

Ms. Randall
Regents Chemistry
Unit 11: Electrochemistry (RedOx)

Unit Objectives:

- Define and identify oxidation reactions
- Define and identify reduction reactions
- Assign oxidation numbers to elements in a compound
- Write and balance half reactions
- Identify oxidizing agents and reducing agents
- Distinguish between voltaic and electrolytic cells
- Identify the components of an electrochemical cell
- Indicate the direction of electrons and ions through an electrochemical cell
- Determine, using Table J, whether a reaction is spontaneous or not

Focus Questions for the Unit:

- How does transfer of electrons relate to changes in oxidation states?
- How can a chemical reaction be used to create electrical energy?

YOU SHOULD BE ABLE TO ANSWER THESE IN DETAIL BY THE END OF THE UNIT

Define the following vocabulary:

Redox
Reduction
Oxidation
Reducing Agent
Oxidizing Agent
Oxidation Number
Half reaction
Electrode
Voltaic Cell
Salt bridge
Electrochemical Cell
Electrolytic Cell
Anode
Cathode

Lesson 1: Oxidation States Review

Objective:

- **Determine the oxidation numbers of atoms and ion in a chemical reaction**

Oxidation numbers (states)

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

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

Poly atomic ions also have oxidation numbers (as ions it's the net ionic charge listed on table E) and as ions you should be able to determine the oxidation numbers of the parts of the whole polyatomic ion.

Selected oxidation numbers: (refer to your periodic table now).

Our key element at the top of the page is carbon. Top right corner of the box shows three selected oxidation numbers. They are -4 , $+2$, and $+4$. There are others, but in our class we'll only use the selected oxidation states on our periodic table.

All group 1 atoms have a +1 oxidation state.

All group 2 atoms have a +2 oxidation state.

**Transitional metals have one or more possible oxidation states
(hence the roman numerals in naming some transitional metal ionic compounds).**

Most of the nonmetals have many possible oxidation states, both positive or negative.

Noble gases have a "0" since they do not make any compounds (except Kr and Xe).

Let's look at these molecules and formula units and figure out the oxidation numbers for each atom or ion inside.

compound or ion	net charge	oxidation numbers of the parts
NaCl	0	Na ⁺¹ Cl ⁻¹
NaOH	0	Na ⁺¹ O ⁻² H ⁺¹
H ₂ SO ₄	0	H ⁺¹ H ⁺¹ S ⁺⁶ O ⁻² O ⁻² O ⁻² O ⁻²
N ₂ O ₅	0	N ⁺⁵ N ⁺⁵ O ⁻² O ⁻² O ⁻² O ⁻² O ⁻²
NH ₃	0	N ⁻³ H ⁺¹ H ⁺¹ H ⁺¹
NH ₄ ⁺¹	+1	N ⁻³ H ⁺¹ H ⁺¹ H ⁺¹ H ⁺¹
MnO ₄ ⁻¹	-1	Mn ⁺⁷ O ⁻² O ⁻² O ⁻² O ⁻²
PO ₄ ⁻³	-3	P ⁺⁵ O ⁻² O ⁻² O ⁻² O ⁻²
HCO ₃ ⁻¹	-1	H ⁺¹ C ⁺⁴ O ⁻² O ⁻² O ⁻²
Cr ₂ O ₇ ⁻²	-2	Cr ⁺⁶ Cr ⁺⁶ O ⁻²

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

Lesson 2: Recognizing Oxidation-Reduction Reactions

Objective:

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

Oxidation-reduction reactions are reactions in which one type of atom increases in oxidation number (is oxidized) and another type of atom decreases in oxidation number (is reduced). Thus to show that a reaction is a redox reaction, you need to determine the oxidation numbers for the atoms in the reactants and products, and decide if any changes in oxidation number value are taking place. There are, however, a few useful generalizations.

- Look for pure elements. If one or more reactants or products are pure elements, the reaction **HAS TO BE** a redox reaction. **WHY?**
- Single Replacement and Combustion reactions are **ALWAYS** redox reactions.
- Synthesis and Decomposition reactions are **USUALLY** redox reactions.
- Double Replacement reactions are **NEVER** redox reactions. This includes acid-base neutralization reactions and most precipitation reactions are not redox reactions.

Oxidation-Reduction (REDOX) reactions are the principle source of energy on earth.

In fact, most chemical reactions are REDOX reactions, which is simply to say what you already know... most chemistry is the result of the competition between elements and compounds for **ELECTRONS!** For example:

- Burning of fossil fuel for energy
- Burning wood, gas, oil for heat
- “Burning” food in our bodies

These oxidations are all accompanied by reductions. There are many other examples of **REDOX** reactions:

- Batteries produce electricity
- Metals corrode and can be regenerated
- Metals can be electroplated to retard corrosion or enhance beauty
- Transformation of organic material into coal and oil over millions of years is an example of a reduction

Let's look in more detail:

REDuction – OXidation Reactions (AKA Redox): reactions that involve the **TRANSFER OF ELECTRONS**; both reduction and oxidation *must* happen **SIMULTANEOUSLY!**

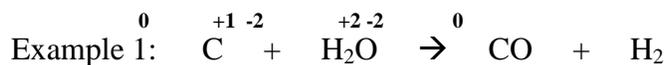
Reduction = **GAIN OF ELECTRONS** by an atom or ion
-OXIDATION NUMBER goes **DOWN/REDUCES**
-Nonmetals **GAIN** electrons

Oxidation = **LOSS OF ELECTRONS** by an atom or ion
- OXIDATION NUMBER goes **UP/INCREASES**
-Metals **LOSE** electrons.

A reaction is REDOX if...OXIDATION NUMBERS CHANGE FOR 2 ELEMENTS WITHIN A REACTION

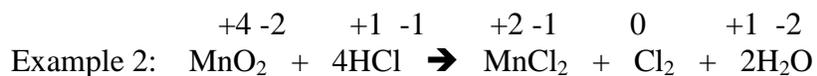
Oxidation = occurs where OXIDATION NUMBER INCREASES for one specific element from reactant to product side of reaction; LEO or OIL

Reduction = occurs where OXIDATION NUMBER DECREASES for one specific element from reactant to product side of reaction; GER or RIG



C = oxidation (charge goes up; electrons are lost)

H⁺¹ = reduction (charge goes down; electrons gained)



Cl⁻¹ undergoes oxidation (charge goes up; electrons are lost)

Mn⁺⁴ undergoes reduction (charge goes down; electrons gained)

Lesson 3: Half reactions

Objective:

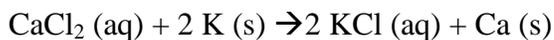
- *Compose redox half reactions*
- *Construct a balanced redox reaction*

Watch the tutorial on my website: [Chemistry tutorial 12.1d half reactions](#)

A “half – reaction” sounds like what it is, a chemical equation showing only the Oxidation OR the Reduction half of the reaction. This is done to focus on one half of the reaction at a time, something chemists do when doing such things as designing batteries or other electrochemical processes, such as refining metals out of ores, or electroplating silver onto a cheaper metal.

Given a half – reaction, you can quickly decide if it is showing oxidation or reduction by just examining the change in oxidation state that occurs in going from the reactant to the product.

For example, given this reaction:



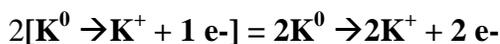
We know this shows reduction, since the charge on Ca went **DOWN** from +2 to 0. Electrons were gained by the Ca^{2+} .



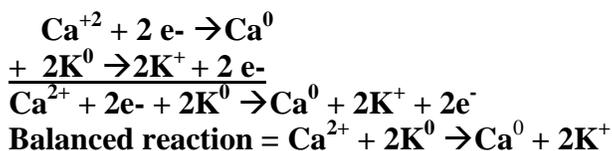
We know this shows oxidation, since the charge on K went **UP** from 0 to +1. An electron was lost by the K.

In this reaction, the Cl^{1-} is not undergoing any chemical change, as can be seen by the fact that it is Cl^{1-} both as a reactant and as a product. A chemist would refer to this as a “spectator ion” for this reason. It is “watching” the other elements do chemical change, but not involved in that change. Spectator ions are often not shown when writing REDOX reactions, showing only the REDOX “skeleton” instead.

Now in order to balance the equations, the number of electrons lost must equal the number of electrons gained. So, the oxidation reaction is multiplied by 2.



This oxidation is combined with the reduction reaction



Notice that the total charge on the reactants side = the total charge on the products side.

This is a good way to check your work:

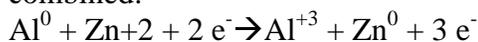
On the REACTANTS side... +2 from Ca^{+2} plus zero on the 2 K atoms = +2

On the PRODUCTS side... zero on the Ca atom plus 2 (1+) from the 2 K^{1+} ions = +2

Let's consider these two half reactions:



Notice Al loses 3 electrons per atom while Zn^{+2} gains 2 electrons per ion. If these two half reactions are combined:



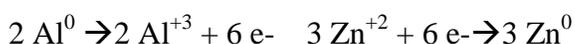
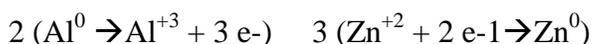
At first glance this may appear balanced, and it IS for MASS, but **not** for CHARGE.

We have to get the total charge to be equal on both sides of the arrow. We will use the same thinking we used when balancing for mass. I think this is most easily accomplished by getting the same number of electrons on both the reactant and product side, so that they will cancel each other out. (The number of electrons lost by Al should equal those gained by Zn^{2+} .)

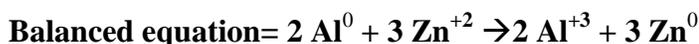
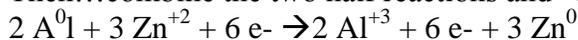
You first have to start with the half reactions...



The lowest common factor between 2 and 3 is **6**. Multiply both half-reactions so that you will get 6 electrons.



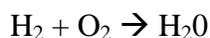
Then...combine the two half reactions and "cancel out" your electrons!



- Half cells are produced when a metal is placed into a salt solution creating source of metal ions to be reduced and metal atoms to be oxidized
- Conservation of mass
- Conservation of charge
- Always ADD ELECTRONS to the side of reaction that has a HIGHER TOTAL CHARGE (remember: electrons are NEGATIVE!)

Special note: In the case of diatomic elements you must balance using coefficients in the half reaction first to make sure the atoms are balanced. This is the only time we do this!

Example: Balance the following equation



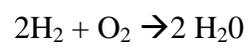
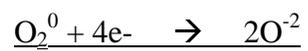
Not balanced:



Balanced for mass in the half reaction, then check to see if the electrons lost = electrons gained.



Then you would multiply the appropriate half reaction to balance electrons lost= electrons gained. Now balance the overall reaction for mass



Lesson 4: Spontaneous Reactions & Reference Table J

Objective:

- Use Table J identify which species is oxidized and reduced and determine if a reaction is spontaneous

For metals, you learned early in the year that the further to the left and the further down a metal is on the Periodic Table, the more reactive it is.

Increasing reactivity, Decreasing ionization energy



Na	Mg	Al		
K	Ca	Ga		
Rb	Sr	In	Sn	
Cs	Ba	Tl	Pb	Bi
Fr				

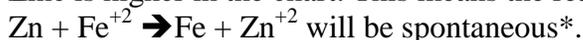
Do you remember why this is?

Consider Francium (Fr). Recall that this element is the largest, keeping its one valence electron in the 7th energy level. Since the valence electron is far out from the nucleus, it is not attracted to the nucleus very much at all. As a result, the valence electron is relatively easily removed from the atom. Remember that metal atoms have LOW IONIZATION ENERGY VALUES (Check Table S – the metal Fr has the lowest value of all!). Now check out Table J. You will notice that the most reactive metals are at the top, and the less reactive ones are towards the bottom. “**More Reactive**” for a **metal** means more likely to be **oxidized**, or to **lose electrons**.

If a two metal system is set up, the higher up metal in Table J will lose electrons to the lower metal’s ion. The reaction would be described as “**spontaneous**,” or in other words, “will happen as soon as the two substances come in contact.”

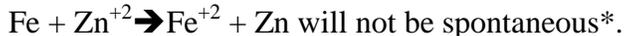
For example: **The two metal system of Zinc and Iron:**

Zinc is higher in the chart. This means the reaction of:



(*Zn is being oxidized & Fe+2 is reduced!)

But the reaction of:



(*Zn⁺² is being reduced & Fe is oxidized!)

IMPORTANT: The easier it is to oxidize a metal, the more difficult it is to reduce its cation.

If Zn is more reactive than Fe, then Fe²⁺ will be more reactive than Zn²⁺

For the non-metals, higher up in the Periodic Table means more reactive. In this case, more reactive means more likely to gain an electron. Remember, F is the most electronegative element (most likely to attract electrons in a bond; highest “electron grabbing ability”).

So, the reaction of F₂ with I⁻ will be spontaneous, but the reaction of I₂ with F⁻ will not occur.

**Table J
Activity Series****

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

**Activity Series based on hydrogen standard

Note: H₂ is *not* a metal

General Rule: elements **HIGHER** on Table J are **MORE** reactive than the elements below them

Spontaneous rxn = rxn occurs w/out adding energy to system

- If the “single” element is more active than the “combined” element, the reaction will be spontaneous.

Non-spontaneous rxn = rxn will not occur unless energy is added to system

- If the “single” element is less active than the “combined” element, the reaction will NOT be spontaneous.

Lesson 5: Electrochemical cells

Objective:

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

LEARN THE FOLLOWING TERMS!!

An electrochemical cell involves a chemical reaction where the flow of electrons is controlled by separating the two halves of a redox reaction, and allowing them contact only through conducting material.

ANODE: Terminal where **Oxidation** occurs (loss of electrons) .

CATHODE: Terminal where **Reduction** takes place (gain of electrons)

SALT BRIDGE: Connects two containers in voltaic cell; provides pathway for flow of ions between the containers

CURRENT FLOW:From anode (losing electrons) to cathode (gaining electrons) A →C

Some Memory Joggers:

1. “LEO goes GER”

Lose electrons-oxidation
Gain electrons-reduction

“OIL RIG”

oxidation is loss
reduction is gain

2. “BIG RED CAT”

The **CAT**hode is the site of **RED**uction and its mass gets **BIG**ger as the cell operates. It is the positively-charged terminal, or “electrode”, of the battery. All of this is the opposite for the anode. The anode is the negative terminal

3. “AN OX” is sometimes used to remember that the **AN**ode is the site of **OX**idation.

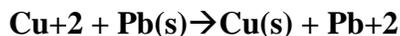
There are **two types** of electrochemical cells:

1) **VOLTAIC CELL:**

- A spontaneous chemical reaction takes place.
- This means, when looking at Table J, the metal which gets oxidized is above the metal which is reduced.
- The cell converts energy from chemical (stored) to electrical.
- An example is a battery.

An electrochemical cell which produces electricity as a result of a *spontaneous* chemical reaction is called a *voltaic* (or galvanic) cell. The chemical reaction is always of the oxidation-reduction type. Batteries and fuel cells are examples of voltaic cells. These devices produce energy for outside use, therefore they are exothermic

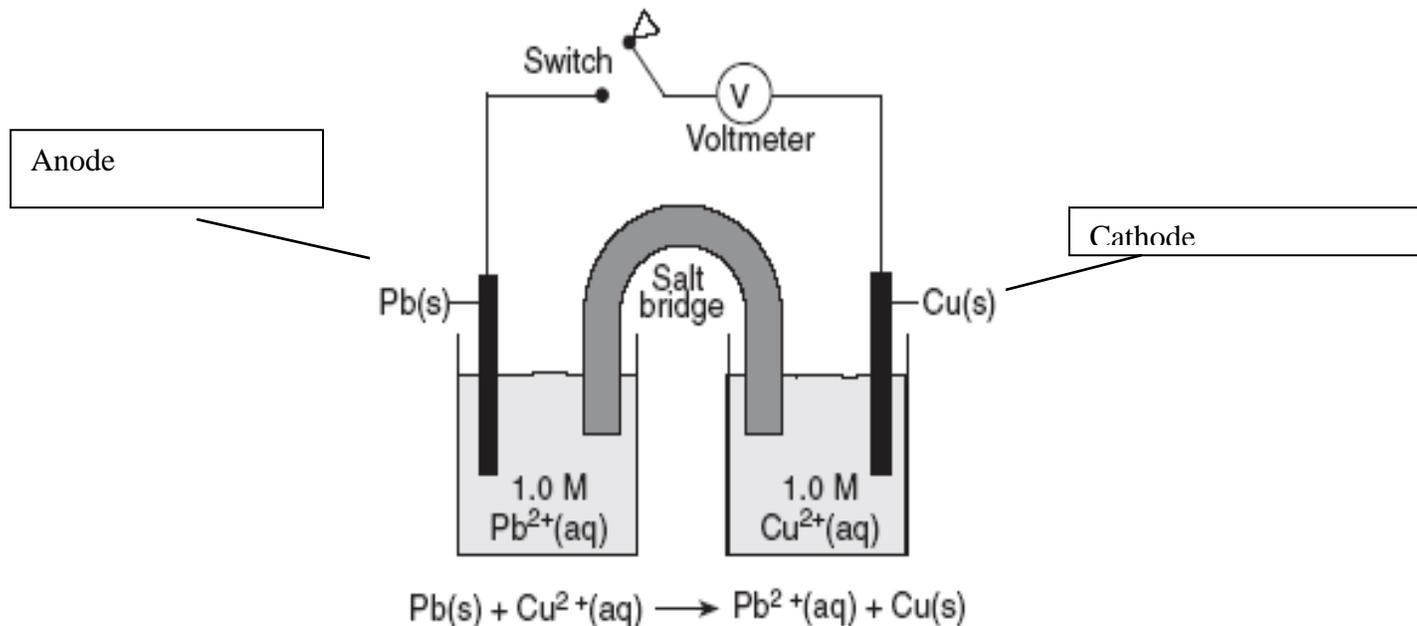
devices. An electrochemical cell is made up of two *half-cells*. Reduction occurs in one half-cell and oxidation in the other. Thus, if the electrochemical cell has the overall reaction,



The reduction half-cell reaction would be: $\text{Cu}^{2+} + 2 \text{e}^- \rightarrow \text{Cu(s)}$
and the oxidation half-cell reaction would be: $\text{Pb(s)} \rightarrow \text{Pb}^{2+} + 2 \text{e}^-$

Electrons are *gained* in a *reduction* half-reaction (GER) and *lost* in an *oxidation* half-reaction (LEO). To form a working voltaic cell, each half-cell must contain a conducting rod known as an *electrode*. The electrode at which *reduction* occurs is called the *cathode*, and the electrode at which *oxidation* occurs is named the *anode*. In our example above, a strip of Cu(s) metal would be the cathode, while a Pb(s) strip would be the anode. If a half-cell reaction does not contain a metal (such as $\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$), an inert electrode is used. Common examples are graphite electrodes and Pt(s) electrodes. The two electrodes are connected to one another by a wire, and are also connected via solution. For example, our cell described by equations above could be formed as illustrated below.

Copy this diagram into your notes!!!!



2) ELECTROLYTIC CELL:

- Electrical energy is converted to chemical energy. Requires an electric current to force a chemical reaction.
- This is how we plate silver or copper onto other metals.
- This is how we recharge a battery.

Both cells have electrodes (terminals) where the chemical reaction....oxidation and reduction.....is taking place.

An electrolytic cell uses an electric current to force a chemical reaction. In this case, electrical energy is converted to chemical energy. The process is endothermic; it is forced to happen by use of an outside energy source like electricity from a battery or a wall socket. An electrolytic cell is used to plate one metal on another. Silver-plated utensils are actually nickel that has been plated with silver. Another application of this idea is recharging your cell phone battery. When you plug it into the wall socket, the electricity is forcing the electrons in a non-spontaneous direction to “rebuild” the chemical reaction that runs the battery. The anode or more active metal is being forced to accept electrons instead of giving them away.

The following is a picture of a basic electrolytic

Copy this diagram into your notes!

