



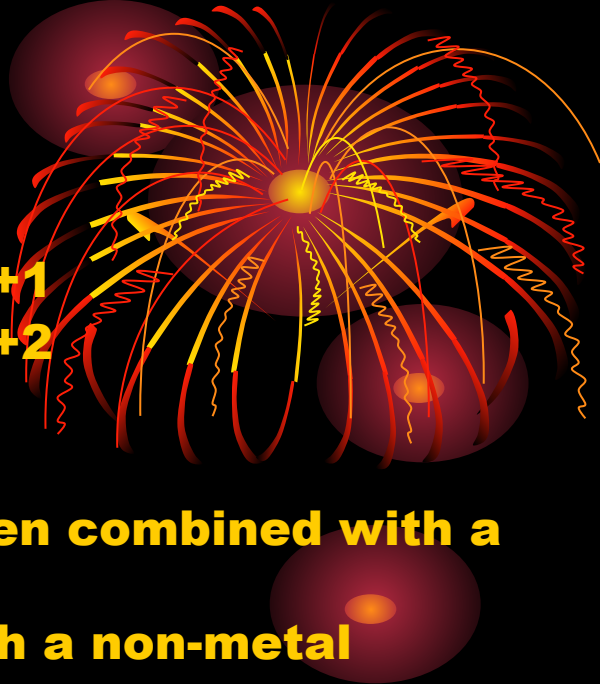
# Chapter 19 & 20

## Oxidation and Reduction “Redox”

# Rules for Assigning Oxidation Numbers



- 1. The CATION is listed first and has a positive oxidation number (charge)**
- 2. The ANION is listed second and has a negative oxidation number (charge)**
- 3. The oxidation # of an atom in an uncombined element = 0**
  - when an element is by itself or diatomic, it has a charge of 0
  - ex/  $\text{Mg}_{(s)}$ ,  $\text{Cl}_{2(g)}$ ,  $\text{O}_{2(g)}$ ,  $\text{Hg}_{(l)}$  all have a charge of 0
- 4. If it is an ion, its ionic charge is its oxidation #.**
  - ex/  $\text{NaCl} \square \text{Na}^+ + \text{Cl}^-$   
          +1     -1



**5. Many charges relate to their group #'s**

**a) elements in group 1 have an oxidation # = +1**

**b) elements in group 2 have an oxidation # = +2**

**c) Aluminum is always +3**

**d) Fluorine is always - 1**

**- the other elements in group 17 are -1 when combined with a metal or hydrogen.**

**e) Hydrogen is always +1 when combined with a non-metal  
it is -1 when combined with a group 1 metal (ex/ NaH)**

**f) Oxygen is always -2 except when in a peroxide, it is -1 Ex/  $H_2O_2$   
and +2 in  $OF_2$**

**6. Oxidation #'s are per atom**

**- ex/  $MgCl_2$  Mg = +2 and Cl = -1 and Cl = -1**

**7. For a compound, the sum of the oxidation #'s will add up to 0**

**- ex/  $BaF_2$  Ba = +2 and F = -1 and F = -1 (remember rule 6)**

$$2 + -1 + -1 = 0$$

**8. For a polyatomic ion, the sum of the oxidation #'s = the charge of the polyatomic ion**

- ex/  $\text{SO}_4^{-2}$

**O = -2 and there are 4 of them**

**so the total charge of the four oxygens = -8 (4 x -2)**

**to figure out the sulfur, we must use the total charge of the ion**

**Sulfur + 4 (oxygens) = -2**

**X + 4 (-2) = -2**

**X = +6**

**Sulfur = +6**

**9. Compounds containing a polyatomic ion must have a total charge of 0 (the same as rule 7)**

ex/  $\text{K}_2\text{CrO}_4$  Find the oxidation numbers of each element

**K = +1 (see rule 5a)**

**since we have 2 K ions, the total charge of the K's is +2**

**O = -2 (see rule 5f)**

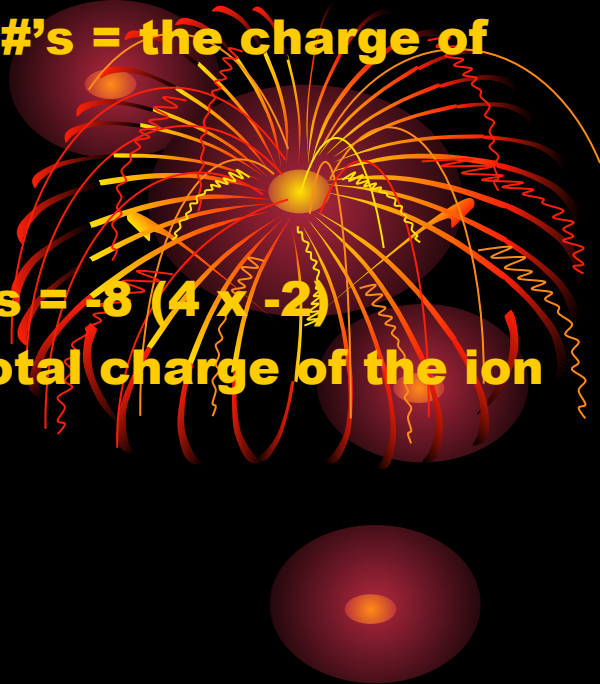
**since we have 4 O ions, the total charge of the O's is -8**

**To figure out the Cr:**

**(2 x K) + Cr + (4 x O) = 0 (see rule 7)**

**( 2 x +1) + Cr + (4 x -2) = 0**

**Cr = +6**



**What is the oxidation #s for each element in the compound**



**K is in group 1, therefore = +1**

**O is always -2 (except in  $H_2O_2$ )**

**To find S, the sum of the oxidation #s must equal 0**

$$2(+1) + X + 4(-2) = 0$$

$$2 + X + -8 = 0$$

$$-6 + X = 0$$

$$X = 6 \quad \square \quad \text{S is +6}$$



**Find the oxidation #'s for each element in the compound  $H_2S_2O_7$**

**H is always +1 when bonded to a non-metal**

**O is always -2 (except in  $H_2O_2$ )**

**To find S, the sum of the oxidation #s must equal 0**

$$2(+1) + 2X + 7(-2) = 0$$

$$2 + 2X + -14 = 0$$

$$-12 + 2X = 0$$

$$2X = 12$$

$$X = 6 \quad \square \quad \text{S is +6}$$

# LEO the LEPer says GER to GENnifer

**L = lose**      **L = lose**

**E = electrons**    **E = electrons**

**O = Oxidation**    **P = Positive**

**G = gain**      **G = gain**

**E = electrons**    **E = electrons**

**R = Reduction**      **N = Negative**

*(Goes down)*



**These are important concepts!!!**

**When an element undergoes OXIDATION, the oxidation number increases □ It gets more POSITIVE (goes up)**

**When an element undergoes REDUCTION, the oxidation number decreases □ It gets more NEGATIVE (goes down)**

**OXIDIZING AGENT – causes something to be oxidized**

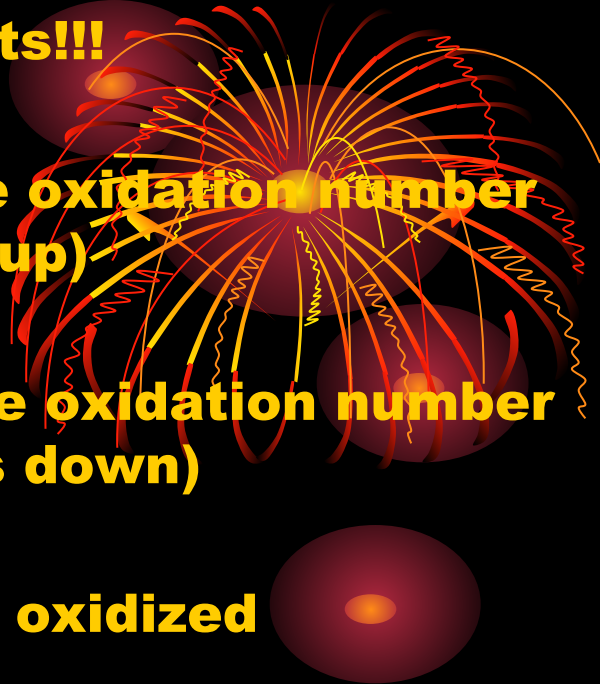
**- It gets REDUCED**

**REDUCING AGENT – causes something to be reduced**

**- It gets OXIDIZED**

**The middle charge of a transition metal can be both a reducing agent and an oxidizing agent because its charge can go up or down.**

**ex/  $\text{Cu}^{+1}$  can go to either  $\text{Cu}_{(s)}$  or  $\text{Cu}^{+2}$**



# Determining what gets: oxidized, reduced, the oxidizing agent & the reducing agent



**Step 1) Write down the oxidation numbers for each element**

**Step 2) Connect the elements on each side of the Rxn. That change**

**Step 3) What gets more negative is what gets reduced and is the oxidizing agent. ALWAYS USE THE REACTANT**

**Step 4) What gets more positive is what gets oxidized and is the reducing agent. ALWAYS USE THE REACTANT**

**A) What gets reduced?  $\text{Cu}^{+2} \square \text{Cu}^0$   
ans.  $\text{Cu}^{+2}$**

**B) What gets oxidized?  $\text{Zn}^0 \square \text{Zn}^{+2}$   
ans.  $\text{Zn}^0$**

**C) What is the reducing agent?  $\text{Zn}^0$   
ans.  $\text{Zn}^0$**

**D) What is the oxidizing agent?  $\text{Cu}^{+2}$   
ans.  $\text{Cu}^{+2}$**



**Ex/ for the following Rxn:**

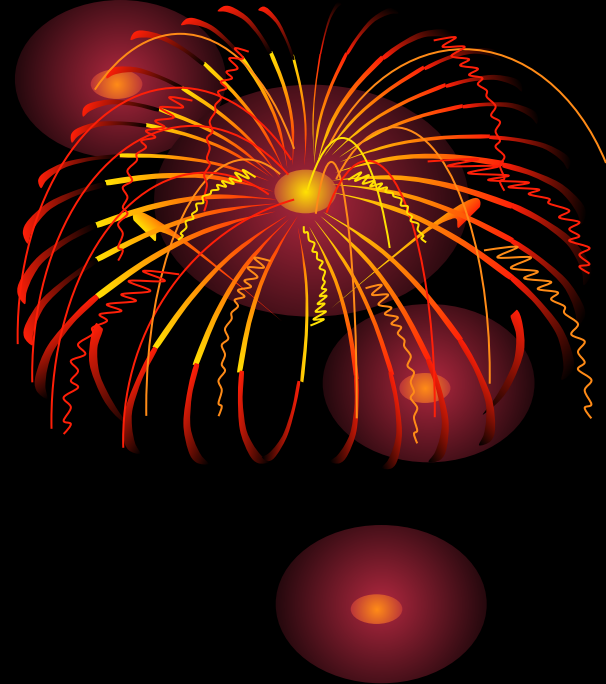


- A) What gets reduced?**
- B) What gets oxidized?**
- C) What is the reducing agent?**
- D) What is the oxidizing agent?**

**Ex/ for the following Rxn:**



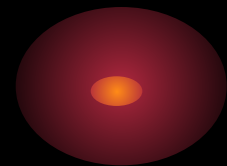
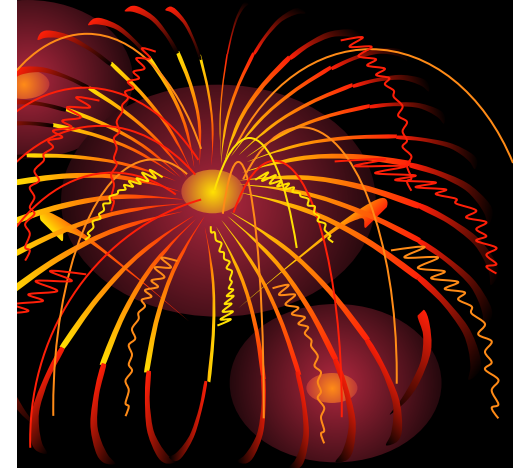
- A) What gets reduced?**
- B) What gets oxidized?**
- C) What is the reducing agent?**
- D) What is the oxidizing agent?**



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

Most	Metals	Nonmetals	Most
	Li	F <sub>2</sub>	Most easily reduced
	Rb	Cl <sub>2</sub>	
	K	Br <sub>2</sub>	
	Cs	I <sub>2</sub>	
	Ba		Most easily oxidized
	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



**Table J is based on the oxidation of Hydrogen gas = 0**

**Most easily oxidized**

**Most easily reduced**

## Using Table J

Table J is used to determine if a Rxn will occur spontaneously.

For Metals: A Rxn will be spontaneous if the substance to be oxidized is above the substance to be reduced.

Ex/ Will the reaction  $\text{Cu}^{+2} + \text{K} \rightarrow \text{Cu} + \text{K}^{+}$  be spontaneous?

$\text{Cu}^{+2} \rightarrow \text{Cu}$  is reduction and  $\text{K} \rightarrow \text{K}^{+}$  is oxidation

Since K is oxidized and above Cu on the table the Rxn will be spontaneous.

Ex/ Will the reaction  $\text{Li}^{+} + \text{Al} \rightarrow \text{Li} + \text{Al}^{+3}$  be spontaneous?

$\text{Li}^{+}$  is reduction

Al is oxidation

Since Al is getting oxidized and is below Li on the table, the Rxn is non-spontaneous.



## Half-Reactions



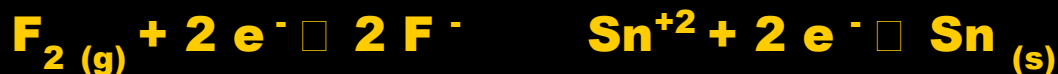
**A half reaction is used to show either a reduction reaction or an oxidation reaction.**

**If the electrons are on the reactant side, it is a reduction  $\frac{1}{2}$  rxn**

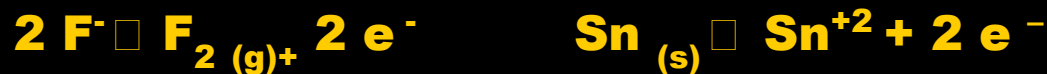
**If the electrons are on the product side, it is an oxidation  $\frac{1}{2}$  rxn**

**(You must still check the charges to make sure they are going up for an oxidation rxn or down for a reduction rxn. This is how I will try to fool you on a test)**

**Reduction  $\frac{1}{2}$  Rxns gain electrons (on the reactant side)**



**Oxidation  $\frac{1}{2}$  Rxns lose electrons (on the product side)**



**Most easily reduced**

**To get an oxidation  $\frac{1}{2}$  Rxn, you must write the reverse Rxn and change the sign of the  $E^\circ$  value**

**Products are Most easily oxidized**

STANDARD ELECTRODE POTENTIALS	
Ionic Concentrations 1 M Water At 298 K, 1 atm	
Half-Reaction	$E^\circ$ (volts)
$F_2(g) + 2e^- \rightarrow 2F^-$	+2.87
$8H^+ + MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O$	+1.51
$Au^{3+} + 3e^- \rightarrow Au(s)$	+1.50
$Cl_2(g) + 2e^- \rightarrow 2Cl^-$	+1.36
$14H^+ + Cr_2O_7^{2-} + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	+1.23
$4H^+ + O_2(g) + 4e^- \rightarrow 2H_2O$	+1.23
$4H^+ + MnO_2(s) + 2e^- \rightarrow Mn^{2+} + 2H_2O$	+1.22
$Br_2(l) + 2e^- \rightarrow 2Br^-$	+1.09
$Hg^{2+} + 2e^- \rightarrow Hg(l)$	+0.85
$Ag^+ + e^- \rightarrow Ag(s)$	+0.80
$Hg_2^{2+} + 2e^- \rightarrow 2Hg(l)$	+0.80
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	+0.77
$I_2(s) + 2e^- \rightarrow 2I^-$	+0.54
$Cu^+ + e^- \rightarrow Cu(s)$	+0.52
$Cu^{2+} + 2e^- \rightarrow Cu(s)$	+0.34
$4H^+ + SO_4^{2-} + 2e^- \rightarrow SO_2(aq) + 2H_2O$	+0.17
$Sn^{4+} + 2e^- \rightarrow Sn^{2+}$	+0.15
$2H^+ + 2e^- \rightarrow H_2(g)$	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
$2H_2O + 2e^- \rightarrow 2OH^- + H_2(g)$	-0.83
$Mn^{2+} + 2e^- \rightarrow Mn(s)$	-1.19
$Al^{3+} + 3e^- \rightarrow Al(s)$	-1.66
$Mg^{2+} + 2e^- \rightarrow Mg(s)$	-2.37
$Na^+ + e^- \rightarrow Na(s)$	-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
$Cs^+ + e^- \rightarrow Cs(s)$	-2.92
$K^+ + e^- \rightarrow K(s)$	-2.93
$Rb^+ + e^- \rightarrow Rb(s)$	-2.98
$Li^+ + e^- \rightarrow Li(s)$	-3.04



**Every  $\frac{1}{2}$  Rxn in this table is a reduction  $\frac{1}{2}$  Rxn**

## Using the Half-Reaction Table

Every Rxn on the table is REDUCTION.

The most easily reduced are at the top, the least easily reduced are at the bottom.

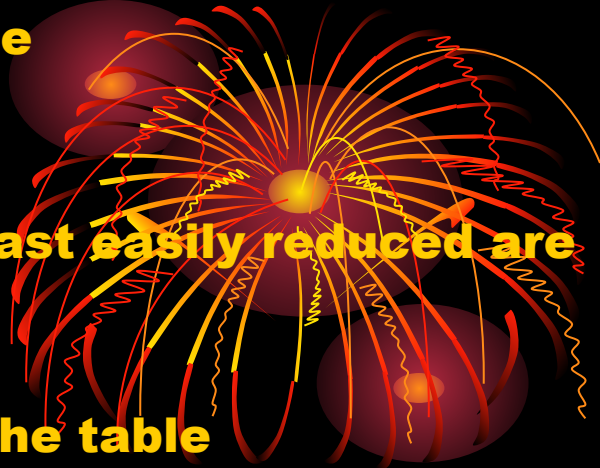
$F_2$  is the most easily reduced element on the table

To find if a Rxn is spontaneous, the element you want to be reduced must be above the element you want to be oxidized.

The element that gets oxidized will be on the product side of the half reaction

The most easily oxidized are at the bottom of the table on the right hand side (product side)

$Li_{(s)}$  is the most easily oxidized element on the table



## Rules for Spontaneous Rxns and the $\frac{1}{2}$ Rxn Table

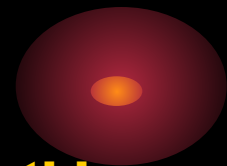
**Substances to the left and those below them on the right react spontaneously with each other.**

**Anything on the left side will react spontaneously with anything below and to the right.**

**$F_2$  reacts spontaneously with  $Cl^-$**

**Anything on the right side will react spontaneously with anything on the left and above it.**

**$Cl^-$  reacts spontaneously with  $F_2$**



**Ex/ Is  $F_2 + CuCl_2 \rightarrow CuF_2 + Cl_2$  spontaneous?**

**Assign oxidation #'s to determine what gets reduced and oxidized.**

**Write the  $\frac{1}{2}$  Rxns. If the reduction  $\frac{1}{2}$  rxn is above the oxidation  $\frac{1}{2}$  rxn, the Overall Reaction will be spontaneous.**

**$F_2$  gets reduced**

**$Cl^-$  gets oxidized (flip the  $\frac{1}{2}$  Rxn**

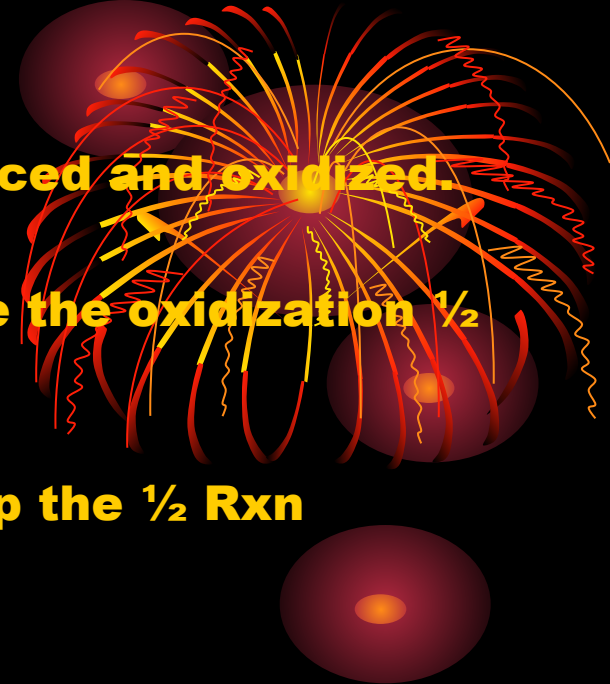


**Since  $F_2$  is reduced and is above  $Cl_2$ , the rxn is spontaneous.**

**Ex/ According to The  $\frac{1}{2}$  Rxn table, which metal will react spontaneously with  $Ag^+$  but not with  $Zn^{+2}$ ?**

**1) Cu      2) Au      3) Al      4) Mg**

**(1) Cu is the only choice between  $Ag^+$  and  $Zn^{+2}$ . On the  $\frac{1}{2}$  Rxn table, it is below  $Ag^+$  and above  $Zn^{+2}$ . On Table J, it is below Zn and above Ag**





**According to Table J, which rxn will take place spontaneously?**

- 1.  $\text{Ni}^{+2} + \text{Pb} \square \text{Ni} + \text{Pb}^{+2}$**
- 2.  $\text{Au}^{+3} + \text{Al} \square \text{Au} + \text{Al}^{+3}$**
- 3.  $\text{Sr}^{+2} + \text{Sn} \square \text{Sr} + \text{Sn}^{+2}$**
- 4.  $\text{Fe}^{+2} + \text{Cu} \square \text{Fe} + \text{Cu}^{+2}$**



**Start by finding what gets oxidized and what gets reduced, then use the table**

- Using table J, what gets oxidized should be above what gets reduced**
- Using the  $\frac{1}{2}$  Rxn table, what gets reduced should be above and on the left (reactant side) and what gets oxidized should be below and on the right (product side)**

# Calculating Spontaneous Rxns using Cell Potential



According to the  $\frac{1}{2}$  Rxn Table, which rxn will take place spontaneously?

1.  $\text{Ni}^{+2} + \text{Pb} \square \text{Ni} + \text{Pb}^{+2}$
2.  $\text{Au}^{+3} + \text{Al} \square \text{Au} + \text{Al}^{+3}$
3.  $\text{Sr}^{+2} + \text{Sn} \square \text{Sr} + \text{Sn}^{+2}$
4.  $\text{Fe}^{+2} + \text{Cu} \square \text{Fe} + \text{Cu}^{+2}$

This same problem can be calculated using the Cell Potential ( $E^0$ )

Steps:

1. Write the Oxidation and Reduction  $\frac{1}{2}$  Rxns
2. Write down their  $E^0$  value from the table, remember, you must change the sign of the  $E^0$  for the oxidation Rxn
3. Add up the  $E^0$  values

A positive  $E^0$  = spontaneous (remember, a  $-\Delta G$  = spontaneous)

A negative  $E^0$  = non-spontaneous (remember, a  $+\Delta G$  = non-spon)

# Electrochemical Cells

## The theory behind an electrochemical cell



1. Redox Rxns are the loss and gain of e-'s.
2. Metals are a "mobile sea of e-'s" □ e-'s can move in a metal
3. Therefore, metals will conduct electricity.
4. When an ionic compound is dissolved in water, it becomes an electrolyte (electrolytes conduct electricity)

Electrochemical Cells - devices that use Redox Rxns to produce electricity

e-'s enter and exit electrochemical cells from electrodes

Electrodes – made of electrical conductors

ex/ Metals or graphite

ex/  $\text{Cu}_{(s)}$  or  $\text{Zn}_{(s)}$

# The RED CAT and AN OX

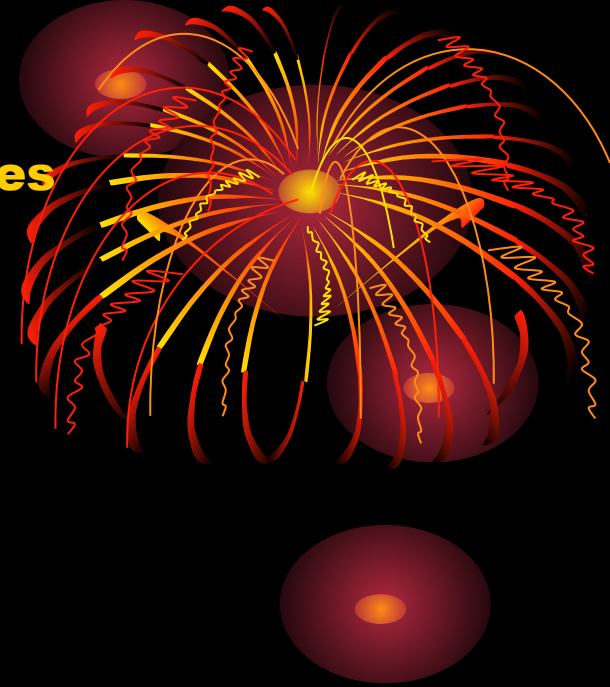
**Oxidation and Reduction Occur at these electrodes**

**In a cell, there are 2 electrodes:**

- 1. Cathode**
- 2. Anode**

**Reduction occurs at the Cathode = RED CAT**

**Oxidation occurs at the Anode = AN OX**



## TYPES OF CELLS

### **1. Electrochemical (Voltaic) (Galvanic) Cells**

**Produce electricity using a spontaneous Rxn  
ex/ Batteries**

### **2. Electrolytic Cells – use electricity to make a non-spon Rxn happen. ex/ Electroplating jewelry, purifying metals**

# Electrochemical Cells

## Contents:

**2 half cells, each one contains an electrode (metal strip) in a solution of the metal's ions.**

***The Cathode is positive***

***The Anode is negative***

**The electrodes are connected by a wire attached to a volt meter.**

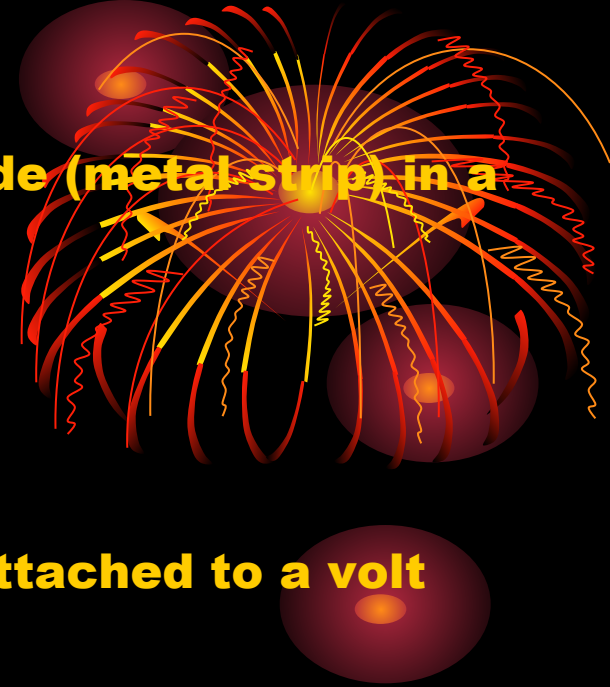
***Electrons pass through the wire from the anode to the cathode!***

**1 salt bridge – The salt bridge connects the half cells**

***The purpose of the salt bridge is to allow ions to flow between the half cells.***

***The ions also flow from the anode side to the cathode side***

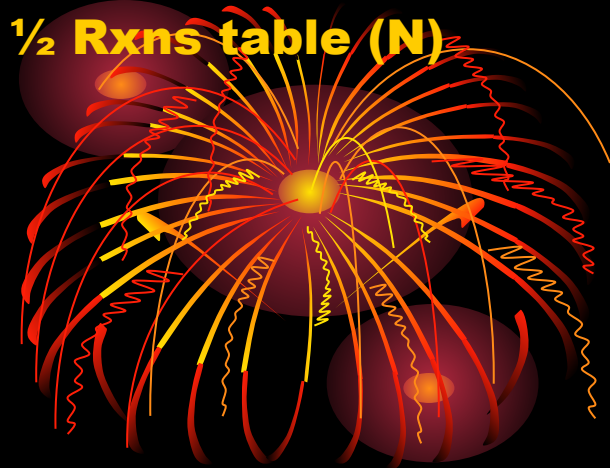
**NO SALT BRIDGE = NO WORK**



**We can calculate the voltage produced using the  $\frac{1}{2}$  Rxns table (N)**

**The voltage produced depends on 3 things**

- 1. [ ] of the solutions**
- 2. types of electrodes**
- 3. pressure if gasses are involved.**



**The voltages in the  $\frac{1}{2}$  Rxn tables are calculated at standard conditions of 1 M solutions and at 1 atm pressure.**

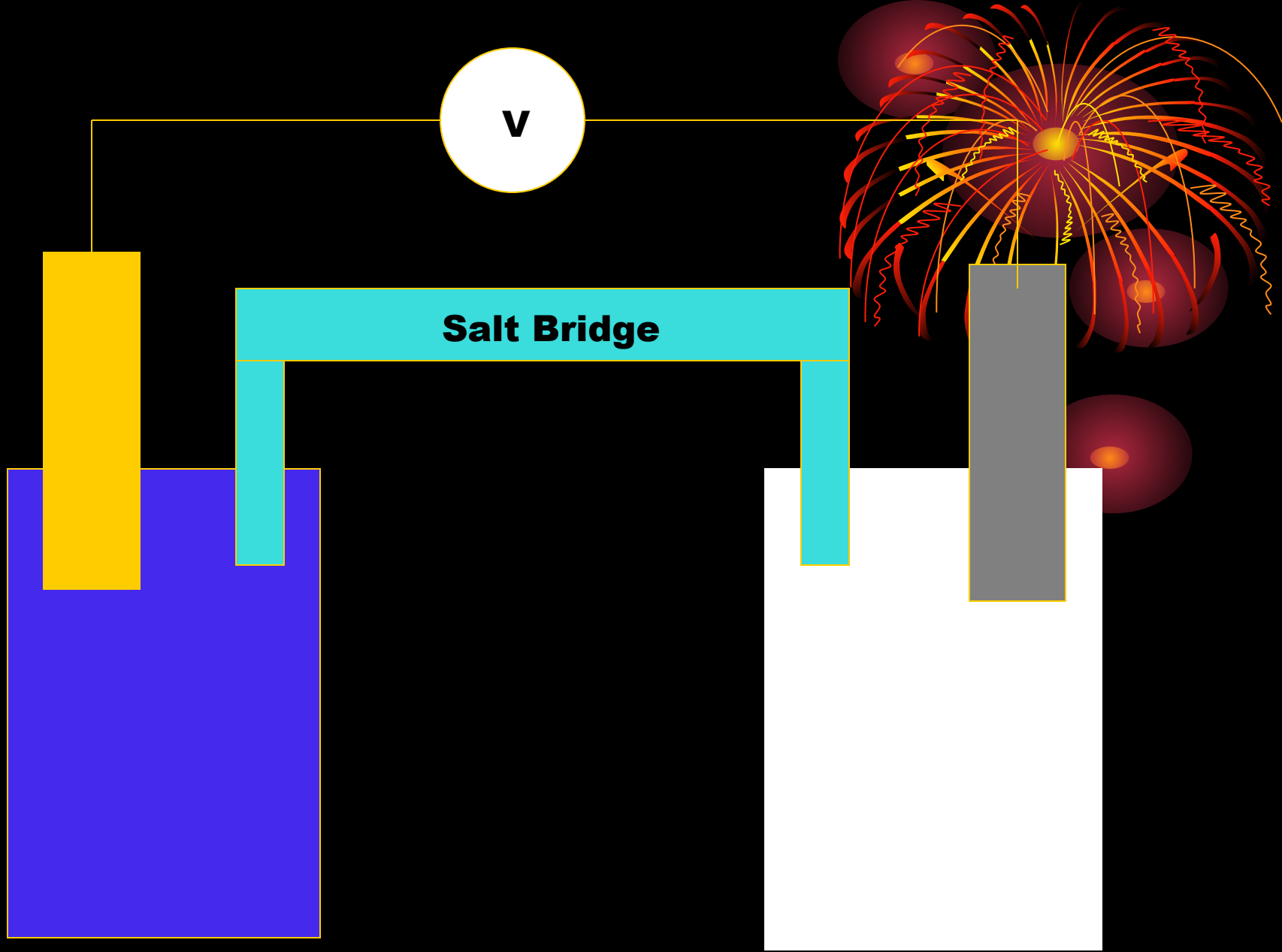


## **Calculating Cell Potential**

**In order to calculate cell potentials it helps to:**

- 1. Draw a picture**
- 2. Write the two  $\frac{1}{2}$  Rxns**
- 3. Flip the oxidation Rxn and Change the sign**
- 4.  $E^0_{\text{total}} = E^0_{\text{Reduction}} + E^0_{\text{Oxidation}}$**

**An electrochemical cell looks like this**



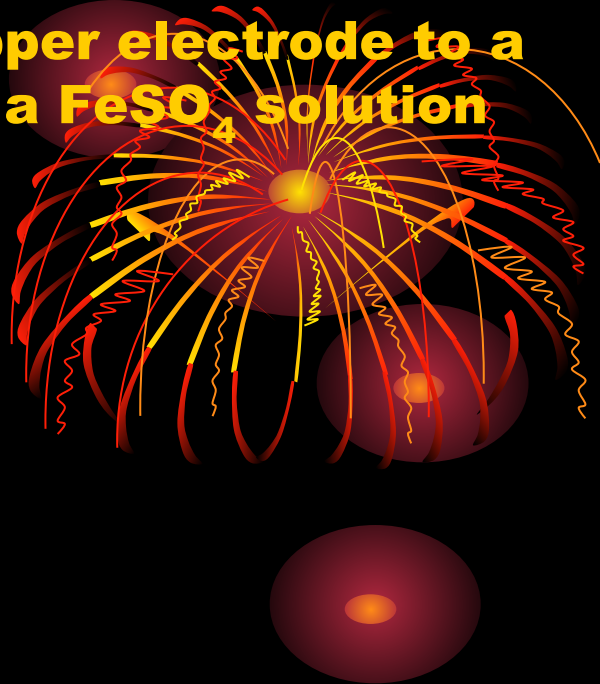
**Ex/ If I connect a half cell of  $\text{Ni}(\text{NO}_3)_2$  w/ a Nickel electrode to a half cell consisting of a Silver electrode in a  $\text{AgNO}_3$  solution with a salt bridge:**



- 1. What gets oxidized?**
- 2. What gets reduced?**
- 3. Which electrode is the anode?**
- 4. Which electrode is the cathode?**
- 5. Which electrode is the + electrode?**
- 6. Which electrode is the - electrode?**
- 7. Which electrode gains mass?**
- 8. Which electrode loses mass?**
- 9. The electrons flow from \_\_\_\_\_ to \_\_\_\_\_ ?**
- 10. The ions flow from \_\_\_\_\_ to \_\_\_\_\_ ?**
- 11. What is the cell potential?**



**Ex/ If I connect a half cell of  $\text{CuSO}_4$  w/ a Copper electrode to a half cell consisting of an Iron electrode in a  $\text{FeSO}_4$  solution with a salt bridge:**



- 1. What gets oxidized?**
- 2. What gets reduced?**
- 3. Which electrode is the anode?**
- 4. Which electrode is the cathode?**
- 5. Which electrode is the + electrode?**
- 6. Which electrode is the – electrode?**
- 7. Which electrode gains mass?**
- 8. Which electrode loses mass?**
- 9. The electrons flow from \_\_\_\_\_ to \_\_\_\_\_?**
- 10. The ions flow from \_\_\_\_\_ to \_\_\_\_\_?**
- 11. What is the cell potential?**

## Electrolytic Cells (Electrolysis)

**Electrolytic Cells use electricity to force a non-spontaneous Rxn to happen.**

**This means that the  $E^0$  will always be negative.**

**What gets Oxidized will be the Rxn on top of the  $\frac{1}{2}$  Rxn chart and must be flipped.**

**What gets Reduced will be below and will be written as is.**

### **HOW TO RECOGNIZE AN ELECTROLYTIC CELL:**

- 1. Both electrodes are in the same solution**
- 2. The electrodes are hooked to a battery or a power source.**
- 3. Often, there is a fork or a spoon.**

**In an electrolytic cell:**

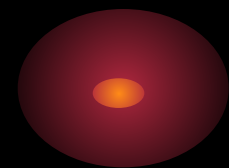
**THE ANODE IS +**

**THE CATHODE IS -**



# Uses of electrolysis:

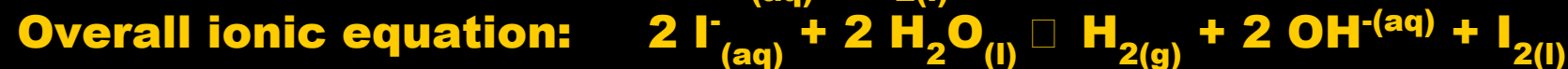
1. To electroplate jewelry or eating utensils.
2. Breaking a compound into its elements
3. Rechargeable Batteries
4. Car Batteries (Lead Storage Battery)



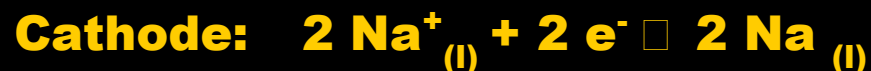
## Ex/ Electrolysis of Water



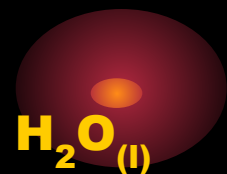
## Ex/ Electrolysis of KI<sub>(aq)</sub>



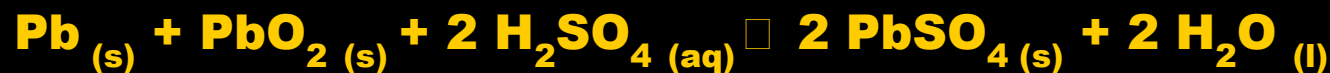
## Electrolysis of Molten NaCl:



## Lead Storage Battery (Car Battery)



## Overall Rxn when discharged (Used to start the car):



## Overall Rxn when the battery is recharging (engine is running):



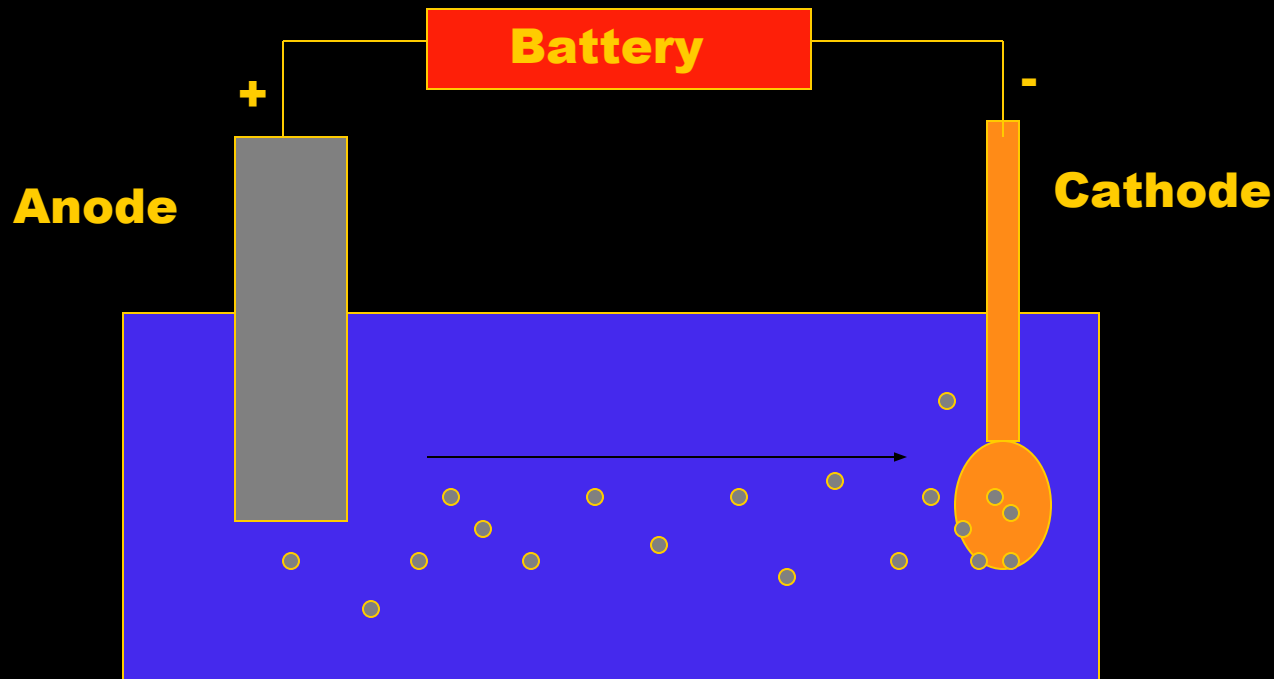
# Electroplating

The object to be plated must be at the cathode (jewelry or fork)

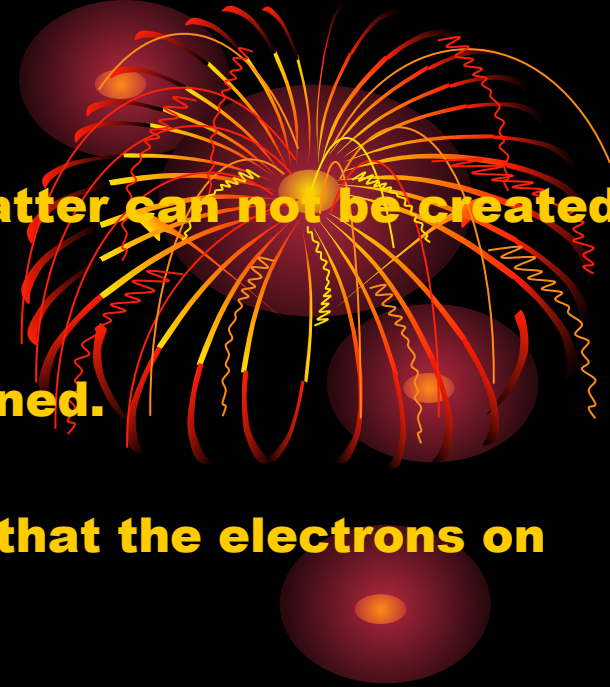
The metal you want to plate with (ex/ Ag or Au) will be the anode and must be in a solution of the same ions (ex/  $\text{Ag}^+$  or  $\text{Au}^{+3}$ )

**Cathode:**  $\text{Ag}^+_{(\text{aq})} + \text{e}^- \rightarrow \text{Ag}_{(\text{s})}$

**Anode:**  $\text{Ag}_{(\text{s})} \rightarrow \text{Ag}^+ + \text{e}^-$



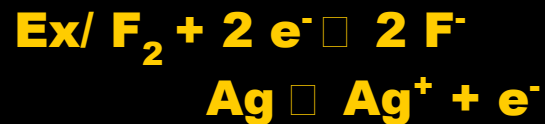
## Balancing Redox Rxns



**The Law of Conservation of Matter states that Matter can not be created or destroyed.**

**Therefore, # of e<sup>-</sup>'s lost must equal # of e<sup>-</sup>'s gained.**

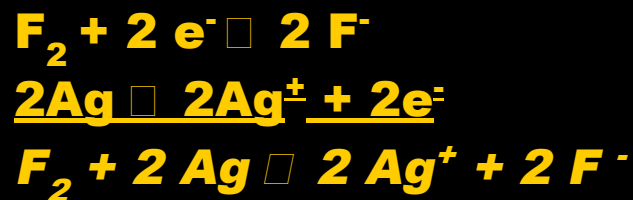
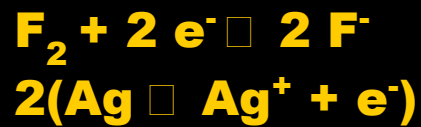
**Using 1/2 Rxns: multiply all of the coefficients so that the electrons on both sides become equal**



**We are gaining 2 electrons but only losing 1 electron. This is a bad thing**

**We try to find a common denominator for the electrons and multiply accordingly.**

**Common denominator is 2, so we must multiply the Ag Rxn's coefficients by 2**



**Ex/ Balance the Redox  $\frac{1}{2}$  Rxns and write the complete balanced Rxn**

