terms as the LOAD. The circuit in the load is designed so that electrons can only pass ONE way through it. A diode is usually placed at the end of the circuit on the + side, preventing electrons from coming in that end, so that if you plug the battery in backwards, the circuit won't be destroyed.

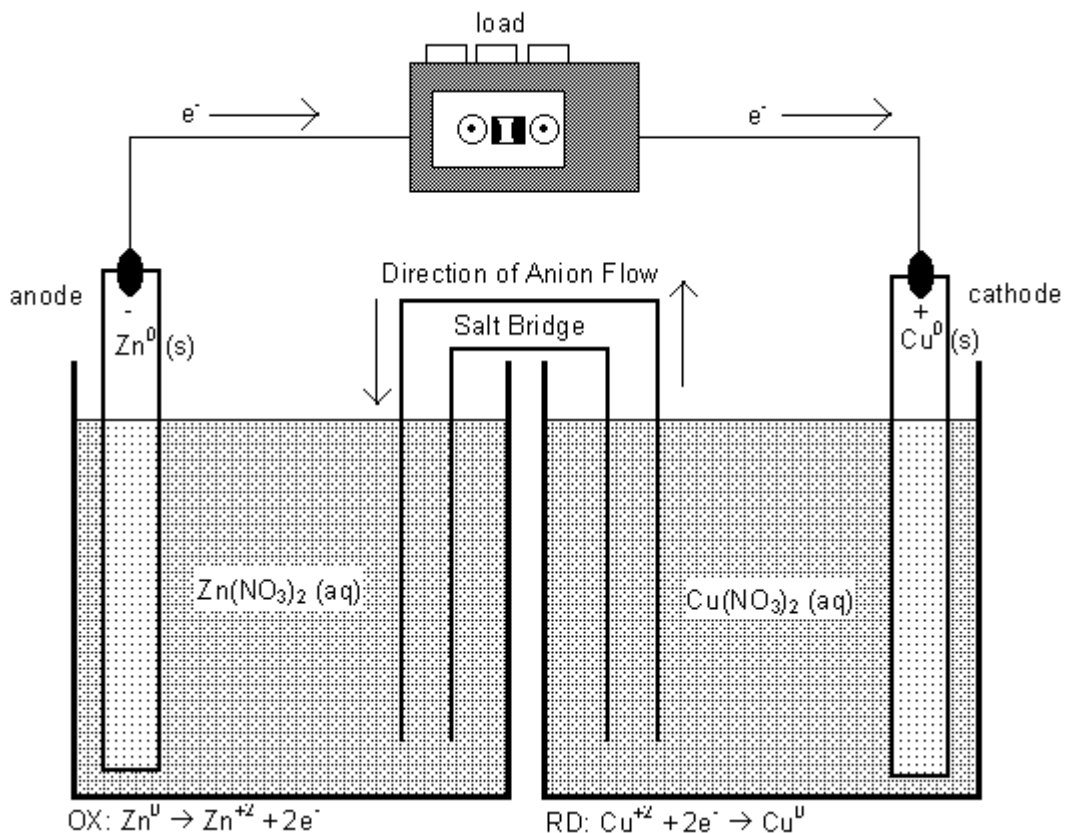
The metal ion in the reduction half-cell ( $Mn^{+4}$  in your typical battery) GAINS electrons (reduces):



The electrons pass from the load into the + end of the cell, where the  $Mn^{+4}$  ions pick them up and reduce to  $Mn^{+2}$ . When all of the  $Mn^{+4}$  is used up, the battery dies.

## TYPICAL WET CELL SCHEMATIC

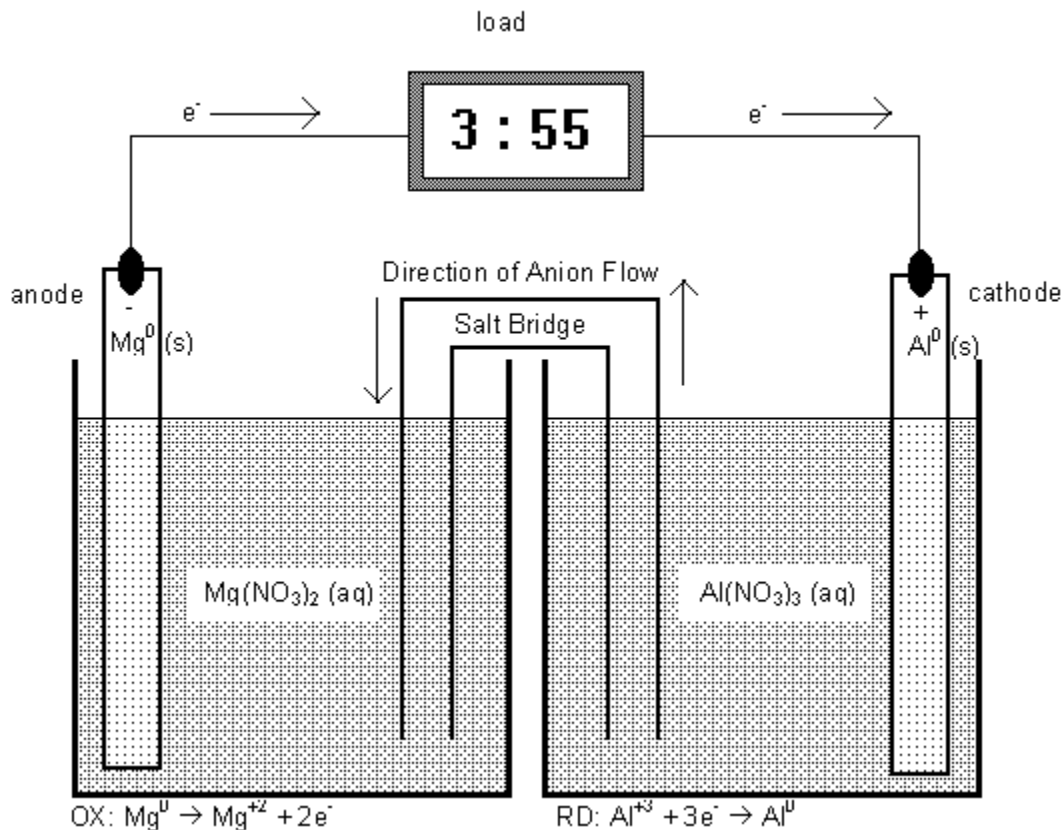
Diagram of Daniell Cell ( $\text{Zn} + \text{Cu}(\text{NO}_3)_2 \rightarrow \text{Zn}(\text{NO}_3)_2 + \text{Cu}$ )



### What's happening during this process?

- 1) The  $\text{Zn}^0$  metal, which is the anode (oxidation electrode, - terminal of the cell) is losing electrons, which the  $\text{Cu}^{+2}$  ions in solution in the reduction half-cell are exerting a pull on these electrons through the wire.
- 2) The  $\text{Zn}^0$  forms  $\text{Zn}^{+2}$ , and dissolves into the solution. The anode becomes smaller over time as it dissolves.
- 3) The electrons pass through the load, powering the device, then out the other side to the cathode ( $\text{Cu}^0$  metal electrode).
- 4) The  $\text{Cu}^0$  has no use for electrons, so they line the surface of the copper metal electrode.
- 5) The  $\text{Cu}^{+2}$  ions in solution are attracted to these electrons, and as soon as the positively charged ions touch the surface of the copper metal, the electrons jump onto the ions, causing  $\text{Cu}^{+2}$  ions to be instantly reduced to  $\text{Cu}^0$ . They stick on to the copper electrode when they are reduced. The size of the cathode increases over time as  $\text{Cu}^0$  atoms stick on to it.
- 6) The nitrate ions in solution travel across the salt bridge from the copper side to the zinc side. More  $\text{Zn}^{+2}$  in the oxidation half-cell requires more nitrate ions to cancel out the charge. The concentration of  $\text{Zn}(\text{NO}_3)_2$  in the oxidation half-cell increases. As copper ions leave the solution in the reduction half-cell, they are no longer needed, so they go across the salt bridge.
- 7) The concentration of  $\text{Cu}(\text{NO}_3)_2$  in the reduction half-cell decreases. As the concentrations of  $\text{Zn}^0$  and  $\text{Cu}(\text{NO}_3)_2$  decrease, the potential difference decreases. When either  $\text{Zn}^0$  or  $\text{Cu}(\text{NO}_3)_2$  get used up, the battery dies. It can be recharged by forcing the electrons to go in the opposite direction, stripping them from the Cu and giving them back to  $\text{Zn}^{+2}$ . This can be done by using a recharger in place of the load.

**Diagram of a voltaic cell carrying out the reaction  $3 \text{Mg} + 2 \text{Al}(\text{NO}_3)_3 \rightarrow 3 \text{Mg}(\text{NO}_3)_2 + 2 \text{Al}$**



**What's happening during this process?**

- 1) The  $\text{Mg}^0$  metal, which is the anode (oxidation electrode, - terminal of the cell) is losing its electrons. The  $\text{Al}^{+3}$  ions in solution in the reduction half-cell exerts a pull on these electrons through the wire.
- 2) The  $\text{Mg}^0$  forms  $\text{Mg}^{+2}$ , and dissolves into the solution. The anode becomes smaller over time.
- 3) The electrons pass through the load, powering the device, then out the other side to the cathode ( $\text{Al}^0$  metal electrode).
- 4) The  $\text{Al}^0$  has no use for electrons, so they line the surface of the aluminum metal electrode.
- 5) The  $\text{Al}^{+3}$  ions in solution are attracted to these electrons, and as soon as the positively charged ions touch the surface of the aluminum metal, the electrons jump onto the ions, causing  $\text{Al}^{+3}$  ions to be instantly reduced to  $\text{Al}^0$ . They stick on to the aluminum electrode when they are reduced. The size of the cathode increases over time.
- 6) The nitrate ions in solution travel across the salt bridge from the aluminum side to the magnesium side. More  $\text{Mg}^{+2}$  in the oxidation half-cell requires more nitrate ions to cancel out the charge. The concentration of  $\text{Mg}(\text{NO}_3)_2$  in the oxidation half-cell increases. As aluminum ions leave the solution in the reduction half-cell, they are no longer needed, so they go across the salt bridge.
- 7) The concentration of  $\text{Al}^{+3}$  in the reduction half-cell decreases. As the concentrations of  $\text{Mg}^0$  and  $\text{Al}^{+3}$  decrease, the potential difference decreases.
- 8) When either  $\text{Mg}^0$  or  $\text{Al}^{+3}$  get used up, the battery dies. It can be recharged by forcing the electrons to go in the opposite direction. NOT ALL BATTERIES ARE RECHARGEABLE – recharging non-rechargeable batteries can cause internal resistance, which builds up heat and can cause the battery to expand and crack.

## DESIGNING A WET VOLTAIC CELL

**1) Choose Two Metal Electrodes** (for the sake of this example, let's choose Al and Ag.)

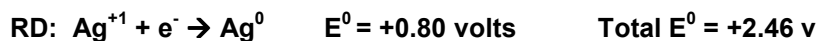
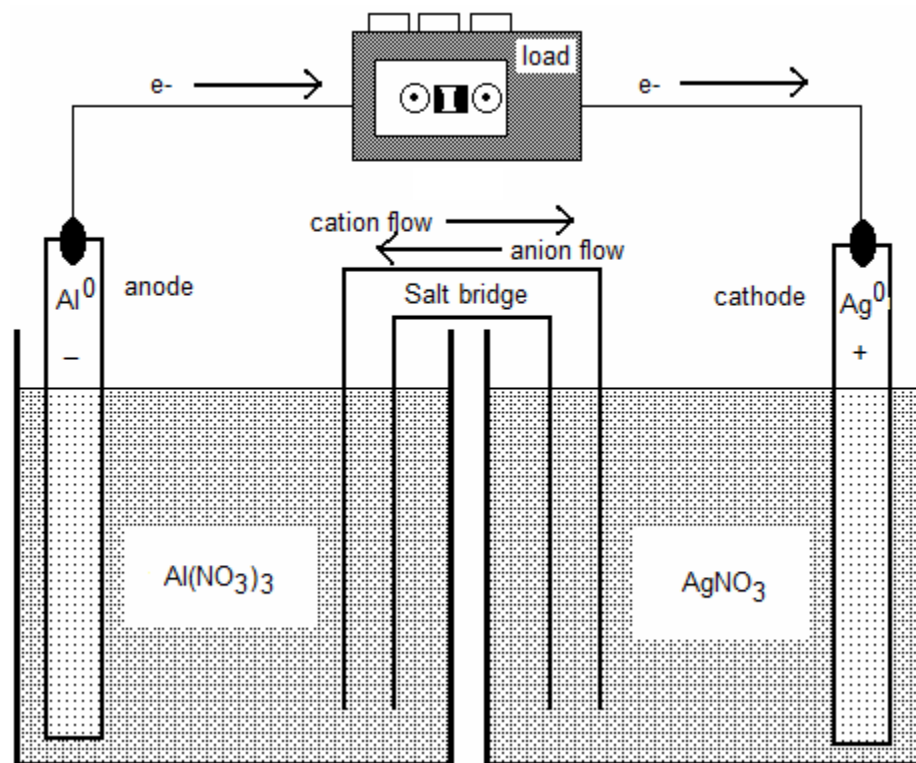
More active metal (- electrode, anode): Al

Less Active Metal (+ electrode, cathode): Ag

**2) Choose the solutions that go with the metals in each half-cell (use rules for formula writing!)**

Solution that goes in the anode half-cell:  $\text{Al}(\text{NO}_3)_3$

Solution that goes in the cathode half-cell:  $\text{AgNO}_3$



1) The oxidized metal (Al, anode) gets smaller as it turns into metal ions and dissolves into the solution. The reduced metal (Ag, cathode) gets larger as it gets coated with the reduced metal ions from solution.

2) The negative ( $\text{NO}_3^{-1}$ ) anions go through the salt bridge from cathode to anode (from the Ag half-cell to the Al half-cell)

3) The positive ( $\text{Al}^{+3}$ ) cations go through the salt bridge from anode to cathode (from the Al half-cell to the Ag half-cell)

4) The salt bridge can contain a solution of any salt, as long as it does not contain Al or Ag.  $\text{NaCl}$  can be used. When the cell is first used,  $\text{Na}^{+1}$  passes into the cathode half-cell, and  $\text{Cl}^{-1}$  passes into the anode half-cell. Later, the ions mentioned in 2) and 3) take their place.

**Table J  
Activity Series\*\***

Most	Metals	Nonmetals	Most
	Li	$\text{F}_2$	
	Rb	$\text{Cl}_2$	
	K	$\text{Br}_2$	
	Cs	$\text{I}_2$	
	Ba		
	Sr		
	Ca		
	Na		
	Mg		
	Al		
	Ti		
	Mn		
	Zn		
	Cr		
	Fe		
	Co		
	Ni		
	Sn		
	Pb		
	** $\text{H}_2$		
	Cu		
	Ag		
Least	Au		Least

\*\*Activity Series based on hydrogen standard

Note:  $\text{H}_2$  is not a metal

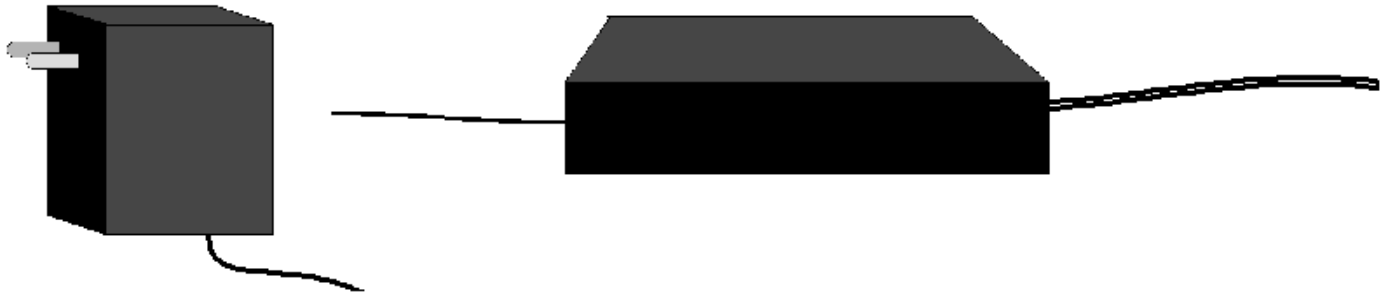
### 3) Other Ways Of Generating Electricity (HW: p. 34)

**Essential Question:** What is "green energy" and how does it work?

The current produced by batteries is called DIRECT CURRENT, or DC. This is because electrons move only in one direction, from anode to cathode, across the wire and through the device (load) that is being powered by the battery.

The electricity generated by power plants is formed when a heat source is used to turn water to high-pressure steam, which pushes turbine fan blades at the rate of 60 rotations per second (50 in Europe), which turns a magnet inside a coil of wire (generator). Changes in the magnetic field induce an electric field in the wire, which creates a flow of electrons out of the wire and out to homes and businesses all over the area.

The electricity made in power plants is ALTERNATING CURRENT, or AC. The electrons reverse their direction of flow 60 times per second (60 Hz), so the + and – poles alternate. The two prongs of a household plug switch off 60 times per second which one is positive and which one is negative. Therefore, if you are going to use a device that runs on batteries on AC current, your device will require an AC to DC converter, which is usually in the form of a "wall wart" block that the prongs are attached to, or an inline "power brick" converter:



This is a "wall wart" DC converter.

This is a power "brick" DC converter.

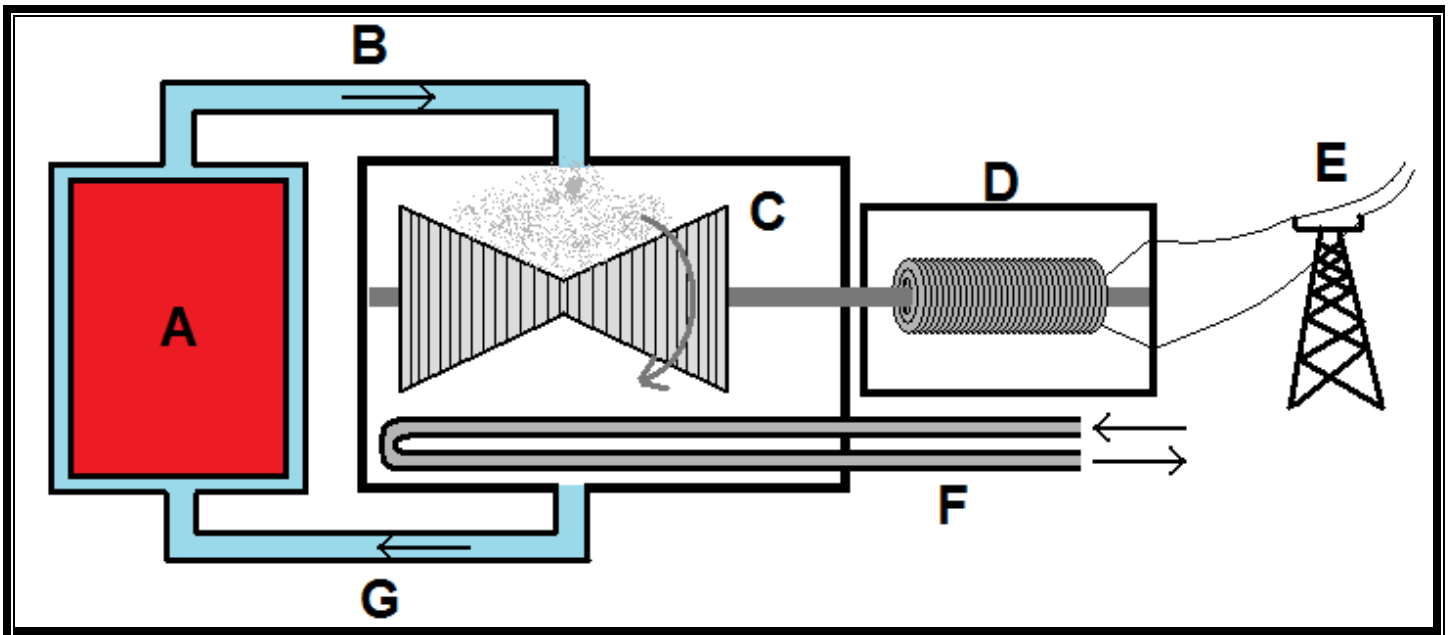
AC comes into the converter on one end, and leaves as DC electricity on the other end. Most devices run on DC converted from AC. Desktop computers and other devices without a "wall wart" use a built-in converter. In the case of a device that makes use of motors, generally, some of the incoming AC current will be sent to the motor, and the rest converted into DC so that the electronics can operate properly. As you'll learn in Physics, electronic devices are designed so that electricity can only run one way through the transistors, resistors, diodes, capacitors and other electronic components that make up the device. This requires DC.

So why don't we just generate DC electricity? Because AC is currently less expensive to produce and much more efficient to transport from the power plant to homes. Back when the fate of AC-generating power plants (championed by George Westinghouse) and DC-generating power plants (championed by Thomas Edison) was being decided for future generations, Edison, in desperation, filmed an elephant being electrocuted by alternating current to show how dangerous it was. This bit of propaganda failed, and today AC current is used for long-range networks of electrical production (the so-called "electrical grid") and DC current is used for short-range use within devices. Portable DC generators, known as batteries, power all of the devices we take with us, and AC from your wall socket (with the help of a converter) is used to recharge them or use them without using up the battery.

DC can not only be generated with a wet or dry cell battery (see the last topic), but refillable fuel cells and solar power. AC can be generated not only with steam turning a turbine, but also through wind and tide. These alternative methods of generating electricity represent the future of humanity's electrical use. They do not produce environmentally dangerous products, are endlessly renewable, and will strongly reduce our dependence on nonrenewable and environmentally hazardous sources like coal or oil.

## 1) Coal, Oil, Nuclear Steam-Based Power Plants (Alternating Current)

Coal is mined, as is uranium. Oil is pumped out of the ground or mined in tar shales or sands. Natural gas is siphoned out of the ground. Each of these fuel sources has its benefits and detriments. All are nonrenewable, which means that once they are gone, they are gone forever (or, in the case of oil and coal, until hundreds of millions of years have passed). Burning fuels generates oxides of carbon, nitrogen and sulfur, each of which is considered to be air pollution. Nuclear fission leaves behind highly radioactive isotopes with long half-lives that we have not yet come up with a way to dispose of safely. The reason we use these fuels is because they are relatively inexpensive at the current time, and we are heavily invested in the technology. It costs billions of dollars to build new power plants. This is the source of the vast majority of electrical production across the world.



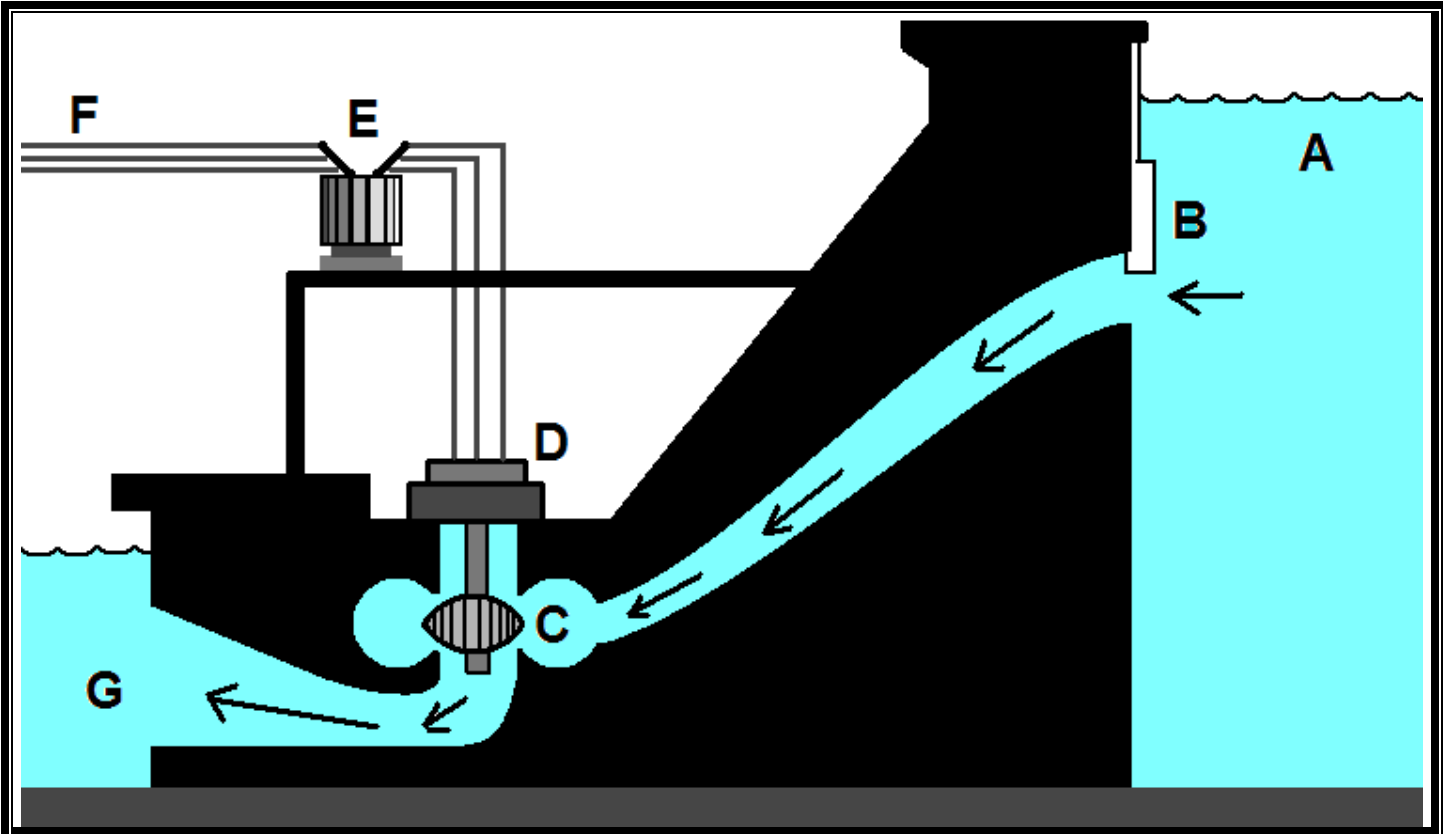
Letter	Part	What Is Happening Here?
A	Heat Source	This is where the heat is generated by burning coal, oil, natural gas, biomass or a nuclear fission reactor core.
B	Hot Water Loop	This is sometimes two loops, one for high pressure water and the other for lower pressure water. In both, water is turned to steam by the heat source and is directed under pressure towards the turbines.
C	Turbines	These consist of row after row of fan blades around a circular spindle. The farther from the center you go, the larger the blades. The pressurized steam catches the fan blades, causing the assembly to turn at the rate of 60 revolutions per second. The axle of the turbine is connected to a magnet in the generator room. To control the speed of the turbines, heat is increased or decreased by computer.
D	Generator	A magnet attached to the turbine axle rotates inside a coil of wire at the rate of 60 rpm, generating an AC electric current.
E	Substation	The electricity is distributed to substations which route the electricity to its final destination; your business or home!
F	Condenser Loop	This brings in cold water from a lake or river, which condenses the steam back into liquid water.
G	Return Loop	The water is pumped back into the heat source, where the process repeats.



## 2) Hydroelectric Generation (Alternating Current)

Just as there is electrical potential difference between the oxidation and reduction half-cells of a voltaic cells, there is a gravitational potential difference between water moving from high elevations to low elevations. This effect is maximized by putting a dam in the way of a river. Water backs up at the high end, forming a lake or reservoir. As the water is let through the dam, it decreases in height, releasing its energy to turning the turbine, which turns a generator to produce electricity. As long as there is water in the reservoir and the dam is standing, it will continue to produce electrical current. The best part? There is no pollution. It can affect the ecosystem of the river, however, but once the dam is established, the ecosystem reaches a new equilibrium.

The greater the difference in height between A and G, the greater the electrical output will be.



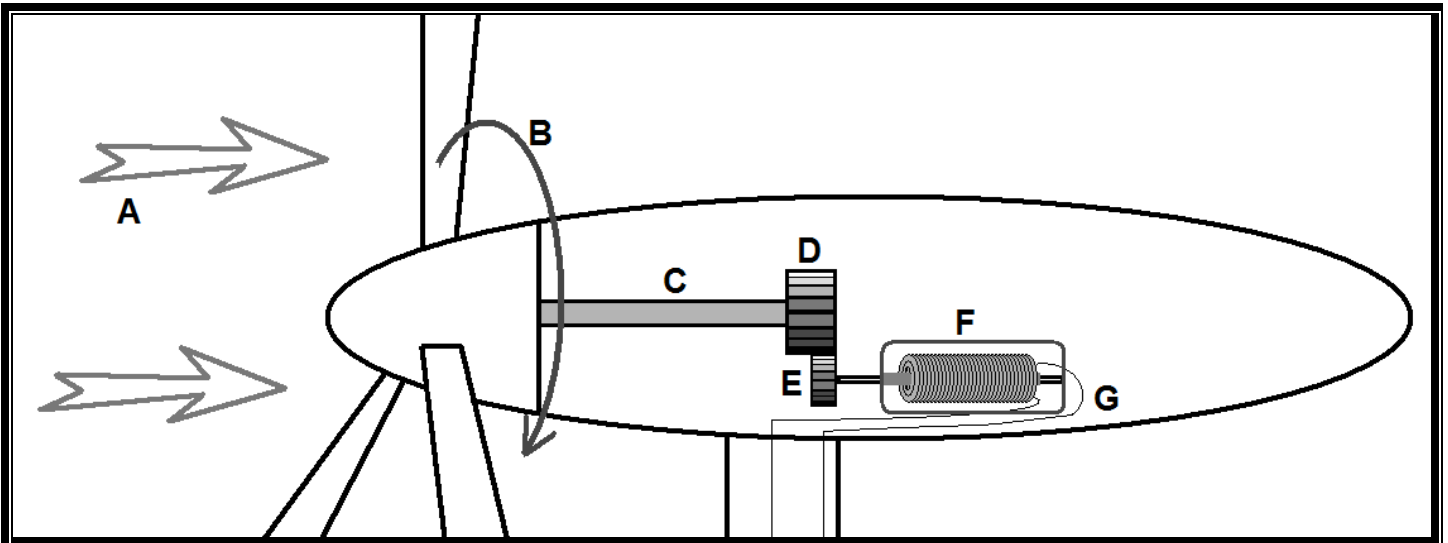
Letter	Part	What Is Happening Here?
A	Reservoir	Water is backed up by the high wall of the dam, where it has a tremendous amount of potential energy compared to the low end of the dam.
B	Sluice Gate	This is opened to let more water in, and closed to let less water or no water into the dam.
C	Turbine	The water coming into the dam converts potential energy into kinetic energy by turning the turbine, which is connected to a magnet in the generator.
D	Generator	Again, a magnet turning in a coil of wire to generate electric current.
E	Transformer	Since the turbine is not turning at the rate of 60 rpm, the transformer converts the slower frequency AC current into 60 Hz current with very high voltage.
F	Transmission Cables	Carry the electricity to substations, which distribute the electricity to where it's needed.
G	Stream Outlet	This is where the water has its lowest potential energy. From this point, the river continues onwards.

### 3) Wind Power (Alternating Current)



On the Great Plains of America, the land is flat and there is nothing to stop the winds from howling over the ground. Trees are rare, hills are gentle and you can see for hundreds of miles on clear days. Popping up all over this landscape are wind turbines, set into large groups called “wind farms”. Each turbine doesn’t generate even a fraction of what is generated in steam turbine or hydroelectric plants; but together in a wind farm, they can generate a lot of electricity. Wind farms are common in areas where there are wide spaces devoid of people. The main challenge of wind turbine generation is getting the electricity the long distances from the wind farm to the populated areas that use the electricity. In recent years, breakthroughs have been made in household wind turbines, and many people now get some of their electricity from wind turbines mounted to the roof of their house.

Each turbine has an anemometer to determine wind speed and a weather vane to detect wind direction. The nacelle that houses the generator inside and holds the rotor assembly that the turbine blades is able to move from side to side to catch the wind as efficiently as possible. This is accomplished with motors built into the tower, and uses a small amount of the energy generated by the turbine. This way, it doesn’t matter what direction the wind is blowing, the turbine will always work at peak efficiency.



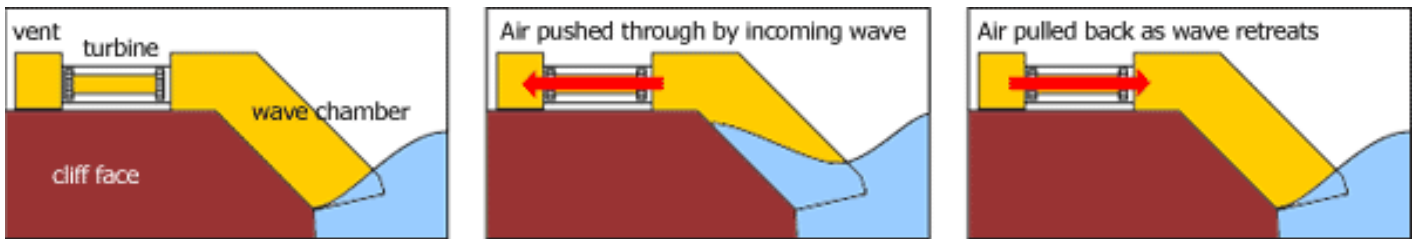
Letter	Part	What is Happening Here?
A	Wind	Wind comes sweeping in from across the landscape. The wind turbine nacelle is designed to rotate on the top of its tower to get maximum wind potential. This system uses a little of the electricity generated by the turbine assembly.
B	Turbine Blades	These aerodynamically designed blades are the length of two tractor trailers and catch the wind in such a way that the rotor assembly (the front cap where the blades are attached) can turn.
C	Low-Speed Shaft	This connects the rotor assembly to the gear box. It turns slowly, but the gear box scales up the speed so that 60 Hz AC current can be produced at the generator.
D	Gear Box	A series of gears designed to convert slow rotational speed to high rotational speed to turn the magnet inside the generator.
E	High-Speed Shaft	Connects the gear box to the magnet in the generator.
F	Generator	Rotates a magnet inside a coil of wire to generate an AC electrical current.
G	Transmission Lines	Go down through the tower into the ground, where the wires are routed to transmission cables to carry the electricity to power substations to be sent out wherever it’s needed.

#### 4) Wave Power (Alternating Current)

This largely experimental technology uses the motion of waves to generate electricity. This is the newest way to generate electricity. It is yet another energy source that does not require burning fuels to make electricity. It must be designed and placed in a way that it has minimal impact on sea life and shipping routes. This is an emerging “green” technology that will continue to evolve in the coming years.

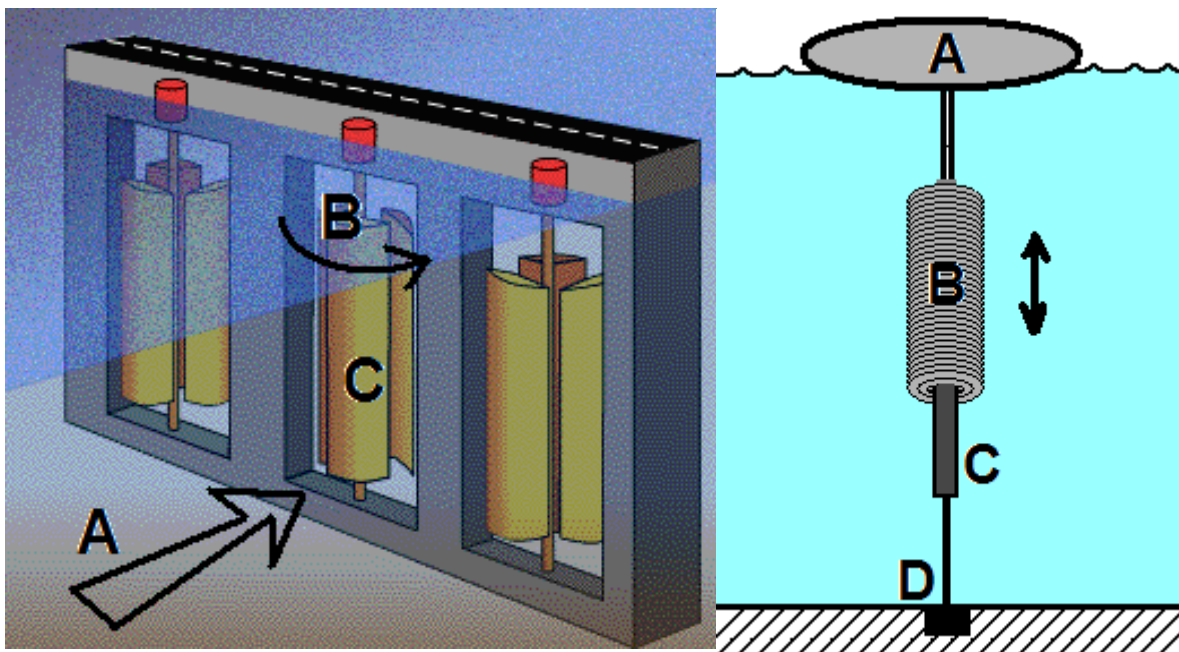
##### Methods of using wave energy:

A) Using water pressure from waves to compress air to spin turbine blades to spin a magnet through a coil of wire repeatedly to generate AC electric current:



B) Underwater “Tidal Fence” Vertical Turbine:

C) Buoy Generator



Part	What It Is	What's Happening	Part	What It Is	What's Happening
A	Water Current	The tide pushes water towards shore.	A	Buoy	This buoy is carried up and down on the ocean waves.
B	Turbine Blades	The water turns the vertical turbine blades.	B	Coil Of Wire	This coil is attached to the buoy and is allowed to move freely over the magnet, generating AC current.
C	Generator	The turbine blades are attached to a coil of wire that rotates around a fixed magnet, generating AC current.	C	Magnet	This magnet is fixed, anchored into the ocean floor.
			D	Anchor	This is where the magnet is anchored to the ocean floor.

## 5) Solar Power (Direct Current)

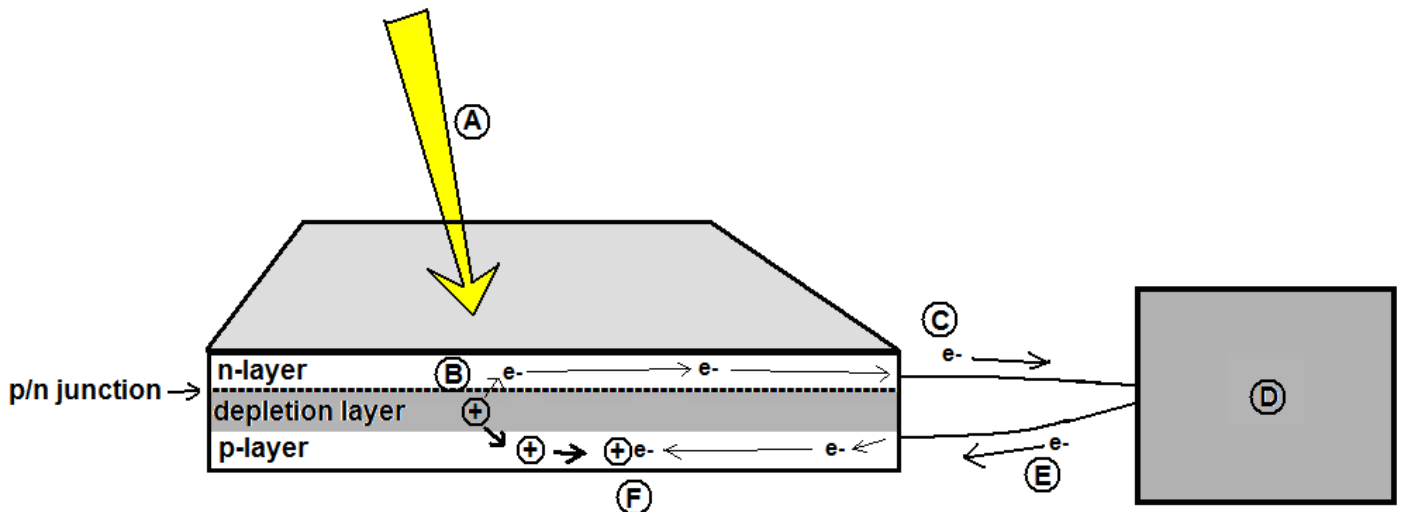
Light can provide a source of energy to ionize atoms. Semiconductors like silicon are especially susceptible to this. Light striking a silicon atom just right can knock an electron out of the valence shell and start it moving towards a region that would prefer to gain electrons. In a photovoltaic solar cell, there are two layers, called the p (for positive) and n (for negative) layers with a boundary between them called the p/n junction. Here's how it works:

**Depletion Layer:** The top part of the p layer, when hit with sunlight, an electron is excited off of an atom, and electric fields move the two apart. The + "hole" that is left behind migrates down into the p-layer, and the electron is pushed across the p/n junction. The + hole is a virtual hole...no atoms actually move. Through a cascade of electron flow, the atoms pass the positive "hole" down further into the p layer. In this manner, fresh electrons are always migrating towards the depletion layer.

**p/n Junction:** The p/n junction is "selectively permeable" in that only negatively charged electrons are allowed to pass through.

**N Layer:** The electrons travel through the n-layer to a conductive layer which carries the electrons to wires that connect the solar cell to the load.

**P Layer:** Electrons go into the p-layer and are captured by the virtual "positive hole" that migrated deeper into the p-layer when it lost an electron. The + hole is then effectively neutralized and the circuit completed.



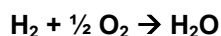
Letter	Part	What Is Happening Here?
A	Sunlight	Light strikes the surface of the solar panel and goes through the thin n-layer to strike the "depletion zone" at the top of the p-layer.
B	n-layer	The light strikes an electron off of an atom in the "depletion zone", creating a virtual "+ hole", which moves deeper into the p-layer, as the electron is forced to cross the n/p junction due to built-in fields into the n-layer.
C	Wire (anode)	Electrons leave the n-layer through a conductive film on top of the n-layer, to a wire.
D	Load	The device that is using the electrons, like a calculator or solar charger.
E	Wire (cathode)	Electrons leave the load and head to the p-layer.
F	p-layer	The + hole, having migrated deeper into the p-layer, captures the electron coming from the wire. This converts the + hole back to a neutral atom.

The process repeats endlessly as long as light is striking the photovoltaic surface of the cell. Since this produces DC current, if you are going to use solar panels to power household appliances, such as a refrigerator or a television, the current must pass through a "DC Inverter" that converts the current to AC power before appliances can use it. Solar power ranges from cheap cells, which power calculators, which operate at 5% efficiency (converting 5% of the incoming solar energy to electricity) to high-end single-crystal cells, which are much more expensive, but operate at 25% efficiency. Higher efficiency means smaller panels to accomplish the same task. There are new "dye-sensitized" cells out there that were designed based on how plants convert sunlight into energy; photosynthesis!

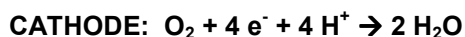
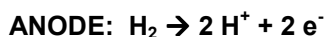
## 6) Hydrogen Fuel Cell Technology (Direct Current)

Wet cells consist of solid electrodes in liquid electrolytes. Dry cells consist of solid electrodes in an electrolyte gel or paste. As the anode gets used up, the cell dies. With a fuel cell, the oxidized species is in the gas phase, and is constantly being replenished. As long as this fuel is supplied to the cell, it will continue to produce electricity at a constant rate. In a hydrogen fuel cell, hydrogen gas is the fuel. The hydrogen can either be stored in pressurized tanks (as in fuel cell cars) or stripped from a molecule of methanol by a catalyst (which produces CO<sub>2</sub> as a byproduct) in the case of fuel cells designed to replace batteries in portable electronics. Hydrogen is also currently produced from decomposing natural gas (methane) and from coal, so we need to find a cleaner way to get our hydrogen.

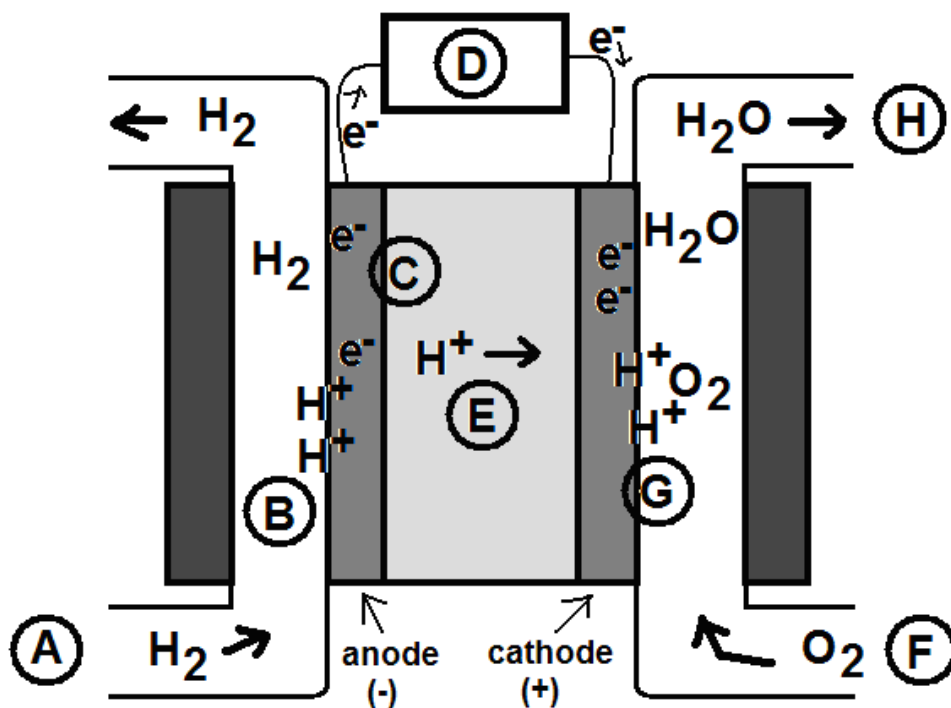
The reaction for this fuel cell is:



The half-reactions for this fuel cell are:



The net voltage of this type of fuel cell is about 0.7 volts per cell. The cells are stacked to form batteries that can generate as much voltage as you need for your application. 10 cells will generate 7 volts, 20 cells will generate 14 volts, and so on.



Letter	Part	What is Happening Here?
A	H <sub>2</sub> inlet	H <sub>2</sub> is pumped into the fuel cell from a tank containing pressurized hydrogen.
B	Anode	H <sub>2</sub> is oxidized to H <sup>+</sup> by a platinum catalyst. The H <sub>2</sub> that is not oxidized recycles back into the cell.
C	Anode	Electrons stripped from the hydrogen move through the anode towards the wire.
D	Load	Electrons travel through the load and power it.
E	Polymer Electrolyte Membrane	The equivalent of a salt bridge, this membrane is engineered to let only cations through it, not anions. The H <sup>+</sup> moves through it, migrating towards the cathode.
F	Air Intake	Air is pumped in from the outside.
G	Cathode	The H <sup>+</sup> reacts with the O <sub>2</sub> and the electrons coming down the cathode to form H <sub>2</sub> O.
H	Exhaust	The only exhaust product from this process is pure water vapor.

## 4) Using Electricity to Decompose Compounds (HW: p. 35, 36)

**Essential Question:** If chlorine can only be found in compounds in nature, then how is it we have chlorine to use in our pools?

Do you remember way back when, you learned about the properties of the elements in various groups? Let's refresh your memory!

Group	Name Of Group	Properties Of Elements In Groups
1	Alkali metals	<ul style="list-style-type: none"> <li>- form +1 ions by losing their only valence electron</li> <li>- have the lowest electronegativity and ionization energy of any group on the Table</li> <li>- have the largest atomic radius of all the groups on the Table</li> <li>- react violently with water to produce a strong base and hydrogen gas</li> <li>- <b>only found in compounds in nature</b></li> <li>- <b>can be extracted from compounds using electrolytic reduction</b></li> </ul>
2	Alkaline Earth metals	<ul style="list-style-type: none"> <li>- form +2 ions by losing their only 2 valence electrons</li> <li>- have the second lowest electronegativity and ionization energy of all the groups on the Table</li> <li>- react quickly with water to produce a weak base and hydrogen gas</li> <li>- <b>only found in compounds in nature</b></li> <li>- <b>can be extracted from compounds using electrolytic reduction</b></li> </ul>
3-12	Transition metals	<ul style="list-style-type: none"> <li>- can lose electrons from valence shell and the one underneath it to form more than one possible charge</li> <li>- form colored compounds</li> <li>- can be removed from compounds using a variety of complex reactions, including single replacement (reacting the compound containing the desired metal with a more active metal)</li> </ul>
17	Halogens	<ul style="list-style-type: none"> <li>- form -1 ions by gaining 1 electron to give them a stable octet</li> <li>- have the highest electronegativity of all the groups on the Table</li> <li>- have the smallest atomic radius of all the groups on the table</li> <li>- react violently with metals to form salts</li> <li>- react with hydrogen to form (mostly) strong acids</li> <li>- <b>only found in compounds in nature</b></li> <li>- <b>can be extracted from compounds using electrolytic oxidation</b></li> </ul>
18	Noble Gases	<ul style="list-style-type: none"> <li>- do not form ions, because they already have a stable octet</li> <li>- have the highest ionization energy of all the groups on the Table</li> <li>- always found in pure, uncombined state in nature (monatomic molecules)</li> </ul>

Even though Group 1 and 2 metals and Group 17 nonmetals are only found in compounds in nature, their pure forms have many uses. For example, sodium vapor is used for yellow street lamps; chlorine is used as a chemical disinfectant. We have developed ways of extracting these highly reactive elements from their compounds, the most common (and cost-effective) of which is called **ELECTROLYSIS**.

**ELECTRO = the use of electricity      LYSIS = the separation of a thing into its parts**

Electrolysis involves melting the ionic solid that contains the element you want to extract and passing electric current through it. This current has to be **DIRECT CURRENT (DC)**, the kind produced directly by batteries or by AC passed through a converter. The electrons must only go one way, which is the only way that the individual elements in the compound can be completely separated.

The – electrode from the DC power supply provides a source of electrons that can **REDUCE** the positive metal ion to form pure metal:  $\text{Na}^{+1} + 1 \text{e}^{-}$  (from the DC power supply)  $\rightarrow$   $\text{Na}^0$  (pure sodium metal).

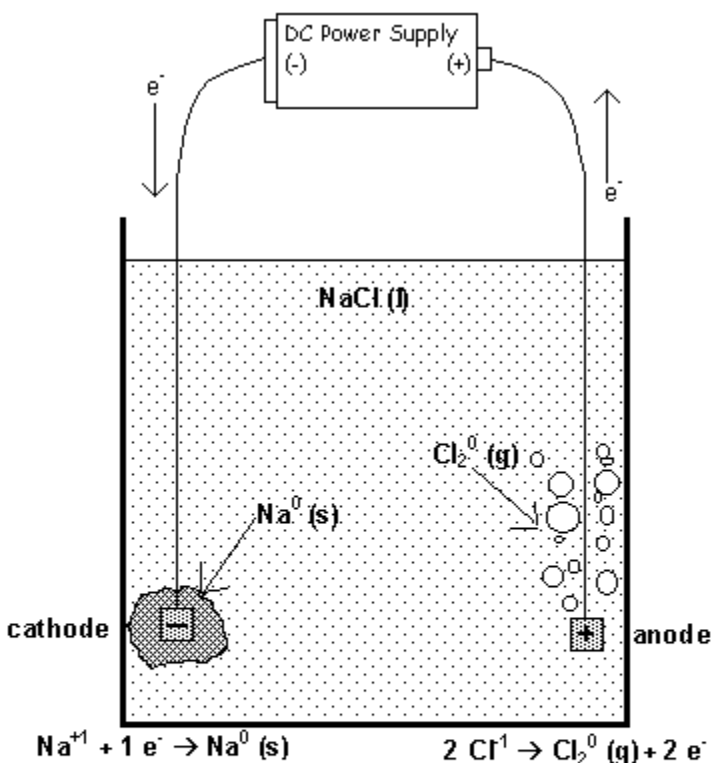
The + electrode from the DC power supply removes electrons from the negative ion, causing the negative nonmetal ions to be reduced to form pure nonmetal gas:  $2 \text{Cl}^{-1} \rightarrow \text{Cl}_2^0$  (pure chlorine gas) +  $2 \text{e}^{-}$  (which are pulled into the + side of the DC supply)

## The electrolysis of a fused solid (NaCl (l))

This is the industrial process used to extract sodium and chlorine from molten sodium chloride. The salt is placed in a furnace and heated to melting. Then electrodes from a source of direct current (DC) are placed in the liquid salt. These electrodes are made of a material with a higher melting point than the salt, so that they don't get melted in the process.

**Though this example shows the electrolytic decomposition of NaCl, this same process can be used to extract AND Group 1 or 2 metal from a compound. It is also used to extract Group 17 nonmetals from those compounds.**

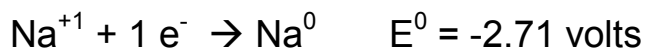
Why do you have to force the salt to decompose? Remember, nature tends towards lower energy states. These active metals and nonmetals are unstable in their pure state, and they become more stable as they form stable octets by bonding. The compounds they form with each other are a lot more stable than the elements were when they were separated. All that energy that was lost when the elements combined has to be given back to the compound to get them to separate again! When the elements have been separated, you have to handle them very carefully! The nonmetal has to be stored in nonreactive containers (metal won't do, because halogens eat metals for breakfast). The metal has to be stored under mineral oil if it is an alkali metal, or sealed up well in their storage containers if they are alkaline earth metals.



■ Represents a platinum catalyst. Platinum is used because it is an excellent conductor of electricity and it does not react with other elements. It is one of the least reactive elements on the periodic table.

### What is Happening Here???

Electrons are coming out of the negative end of the DC power supply. These travel down into the liquid NaCl, into the electrode. Since electrons are negative, they attract the positively charged  $\text{Na}^{+1}$  ions to it. The  $\text{Na}^{+1}$  is forced to eat the electrons, reducing them to metallic  $\text{Na}^0$ , which forms around the electrode.



Electrons are being yanked towards the positive side of the DC power supply. The  $\text{Cl}^{-1}$  ions are forced to give up their electrons and be oxidized into  $\text{Cl}_2^0 (\text{g})$  molecules, which bubble up and are trapped for later use. The electrons ripped from the  $\text{Cl}^{-1}$  are pulled back into the positive side of the DC power supply, completing the circuit.



This same setup can be used to decompose any binary compound containing an alkali metal and halogen. The metal ion will always be reduced at the negative cathode and the nonmetal ion will always be oxidized at the positive anode. This has to be carried out in the equivalent of a blast furnace...NaCl melts at 1074 K!

**Since the total voltage of this cell is -4.07 volts, DC current with at least that much voltage must be added to make this electrolysis occur.**

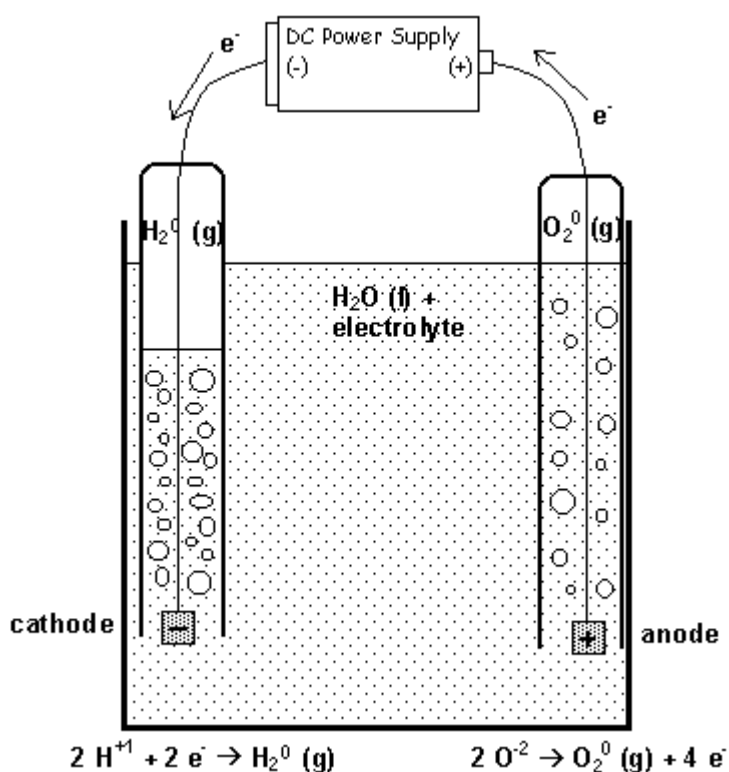
## The electrolysis of water (H<sub>2</sub>O (l))

The “hydrogen economy”. This is a phrase that has become more commonplace as the United States looks to move beyond oil to more renewable energy sources. Hydrogen is a good fuel, as it is plentiful, burns cleanly with oxygen to produce water as the only product, and as a gas it is compressible for storage purposes. The only problem with hydrogen is that, being a Group 1 element, it is only found in compounds here on Earth. The most common compound? Water, of course. Hydrogen can be easily extracted from methanol (CH<sub>3</sub>OH), but that produces CO<sub>2</sub>, which is one of the things we are hoping to eliminate by moving away from fossil fuels!

OK, so hydrogen is found in water. Water is a very stable compound, and cannot be decomposed without either adding extreme heat (well beyond that needed to boil water) (which is not cost-effective) or applying electric current. Just like sodium and chlorine can be extracted from sodium chloride using electricity, hydrogen can be very effectively extracted from water using the same method. Since water is already a liquid at room temperature, you don't have to add any heat. The products, hydrogen and oxygen gas, can be simply trapped in test tubes if you are doing the extraction on a small scale.

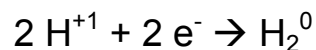
Water does not conduct electricity well (it is a nonelectrolyte), unlike ionic liquids which have freely-moving ions. To allow electricity to pass through the water, an electrolyte is added. You have to be careful as to which one you add, though...adding a salt might interfere with the electrolysis and give you unwanted products. Adding an acid, on the other hand, will work well. Sulfuric acid (H<sub>2</sub>SO<sub>4</sub>) is a common electrolyte to use, because it will not interfere with the process in any way.

The source of electricity can come from solar power, or from nuclear power in a world free of burning fossil fuels. In this way, we can have a source of chemical power that is totally renewable, completely clean, and, with proper engineering, extremely safe.



### What is Happening Here???

Electrons are coming out of the negative end of the DC power supply. These travel down into the liquid water (which has electrolyte like sulfuric acid dissolved in it to conduct electricity, since water does not conduct on its own), into the electrode. Since electrons are negative, they attract the positively charged  $H^+$  ions to it. The  $H^+$  is forced to eat the electrons, reducing them to  $H_2$  gas, which bubbles up into the test tube and is trapped by water displacement.



Electrons are being yanked towards the positive side of the DC power supply. The  $O^{2-}$  ions are forced to give up their electrons and be oxidized into  $O_2(g)$  molecules, which bubble up and are trapped by water displacement. The electrons ripped from the  $O^{2-}$  are pulled back into the positive side of the DC power supply, completing the circuit.



**Notice that there is twice as much trapped  $H_2(g)$  as there is  $O_2(g)$ ?** This is because the formula of water is  $H_2O$ . When the water is decomposed, there are two moles of  $H_2(g)$  produced for every one mole of  $O_2(g)$  produced:





## Electroplating One Metal Onto Another

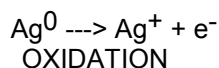
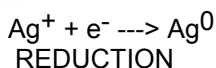
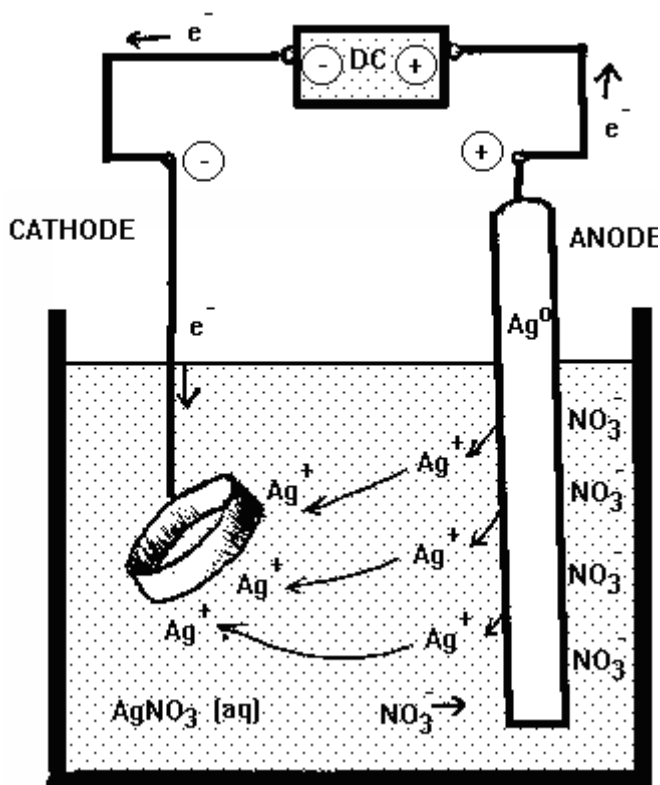
Metals are excellent for building things, like building structures, ships, aircraft components, artificial joints, bearings and jewelry. Many metals that are excellent for making these things have a fatal flaw; they wear easily or they corrode easily. Or perhaps they are just too expensive. To help make the base metal part more durable so it lasts longer or looks prettier, it is often coated with a thin layer of a more durable metal, or a more decorative metal.

Steel can be used to make bearings, metal parts that move over other metal parts in an engine. Because metal is rubbing on metal, over time the metal can wear away, exposing surface area that can rust. Electroplating steel with a less active metal can provide extra durability. The less active metal might not have the tensile strength that steel has, but is better at resisting wear due to use. In this way, both metals can be put to good use doing what they both do best.

Motorcycle enthusiasts love chrome. It's bright, shiny and makes their bikes flash in the sun. Chromium is very expensive, though, and it would be too expensive to make the metal parts of the bike out of pure chromium. In addition, chromium is a poor choice of metal to use to make engine and exhaust parts with. Use steel or aluminum, which are strong enough to handle the rigors of heavy use, and electroplate those parts with a layer of chromium! You get all the benefits of steel or aluminum construction, with the shiny chromium finish!

How do you get the plated metal onto the base metal? It's fairly easy. Just take the base metal part that you want to coat with the plating metal and place it in a bath that has a salt containing the metal you want to plate with dissolved in water. Put bars of the plating metal into the salt bath as well. Attach the base metal part to the – end of a DC power supply, and the plating metal bars to the + end of the DC power supply. Turn on the current. The longer you let the electricity run, the thicker the metal plating will be. The best part? The coating will be of an even thickness over the entire base metal part.

The diagram below demonstrates how silver can be electroplated onto a ring made of a base metal (like nickel). The ring is hooked up to the – electrode, the plating metal (silver) is hooked up to the + electrode, and both items are immersed in a bath containing silver nitrate.

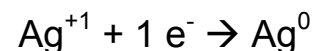


### What is Happening Here?

Silver is being oxidized at the + electrode as its electrons get stripped off and pulled into the + end of the DC power supply. The silver dissolves into solution as silver ions.



Meanwhile, over at the – electrode, electrons are being pumped out of the DC power supply and into the ring. Metal atoms don't gain electrons, so the electrons just hang out on the surface of the ring, giving the ring a negative charge. This attracts the positive silver ions in the solution, which migrate to the ring and touch it. As soon as a silver ion touches the ring, an electron jumps off the ring onto the silver ion, reducing it to solid silver metal, which freezes on to the ring.

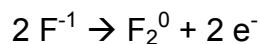


As more and more silver ions reduce and plate on to the ring, the thickness of the plating increases. The final thickness depends on how much current is being used, and how long you let the process continue.

## Practical Examples:

### 1) When KF (s) is decomposed, which element will form at the anode?

The anode is where OXIDATION takes place. In an electrolytic cell, the negative ion is oxidized.  $F^{-1}$  is the negative ion. It gets oxidized to pure  $F_2$ :



Therefore, fluorine,  $F_2$ , will form at the anode.

### 2) When KF (s) is decomposed, which element will form at the cathode?

The cathode is where REDUCTION takes place. In an electrolytic cell, the positive ion is reduced.  $K^{+1}$  is the positive ion. It gets reduced to pure K:



Therefore, potassium, K, will form at the cathode.

### 3) When $H_2O$ is electrolytically decomposed, which substance can be added to make water conduct electricity?

Since water does not conduct electricity on its own, **an acid, ( $H_2SO_4$ )** should be added to provide ions that will allow electricity to flow through the water without affecting the outcome of the electrolysis.

### 4) What solution can be used if you want to copper-plate a zinc penny?

To copper plate a zinc penny, you need to use a solution of a salt containing the plating metal, in this case, copper. You want to choose a soluble salt (from Reference Table F), and you can't beat a nitrate for solubility. **Copper (II) nitrate,  $Cu(NO_3)_2$** , is the best salt to use for this purpose.

### 5) If you want to chrome-plate an aluminum motorcycle tailpipe, what gets hooked up to what electrode of the DC power supply?

You will need to **hook the tailpipe up to the NEGATIVE electrode, and the bar of chromium to the POSITIVE electrode**. Both are then immersed in a solution of chromium salt (perhaps chromium nitrate), and then turn the juice on until you've got the thickness of chromium coating that you want!

### 6) What procedure would you go through to extract fluorine gas from lithium fluoride?

- Place solid LiF into a reaction cell.
- Heat the LiF to melting.
- Place electrodes, hooked into a DC power supply, into the liquid salt.
- Turn the electricity on.

The lithium will REDUCE at the – cathode:  $Li^{+1} + 1 e^{-} \rightarrow Li^0$

The fluoride will OXIDIZE at the + anode:  $2 F^{-1} \rightarrow F_2^0 + 2 e^{-}$

The fluorine is released as a gas which bubbles up from the liquid salt, all you have to do is trap it and it's yours!

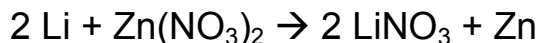
Student Name: \_\_\_\_\_ Grades: \_\_\_\_\_, \_\_\_\_\_, \_\_\_\_\_, \_\_\_\_\_  
HalfRxns Voltaic

## 1) Half-Reactions Homework

A) Determine the oxidation numbers of each element in the following substances. Report only the individual oxidation number, not the total contribution. Hint: O will always be -2. If you put anything other than -2 down for O, you know you are doing it incorrectly.

1) N <sub>2</sub>	N:		
2) ZnCl <sub>2</sub>	Zn:	Cl:	
3) K <sub>2</sub> SO <sub>3</sub>	K:	S:	O:
4) S <sub>8</sub>	S:		
5) Fe	Fe:		
6) PbO	Pb:	O:	
7) PbO <sub>2</sub>	Pb:	O:	
8) CuCO <sub>3</sub>	Cu	C	O
9) Na <sub>2</sub> CrO <sub>4</sub>	Na:	Cr:	O:
10) Cu(ClO <sub>3</sub> ) <sub>2</sub>	Cu:	Cl:	O:
11) Cu(ClO <sub>2</sub> ) <sub>2</sub>	Cu:	Cl:	O:
12) Cu(ClO) <sub>2</sub>	Cu:	Cl:	O:
13) K	K:		
14) NO <sub>2</sub>	N:	O:	
15) SO <sub>2</sub>	S:	O:	
16) Pb(Cr <sub>2</sub> O <sub>7</sub> ) <sub>2</sub>	Pb:	Cr:	O:

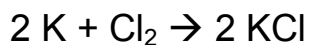
B) Write all of the charges for each species above their symbols, then write the oxidation half-reaction and reduction half-reaction, identify the oxidizing and reducing agents and any spectator ions.



Oxidation Half-Reaction: \_\_\_\_\_

Reduction Half-Reaction: \_\_\_\_\_

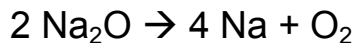
Oxidizing Agent: \_\_\_\_\_ Reducing Agent: \_\_\_\_\_ Spectator Ion: \_\_\_\_\_



Oxidation Half-Reaction: \_\_\_\_\_

Reduction Half-Reaction: \_\_\_\_\_

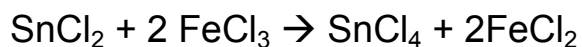
Oxidizing Agent: \_\_\_\_\_ Reducing Agent: \_\_\_\_\_ Spectator Ion: \_\_\_\_\_



Oxidation Half-Reaction: \_\_\_\_\_

Reduction Half-Reaction: \_\_\_\_\_

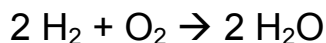
Oxidizing Agent: \_\_\_\_\_ Reducing Agent: \_\_\_\_\_ Spectator Ion: \_\_\_\_\_



Oxidation Half-Reaction: \_\_\_\_\_

Reduction Half-Reaction: \_\_\_\_\_

Oxidizing Agent: \_\_\_\_\_ Reducing Agent: \_\_\_\_\_ Spectator Ion: \_\_\_\_\_

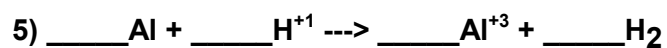
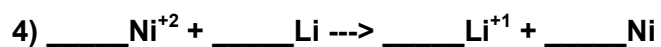
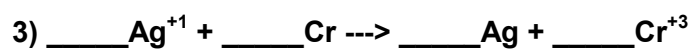
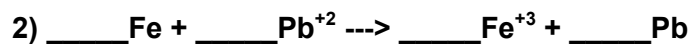
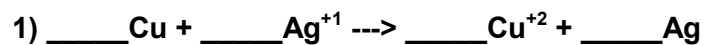


Oxidation Half-Reaction: \_\_\_\_\_

Reduction Half-Reaction: \_\_\_\_\_

Oxidizing Agent: \_\_\_\_\_ Reducing Agent: \_\_\_\_\_ Spectator Ion: \_\_\_\_\_

C) Balance the following redox reactions by the half-reaction method, rewriting the balanced equations below the given unbalanced equation. Show your work below each reaction and put coefficients in the spaces provided:



## 2) Electrochemistry Homework

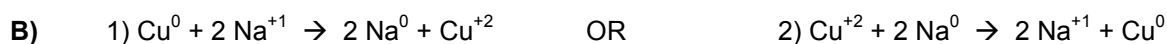
1) Write the symbol & charge of a metal ion that has a greater potential to undergo oxidation than Na \_\_\_\_\_

2) Write the symbol & charge of a nonmetal that has a greater potential to undergo reduction than  $\text{Cl}^{-1}$  \_\_\_\_\_

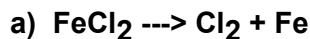
3) When the following pairs of elements are reacted, indicate which element will undergo oxidation and which element will undergo reduction based on each element's relative reactivity on Reference Table J.

Pair	Oxidation	Reduction	Pair	Oxidation	Reduction
Mg and Cu			Ni and Pb		
Zn and Li			Na and Cu		
Mn and Ba			Au and Ag		

4) Two pairs of ionic reactions are listed below. Either 1) or 2) will react. For each pair, use Reference Table J to determine which will spontaneously react and circle that pair (either 1) or 2)).

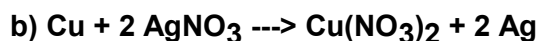


5) For the following redox reactions, identify the charges of each species, write the oxidation and reduction half-reactions and determine the net voltage and write if the reaction is spontaneous or nonspontaneous.



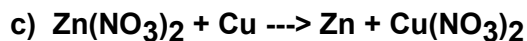
OX:  $E^0$  Total  $E^0$  for reaction Spontaneous?

RD:



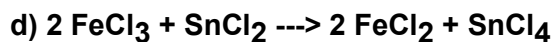
OX:  $E^0$  Total  $E^0$  for reaction Spontaneous?

RD:



OX:  $E^0$  Total  $E^0$  for reaction Spontaneous?

RD:

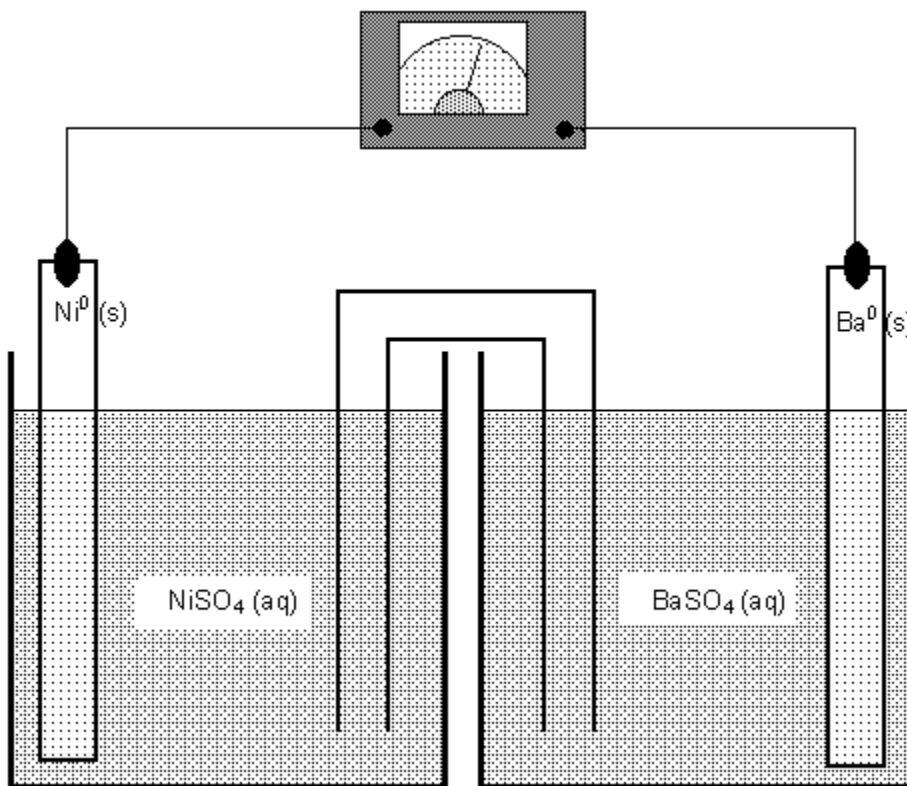


OX:  $E^0$  Total  $E^0$  for reaction Spontaneous?

RD:

6) Given the following nickel/barium voltaic cell diagram, label the anode, cathode, + electrode, - electrode, direction of electron flow, direction of anion flow, salt bridge and load Base your decisions on Reference Table J.

- a) According to Table J, \_\_\_\_\_ will undergo oxidation and \_\_\_\_\_ will undergo reduction.  
 b) Label the diagram below:



- c) Oxidation Half-Reaction: \_\_\_\_\_  $E^0$  \_\_\_\_\_  
 d) Reduction Half-Reaction: \_\_\_\_\_  $E^0$  \_\_\_\_\_ Total  $E^0$ : \_\_\_\_\_  
 e) Oxidizing Agent: \_\_\_\_\_ Reducing Agent: \_\_\_\_\_

\_\_\_\_\_ f) Which of the following statements are true?

- 1) The Ni electrode is getting larger because the nickel atoms are being oxidized to form nickel ions.
- 2) The Ni electrode is getting larger because the nickel ions are being reduced to form nickel atoms.
- 3) The Ni electrode is getting smaller because the nickel atoms are being oxidized to form nickel ions.
- 4) The Ni electrode is getting smaller because the nickel ions are being reduced to form nickel atoms.

\_\_\_\_\_ g) Which of the following statements are true?

- 1) The Ba electrode is getting larger because the barium atoms are being oxidized to form barium ions.
- 2) The Ba electrode is getting larger because the barium ions are being reduced to form barium atoms.
- 3) The Ba electrode is getting smaller because the barium atoms are being oxidized to form barium ions.
- 4) The Ba electrode is getting smaller because the barium ions are being reduced to form barium atoms.

h) What direction are the cations going across the salt bridge? \_\_\_\_\_

i) Can  $\text{Ba}(\text{NO}_3)_2(\text{aq})$  be used to make the salt bridge? Why or why not?

7) Using the metal side of Reference Table J table, design two wet voltaic cells. There will be an extra five points awarded if you design a voltaic cell that packs the maximum possible voltage allowed for using the solid electrodes from AE Table D.

**FIRST CELL:**

- 1) More active metal: \_\_\_\_\_ Less active metal: \_\_\_\_\_
- 2) Solution with more active metal: \_\_\_\_\_ Solution with less active metal: \_\_\_\_\_
- 3) Balanced reaction for the cell: \_\_\_\_\_
- 4) OX: \_\_\_\_\_  $E^0$ : \_\_\_\_\_
- 5) RD: \_\_\_\_\_  $E^0$ : \_\_\_\_\_ Total  $E^0$  of the cell: \_\_\_\_\_
- 3) Draw a wet-cell diagram that contains your chosen metals and solutions, labeled with the following:
- a) anode
  - b) cathode
  - c) + electrode
  - d) – electrode
  - e) load
  - f) salt bridge
  - g) direction of electron flow
  - h) Direction of anion flow across the salt bridge
  - i) Direction of cation flow across the salt bridge

**Make sure you place your metals and solutions in the proper place.**



**SECOND CELL:**

- 1) More active metal: \_\_\_\_\_ Less active metal: \_\_\_\_\_
- 2) Solution with more active metal: \_\_\_\_\_ Solution with less active metal: \_\_\_\_\_
- 3) Balanced reaction for the cell: \_\_\_\_\_
- 4) OX: \_\_\_\_\_  $E^0$ : \_\_\_\_\_
- 5) RD: \_\_\_\_\_  $E^0$ : \_\_\_\_\_ Total  $E^0$  of the cell: \_\_\_\_\_
- 3) Draw a wet-cell diagram that contains your chosen metals and solutions, labeled with the following:
- a) anode
  - b) cathode
  - c) + electrode
  - d) – electrode
  - e) load
  - f) salt bridge
  - g) direction of electron flow
  - h) Direction of anion flow across the salt bridge
  - i) Direction of cation flow across the salt bridge

**Make sure you place your metals and solutions in the proper place.**

### 3) Other Ways Of Generating Electricity Homework

A) List two ways in which alternating current differs from direct current.

**B) RESEARCH!!! Please answer any THREE of the questions below. Type your answers on a separate page, and include a reference for each one.**

1) What is one benefit of powering your home with solar power from solar cells on the roof of your house? What is one drawback to using solar power for this purpose? How does the government make it more affordable for homeowners to have solar cells installed on their homes?

2) What is one benefit of using refillable methanol cartridges for the hydrogen source in a fuel cell designed for portable electronics such as notebook computers and cell phones? What is one drawback to using methanol cartridges for this purpose? Which would you rather use, a cell phone that runs on a rechargeable battery, or one that runs on a methanol-powered fuel cell? Explain your answer.

3) Photovoltaic solar panels are only one way to harness the sun's energy to make electricity. Describe another method of harnessing the sun's energy to make electricity. Which method is more efficient at converting solar energy to electricity?

4) What is one possible environmental impact of using tidal fence turbines to produce electricity, especially in estuaries? How can this impact be minimized?

5) If you lived in an area perfect for wind farming, would you be for or against putting a wind farm within 2 miles of your house? Explain your answer.

6) How much electricity (in megawatts) does the hydroelectric plant at Hoover Dam produce? What is the name of the reservoir created by the dam? How many homes can be powered by the electricity generated at Hoover Dam? How does this compare with the megawatt output and number of houses that can be powered by the Indian Point nuclear power plant in Buchanan, NY?

**C) +10 Extra Credit: AC current is generated by either rotating a magnet inside a fixed coil of wire, or rotating a coil of wire outside a fixed magnet. This rotation is created by using turbines that convert steam, wind, water and wave motion to rotating motion. Design a method of generating AC power that is creative, unique and thoroughly yours. In order to get the extra credit, you must:**

**1) Describe the method of how the wire or magnet is going to be rotated**

**2) The source of energy that will create the rotation**

**3) A sketch that shows the setup of your generator**

**4) Turn in a completely unique design. No one else can have the same design as you. If I can find this design on the Internet, then the extra credit will not be awarded.**

**5) The design must be able to, in principle, work. In other words, it must be a practical method of producing electricity.**

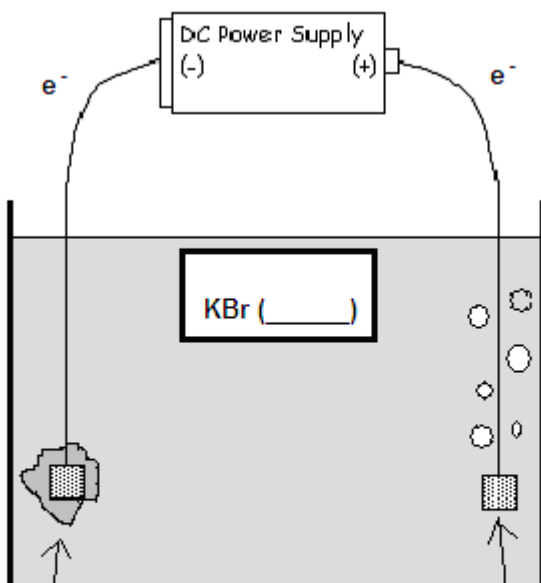
You have FIVE school days in which to complete this extra credit and turn it in, presented NEATLY.

## 4) Using Electricity to Decompose Compounds Homework

### A) Potassium bromide can be decomposed using electrical current.

1) Complete and balance the decomposition reaction:  $\text{KBr} \rightarrow \text{_____} + \text{_____}$

2) Write the oxidation and reduction half-reactions that occur when potassium bromide is decomposed and determine the voltage of this reaction. Then label the diagram, putting arrows to show direction of  $e^-$  flow.



OX: \_\_\_\_\_

$E^0$ : \_\_\_\_\_

RD: \_\_\_\_\_

$E^0$ : \_\_\_\_\_

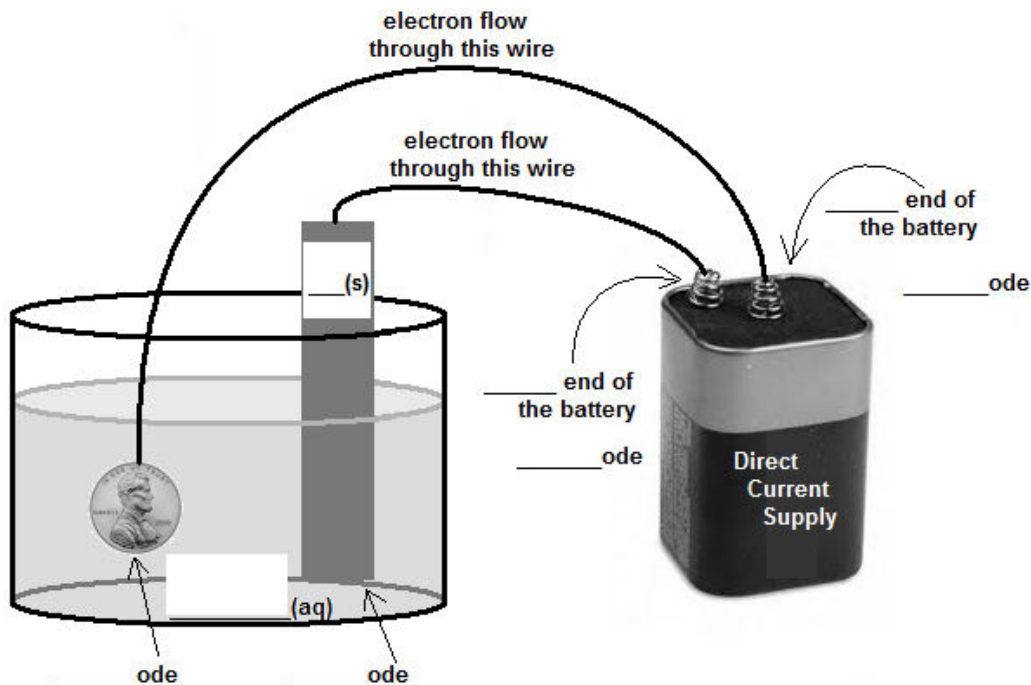
Total Voltage: \_\_\_\_\_

How many 1.5 volt cells will be the minimum required to make this electrolytic cell operate? \_\_\_\_\_

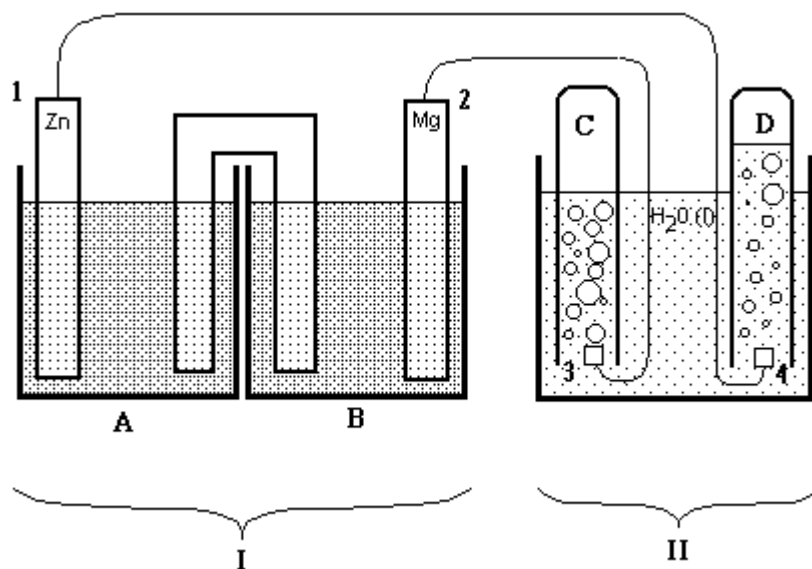
\_\_\_\_\_ forms here at the \_\_\_\_\_ ode  
which is the \_\_\_\_\_ electrode.

\_\_\_\_\_ forms here at the \_\_\_\_\_ ode  
which is the \_\_\_\_\_ electrode.

### B) In this diagram, copper is being plated onto a zinc penny. Please place your answers in the provided spaces. Draw arrows for the direction of electron flow. Choose a suitable soluble compound for the solution.



**C) The diagram below shows an electrolytic cell hooked up to, and being powered by a voltaic cell. A few clues have been included to help you answer the questions below. Reference Table J is also needed.**



\_\_\_\_\_ 1) Which symbol represents the voltaic cell?

- a) A                      b) B                      c) I                      d) II

\_\_\_\_\_ 2) Which symbol represents the electrolytic cell?

- a) C                      b) D                      c) I                      d) II

3) Which metal, Zn or Mg, is listed higher on the Metal Activity Table? \_\_\_\_\_

4) Which metal will undergo oxidation, Zn or Mg? \_\_\_\_\_

**5) Draw the direction that the electrons will go on both wires in the above diagram, using arrows, based on your answer to #4.**

6) Electrons flow from \_\_\_\_\_ to \_\_\_\_\_ and then from \_\_\_\_\_ to \_\_\_\_\_ through the wires.

7) Which symbol represents the anode of the voltaic cell? \_\_\_\_\_

8) Which symbol represents where H<sub>2</sub> (g) is formed? \_\_\_\_\_

9) Which symbol represents the cathode of the electrolytic cell? \_\_\_\_\_

10) Which symbol represents the cathode of the voltaic cell? \_\_\_\_\_