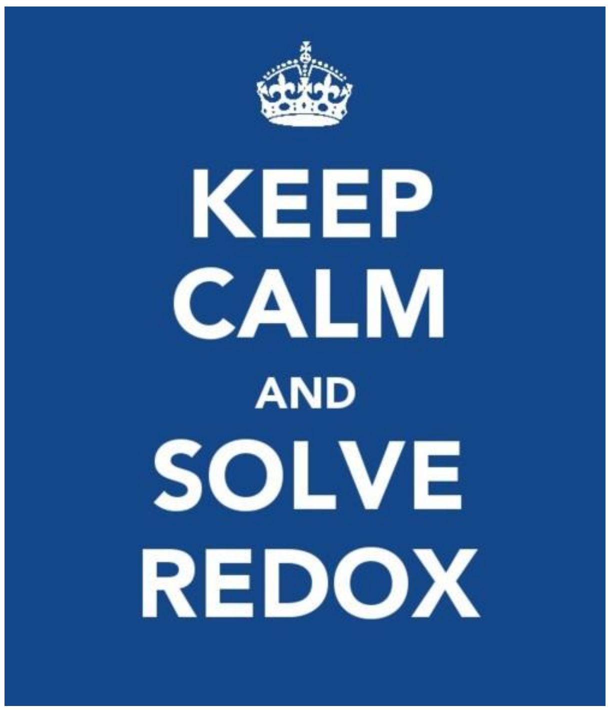
Name:

Regents Chemistry:

Notes: Unit 12 Electrochemistry



Name:

KEY IDEAS:

- An oxidation-reduction (redox) reaction involves the transfer of electrons (e-). (3.2d)
- Reduction is the gain of electrons. (3.2e)
- A half-reaction can be written to represent reduction. (3.2f)
- Oxidation is the loss of electrons. (3.2g)
- A half-reaction can be written to represent oxidation. (3.2h)
- In a redox reaction the number of electrons lost is equal to the number of electrons gained. (3.3b)
- Oxidation numbers (states) can be assigned to atoms and ions. Changes in oxidation numbers indicate that oxidation and reduction have occurred. (3.2i)
- An electrochemical cell can be either voltaic or electrolytic. In an electrochemical cell, oxidation occurs at the anode and reduction at the cathode. (3.2j)
- A voltaic cell spontaneously converts chemical energy to electrical energy. (3.2k)
- An electrolytic cell requires electrical energy to produce chemical change. This process is known as electrolysis. (3.2l)

Lesson 1: Oxidation Numbers

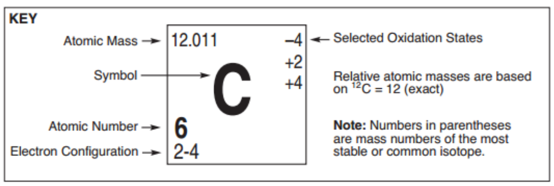
> ***Objective:***
> - ***Determine the oxidation numbers of atoms and ion is a chemical reaction***

Problem: What is electricity and how is it formed?

In electrochemical reactions, electrons are transferred from one species to another.

In order to keep track of what loses electrons and what gains them, we assign

____.



<u>Rules for Assigning Oxidation Numbers</u>

1. Uncombined elements (not combined with any other element) have an oxidation # of ZERO (this includes diatomic elements)
   Example: $Na = 0$
   $Cl_2 = 0$
2. If an element has only one charge listed on the periodic table, then that is its oxidation number *when in a compound*.
   Example: Group 1 are always +1 and group 2 are always +2
3. If a nonmetal atom is the negative ion in an ionic compound, then the top charge listed is its oxidation state *(recall from building ionic formulas from names)*
   Examples: *HF*
   NaCl
4. H is +1 in the front and -1 in the back.
   Examples: H_3P *P is -3* *H is +1*
   LiH *Li must be +1* *H is -1*

THIS IS A GENERAL RULE…H with a metal acts as a nonmetal (-1); otherwise +1

5. **The Sum of the oxidation #'s in a compound must = ZERO.**
   Be sure to multiply the oxidation # by the # of atoms indicated by the subscript.

   Example: $CaCl_2$

6. **Sum of the oxidation # in an ion must equal the charge of the ion listed**

Example: $Cr_2O_7^{2-}$

Cr: 2(+6) = +12

O: 7(-2) = -14

-2

7. If an element has more than one charge listed, use the other charges to figure it out.

- Assign oxidation number to elements with 1 oxidation # first.
- Assign it an oxidation number which will make the compound = ZERO

Example: $FeCl_3$

Example: Determine the Oxidation Number of each atom

Sn	= 0
N_2	= 0
Cl^-	= -1
Ca^{+2}	= +2

LiF	Li= +1 F=-1
$NaBrO_3$	Na= +1 O= -2 Br= +5
$MgSO_4$	Mg= +1 O= -2 S= +6
$CaClO_3$	Ca= +2 O= -2 Cl= +4

NOT IN VIDEO – will cover in class

Exceptions to the rule (oxygen):

Oxygen is -2 except when bonded to fluorine

Example: OF_2 F must be -1 O is +2

Oxygen is -1 in peroxides

Example: K_2O_2 each K is +1 O is -1

Lesson 2: Identifying a REDOX Reaction

> ***Objective:***
> - ***Determine if a reaction is a Redox reaction***
> - ***Determine which species is oxidized and which is reduced***

REDOX Reactions = Reactions involving the transfer of electrons

OXIDATION: LOSS OF ELECTRONS by an atom or ion

- OXIDATION NUMBER goes UP/INCREASES because negative charges (electrons) are lost
- Becomes more (+) charged

REDUCTION: GAIN OF ELECTRONS by an atom or ion

- OXIDATION NUMBER goes DOWN/REDUCES because negative charges (electrons) are gained
- Becomes MORE (-) charged

EXAMPLE:

$$\underset{0}{Zn}(s) + 2\,\underset{+1}{H^{+}}(aq) \longrightarrow \underset{+2}{Zn^{2+}}(aq) + \underset{0}{H_2}(g)$$

- Zinc loses two electrons (oxidized) to go from neutral zinc metal to the Zn^{2+} ion
- Each of the H^+ gains an electron (reduced) and they combine to form H_2

Trick: **LEO** (the lion says…) **GER**

LEO:

GER:

Lesson 2: Identifying a REDOX Reaction

EXAMPLE: Oxidation #'s can change as a result of a rxn.

$2Na + Cl_2 \rightarrow 2NaCl$

Na went from 0 to +1: Na is OXIDIZED

Cl went from 0 to -1: Cl is REDUCED

EXAMPLE: Oxidation #'s can change as a result of a rxn.

$H_2 + O_2 \rightarrow 2H_2O$

H went from 0 to +1: H is OXIDIZED

O went from 0 to -2: O is REDUCED

<u>*Tip: For Polyatomic Ions</u>**

- If the polyatomic ion is unchanged (single or double replacement), you can use the charge of the entire ion (Look up charge on TABLE E)
- This is the oxidation # (easier than assigning oxidation numbers to each individual atom)
- Ex. NO_3^{-1}

SPECTATOR IONS: The ion that does have a change in oxidation #

- $Cu^0 + 2Ag^{+1}NO_3^{-1} \rightarrow Cu^{+2}(NO_3)_2^{-1} + 2Ag^0$
- NO_3^- is the spectator ion

Charges added here to compounds to show oxidation states – remember compound formulas as normally written do not include any charges!!

EXAMPLE:

$Li + NaF \rightarrow Na + LiF$

$K_2O + Li \rightarrow Li_2O + K$

$2Na + Cl_2 \rightarrow NaCl$

$LiClO_3 \rightarrow LiClO + O_2$

Lesson 2: Identifying a REDOX Reaction

ARE THESE REDOX REACTIONS?

$NaClO_3 \rightarrow 2NaCl + 3O_2$	Yes
$Li + F_2 \rightarrow LiF$	Yes
$Mg + 2HNO_3 \rightarrow Mg(NO_3)_2 + H_2$	Yes
$NaOH + LiBr \rightarrow LiOH + NaBr$	No

TRICK:

- Single Replacement Rx's are <u>ALWAYS</u> Redox

  $Zn(s) \quad + \quad HCl(aq) \quad \rightarrow \quad H_2(g) \quad + \quad ZnCl_2\,(aq)$

- Double Replacement Rx's are <u>NEVER</u> redox

  $NaOH_{(aq)} \quad + \quad HCl_{(aq)} \quad \rightarrow \quad H_2O_{(l)} \quad + \quad NaCl_{(aq)}$

NOT IN VIDEO – WILL DISCUSS IN LATER LESSON:

CAN THESE REACTIONS HAPPEN SPONTANEOUSLY?

- **SPONTANEOUS REACTION** = occurs w/out adding energy to system
- If the "single" element is more active than the "combined" element, the reaction will be spontaneous.

  Ex: $Zn \quad + \quad PbCl_2 \quad \rightarrow \quad ZnCl_2 \quad + \quad Pb$

- **NON SPONTANEOUS REACTION** = Reaction WILL NOT occur unless energy is added to system

  Ex: $Zn \quad + \quad AlCl_3 \quad \rightarrow$ No rxn

EXAMPLE:

Comparing nickel and aluminum. Which would be oxidized and which would be reduced? (the higher element on Table J always gets to do what it wants, metals to lose electrons, non-metals to gain them.

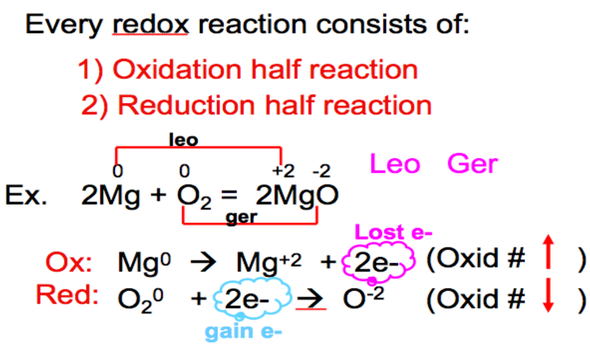
Oxidized: Al

Reduced: Ni

Objective:

- ***Compose redox half reactions***
- ***Construct a balanced redox reaction***

HALF REACTIONS: Show either the oxidation or reduction portion of a redox reaction, including the **electrons gained or lost**.



Careful…To balance the O_2 reduction, remember you need 2 electrons per oxygen atom and you end up with 2 oxygen ions.

Rules for Setting Up Half Reactions:

1. Assign oxidation numbers to all elements in reaction and determine if it is a redox reaction (look for the change in oxidation # of 2 elements
2. Determine which species is oxidized (loses electrons) and which is reduced (gains electrons) [use brackets].
3. Then break the overall reaction into oxidation and reduction reactions called HALF REACTIONS by pulling out brackets. ***spectator ions (those that do not change oxidation states) are left out of the half reactions.
4. Fill in electrons for each half reaction:

   oxidation: electrons (lost) on right side

   reduction: electrons (gained) on left side
5. Check to see if charges are equal on both sides of each half reaction

EXAMPLE:

red

$$\overset{0}{Cu} + 2\overset{+1}{Ag}\overset{-1}{NO_3} \rightarrow \overset{+2}{Cu}\overset{-1}{(NO_3)_2} + 2\overset{0}{Ag}$$

ox

ox: $Cu^0 \rightarrow Cu^{+2} + 2e^-$

Red: $Ag^{+1} + 1e^- \rightarrow Ag^0$

EXAMPLE: Set up both half reactions for the following reaction:

$Ca(s) + Cu^{2+}(aq) \rightarrow Ca^{2+}(aq) + Cu(s)$

Balancing Redox Reactions

In All redox reactions there is a **CONSERVATION** of:

- **MASS**
- **CHARGE**

Steps to balance redox reactions:

1. Write out the 2 half reactions (including electrons)
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 original equation

EXAMPLE:

$$\overset{0}{Cu} + \overset{+1}{Ag}\overset{-1}{NO_3} \rightarrow \overset{+2}{Cu}\overset{-1}{(NO_3)_2} + \overset{0}{Ag}$$

red

ox

$$Cu^{0} \rightarrow Cu^{+2} + 2e-$$

Red: $2(Ag^{+1} + 1e- \rightarrow Fe^{0})$ $\quad 2Ag^{+1} + 2e- \rightarrow Fe^{0}$

$$Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$$

EXAMPLE: Balance this Reaction

$Ca(s) + Cu^{2+}(aq) \rightarrow Ca^{2+}(aq) + Cu(s)$

EXAMPLE: Balance this Reaction

$K + ZnCl_2 \rightarrow KCl + Zn$

Lesson 4: Electrochemical Cells (Voltaic Cell or Battery)

Objective:
- ***Determine the flow of electrons in a battery (voltaic cell)***
- ***Identify the anode and cathode in a voltaic cell***

Problem: What is electricity and how is it formed?

- In a spontaneous redox rxn, electrons are transferred and ENERGY is released.
- This energy can be used to do work is a voltaic cell

Voltaic Cell (battery)

- SPONTANEOUS redox reaction
  - Converts chemical energy into electrical energy
- The flow of electrons produces electricity

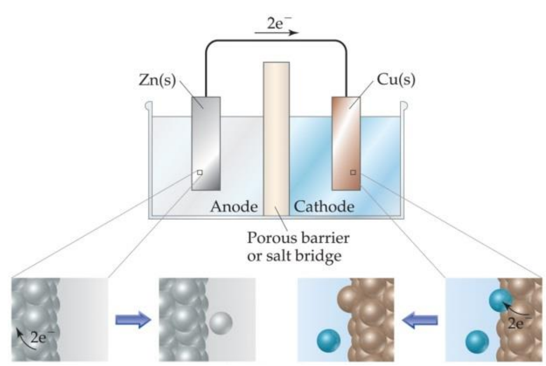
How does a voltaic cell work?

- Electrons flow SPONTANEOUSLY from the **ANODE** (more active metal) to the **CATHODE** (less active metal) *[see Table J]*
- Anode: Negative electrode where oxidation (loss of e-) occurs
- Cathode: Positive electrode where reduction (gain of e-) occurs

Remember: AN OX RED CAT

EXAMPLE:

- Zn is higher on table J so electrons flow from Zn to Cu.
- Zn is anode and Cu is cathode



Lesson 4: Electrochemical Cells (Voltaic Cell or Battery)

Parts of a Voltaic Cell

- **2 half cells** (1 for oxidation and 1 for reduction half reactions)
- **Electrodes** (site of ox and red)
- **Wire** (connects to electrodes- *allows electrons to flow*)
- **Salt Bridge** (allows ions to flow and prevents polarization of cells
- **Voltmeter** (measures electric current)

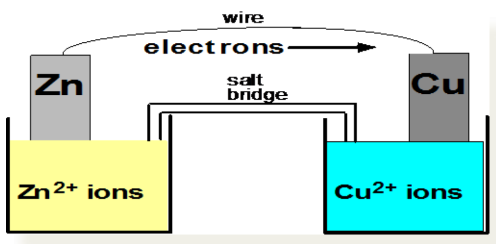
How to Label a Voltaic Cell

1. Determine which electrode is the anode and which is the cathode
2. Identify where oxidation and reduction take place
3. Determine the direction of electron flow

   ****Remember:
   - Red Cat
   - An Ox

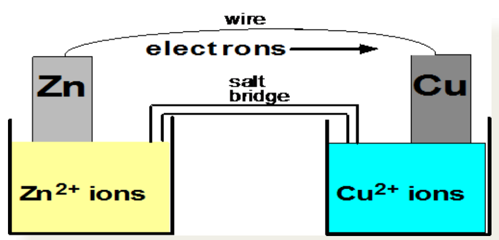
Using table J to determine flow of electrons

- More active metal is oxidized
- ELECTRONS flow from **HIGH** to **LOW**



What happens to the electrodes as electrons flow?

- Cathode (Cu) increases in mass (FAT RED CAT)
- Anode (Zn) decreases in mass (EATS AN OX)



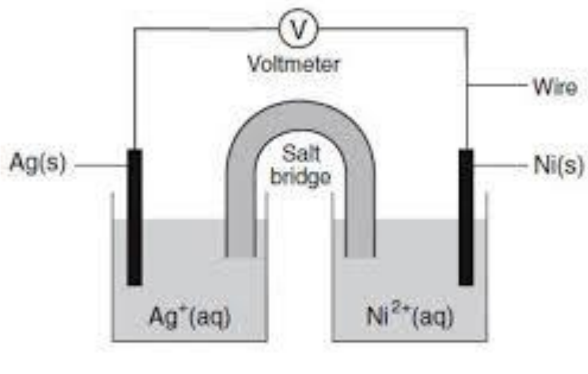
Function of the Salt Bridge

- The salt bridge allows for the flow of ions to prevent a build up of charge at each electrode
- Positive ions flow to cathode
- Negative ions flow to anode

Removing the salt bridge would result in:

- The voltage will go to zero because electrons stop flowing.

EXAMPLE: Label the following: anode, cathode, where ox and red take place and direction of e- flow and which electrode increases in mass and which decreases in mass

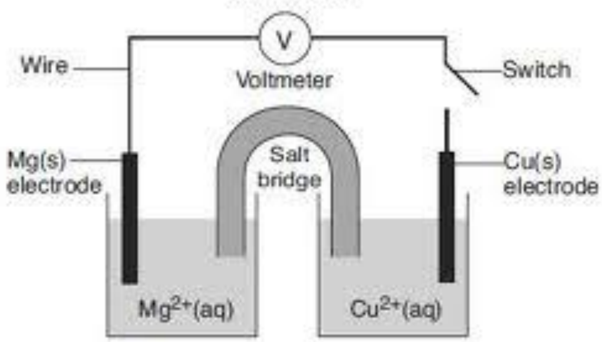


$2Ag^{+}(aq) + Ni(s) \rightarrow 2Ag(s) + Ni^{2+}(aq)$

- Ni is oxidized (anode)
- Ag is reduced (cathode)
- Electrons flow from Ni to Ag
- Ag electrode increases in mass and Ni decreases in mass

EXAMPLE: Label the following: anode, cathode, where ox and red take place and direction of e- flow and which electrode increases in mass and which decreases in mass

Voltaic Cell



$Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$

- Mg is oxidized (anode)
- Cu is reduced (cathode)
- Electrons flow from Mg to Cu
- Cu electrode increases in mass and Mg decreases in mass

Lesson 5: Electrochemical Cells (Electrolytic Cells)

Objective:

- ***Identify an electrolytic cell***
- ***Differentiate between an electrolytic cell and a voltaic cell***

ELECTROLYTIC CELL:

- **Non-Spontaneous** reaction
- Outside **power source** must be supplied to transfer electrons
- Converts **electrical energy** into **chemical energy.**
- Oxidation still takes place at the anode and reduction at the cathode and electrons travel from anode to cathode, **SO** their **CHARGES** are **REVERSED** (cathode – and anode +)

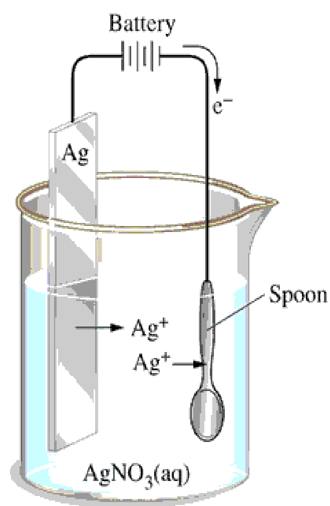
ELECTROPLATING:

- The process of adding a layer (plate) of metal on the surface of another object.

Ex. Gold plated jewelry & Chrome Bumpers

ELECTROPLATING PROCESS:

- The battery forces electrons to travel to the spoon.
- The spoon is negative and will attract Silver (+) ions.
- The silver ions will reduce (stick) onto the spoon, plating it.
- The Ag bar anode loses e- and will eventually disappear.



At the anode:

$Ag \rightarrow Ag^+ + 1e^-$

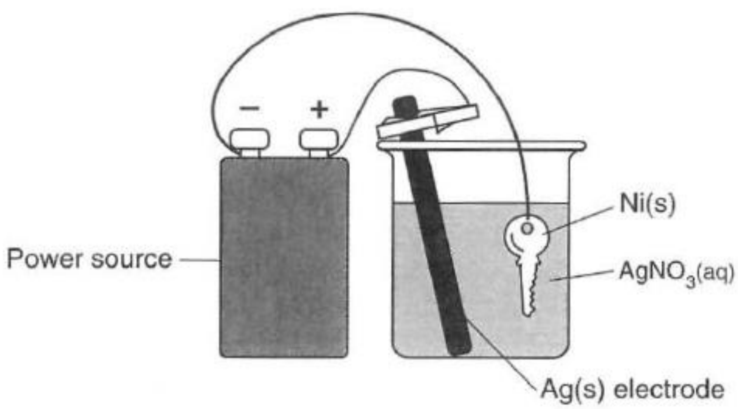
At the cathode:

$Ag^+ + 1e^- \rightarrow Ag$

COMPARISON BETWEEN ELECTROLYTIC AND VOLTAIC CELLS

Voltaic	Electrolytic
An Ox, Red Cat	An Ox, Red Cat
Electrons flow A→ C	Electrons flow A→ C
Spontaneous	Not Spontaneous
Makes electric	Needs electric
Salt bridge	No Salt bridge, same container
Anode -, cathode +	Anode +, cathode –
Chemical energy spontaneously changes to electrical energy	Electric energy changes to chemical energy

EXAMPLE:



1. **Identify the cathode in the cell.**

   The Ni key

2. **What is the purpose of the battery?**

   To force a nonspontaneous reaction