# <u>**Oxidation & Reduction**</u>:

<u>**Oxidation**</u> is the loss of electrons by an atom or group of atoms. When an atom oxidizes, it becomes more positive. Oxidation does not necessarily have to involve oxygen. It is true that when a metal reacts with oxygen, it oxidizes. An example of oxidation is when iron rusts. Iron loses electrons and goes from the neutral state(charge = **0**) to a charge of **+3** as illustrated by the following reaction: $4Fe + 3O_2 \rightarrow 2Fe_2O_3$. When an atom or group of atoms is oxidized, it is referred to as a <u>**reducing agent**</u>. In the above reaction, iron is the reducing agent.

<u>**Reduction**</u> is the gain of electrons by an atom or group of atoms. When an atom is reduced, it becomes more negative. *In mathematics, when a number becomes more negative, it decreases; hence, the number is reduced*. In the reaction $4Fe + 3O_2 \rightarrow 2Fe_2O_3$, oxygen is reduced because it gains electrons. When an atom or group of atoms is reduced, it is referred to as an <u>**oxidizing agent**</u>. In the above reaction, oxygen is the oxidizing agent.

Assigning oxidation numbers to the elements in a reaction is the way to determine if a chemical reaction is a redox reaction. Notice that all of the alkali earth metals always have oxidation state **+1**. All alkaline-earth metals always have oxidation state **+2**. With a few exceptions, all halogens have oxidation state **–1**. Oxygen always has oxidation state **–2** as well as most of the chalcogens. Hydrogen has oxidation state **+1**. With some exceptions, the lanthanides, actinides, and most of the transition metals have oxidation states **+3**. However, zinc and copper have oxidation state **+2**. Boron and aluminum have oxidation number **+3** with no exceptions. Silver has oxidation state **+1** with no exceptions. Iron can have oxidation state **+2**(*ferrous ion*) or **+3**(*ferric ion*). If you are in doubt as to what oxidation number an element in a compound has, figure out the oxidation numbers of the other elements in the compound. <u>**Remember that the oxidation numbers in a compound must add to zero**</u>.

Not all chemical reactions are oxidation-reduction(***Redox***) reactions. In some reactions, the atoms are displaced or replaced. The way that you can tell if a reaction is a redox reaction is if the oxidation numbers or charges on some of the atoms change. The below reactions will illustrate both redox and non-redox reactions:

1) $HCl + NaOH \rightarrow H_2O + NaCl$ is not a redox reaction. The oxidation number on sodium(Na) is **+1** on both sides; the oxidation on chlorine(Cl) is **–1** on both sides; and the oxidation numbers of hydrogen(H) is +1 and oxygen(O) is **–2** on both sides. This reaction is a double displacement reaction.

2) $NF_3 + AlCl_3 \rightarrow N_2 + Cl_2 + AlF_3$ is a redox reaction. First, we assign oxidation numbers to the elements as follows: $\overset{+3}{N}\overset{-3}{F_3} + \overset{+3}{Al}\overset{-3}{Cl_3} \rightarrow \overset{0}{N_2} + \overset{0}{Cl_2} + \overset{+3}{Al}\overset{-3}{F_3}$. To see this, the above rule says that the halogens have oxidation number **–1**. Using this, we can deduce that aluminum(Al) and nitrogen(N) have oxidation number **+3** on the left side of the equation. However, nitrogen and chlorine have oxidation number **0** on the right side of the equation(*Any other number added to itself will not yield* ***0*** *for* $N_2$*,* $O_2$*, and* $H_2$.). Hence, oxidation and reduction have taken place in the reaction. The above reaction must be balanced. We first balance the elements that have

been oxidized and reduced as follows: $2\overset{+3\,-3}{NF_3} + 2\overset{+3\,-3}{AlCl_3} \rightarrow \overset{0}{N_2} + 3\overset{0}{Cl_2} + AlF_3$. Then we label the loss and gain of electrons in the reaction as follows:

Gain of 6 e-'s

$$2\overset{+3\,-3}{NF_3} + 2\overset{+3\,-3}{AlCl_3} \rightarrow \overset{0}{N_2} + 3\overset{0}{Cl_2} + AlF_3$$

Loss of 6 e-'s

We find the lowest number that both 6 and 6 will divide into evenly = LCM least common multiple = 6. We multiply each number to get the LCM. Those numbers multiply the coefficients of the compounds that contain the oxidized and reduced elements. The above process is the same as finding the lowest denominator to add fractions as shown below: $\frac{1}{2} + \frac{1}{3} = ?$, LCM $= 6$. Next, we balance any additional compounds in the reaction to get the below result: $2NF_3 + 2AlCl_3 \rightarrow N_2 + 3Cl_2 + 2AlF_3$.

Using the above techniques, try balancing the below equations:

1) $Zn + HNO_3 \rightarrow Zn(NO_3)_2 + NH_4NO_3 + H_2O$

2) $NaI + H_2SO_4 \rightarrow H_2S + I_2 + Na_2SO_4 + H_2O$

3) $I_2 + HClO + H_2O \rightarrow HIO_3 + HCl$