**Electrochemistry**

The interconnection between electricity and chemical reactions is **electrochemistry**. Specifically electrochemistry deals with the transfer of electrons from one substance to another. These reactions can produce electricity in a spontaneous and non-spontaneous manner. In this unit you will learn about two important chemical reactions involving the transfer of electrons; oxidation and reduction. You will also learn about how the energy is harvested from oxidation and reduction reactions using special cells called galvanic and electrolytic cells. Practical application of electrochemistry will also be learnt.

Oxidation and Reduction Reactions (REDOX)

Oxidation and reduction reactions involve the transfer of electrons.

Consider:

Overall equation

 Zn(s) + CuSO4(aq) 🡪 Cu(s) +ZnSO4(aq)

Total Ionic equation

 Zn(s) + Cu2+(aq) +SO42-(aq) 🡪 Cu(s) + Zn2+(aq) + SO42-(aq)

Net Ionic equation

 Zn(s) + Cu2+(aq) 🡪 Cu(s) + Zn2+(aq)

Notice what happens to the reactants in this equation.

The zinc atoms lose electrons to form zinc ions. (Zn(s) 🡪 Zn2+(aq))

The copper ions gain electrons to form copper atoms. (Cu2+(aq) 🡪 Cu(s))

An oxidation reaction describes a loss of electrons, while a reduction reaction occurs when there is a gain of electrons. These reactions always occur in tandem and hence are referred to a REDOX.

 See page 655 to 656 for an example problem.

One easy way of remembering REDOX reactions is the expression

 “**LEO the Lion says GER”**

 **Lose Electrons Oxidation, Gain electrons Reduction**

REDOX Reactions of Non-metals

With non-metals there is not a transfer of electrons but a sharing of electrons. An apparent charge is assigned to each atom, called an **oxidation number**.

The more electronegative atom is considered negative and the less electronegative atom is considered positive.

 Example:

 CO2 Oxygen is more electronegative and will form a 2- ion with a metal

 Therefore, the oxidation number for oxygen is 2-.

 To find the oxidation number for carbon

 Two oxygen atoms = 2(2-) = total charge of 4-

 Thus the carbon atom must have a charge 4+

Rules for Assigning Oxidation Numbers

|  |  |
| --- | --- |
| Rule | Examples |
| 1. the oxidation number of an atom in an uncombined state is always 0. | K, O2, N, Li etc all have an oxidation number of 0 |
| 2. the oxidation number of a number of a simple ion. | The oxidation number of Mg in Mg2+ is 2+ |
| 3. the oxidation number of hydrogen in most compounds is 1+. Exception: metal hydrides. | The oxidation number of H in H2O, HCl, HNO3 is 1+The oxidation number of H in sodium hydride, NaH is 1- |
| 4. the oxidation number of oxygen in most compounds is 2-. | The oxidation number of O in H2O, MgO, and HNO3 is 2-The oxidation number of O in hydrogen peroxide, H2O2 is 1- |
| 5. the oxidation number of Group 1 element ions to 1+. The oxidation number of Group 2 element ions is 2+. | The oxidation number of K in K+ is 1+The oxidation number of Mg in Mg2+ is 2+ |
| 6. the sum of oxidation number in a compound must equal 0. | In H2O2(1+) + (2-) = 0 |
| 7. the sum of oxidation numbers in a polyatomic ion must equal the charge of the ion. | In hydroxide, OH-(2-) + (1+) = 1- |

 See page 658-659 for an example problem.

Example: Lithium nitrate, LiNO3

 Known oxidation numbers: Li = 1+, N = ?, O3 = 3(2-)

 Solve for oxidation number of N

 (1+) + x + 3(2-) = 0

 x – 5 = 0

 x = 5+

Note: REDOX reactions have only occurred if there is an oxidation number that decreased and an oxidation number that has increased.

Balancing REDOX Reactions

Method 1: Oxidation Number Method

The increase in oxidation number for an atom/ion must equal the decrease in the oxidation number of another atom/ion.

* Determine the atom/ion that is oxidized
* Determine the atom/ion that is reduced
* Determine the ratio of oxidized to reduced atoms that will produce a net oxidation to reduction of zero
* Add the ratio as coefficients the formulas in the equation which contain the elements whose oxidation number is changing
* Balance the rest of the equation
* Use water molecules, hydrogen ions, or hydroxide ions if necessary

See page 666-667 for an example problem

Method 2: Half-Reaction Method

* Write the unbalanced half-reactions
* Balance any atoms other than oxygen and hydrogen
* Balance any oxygen atoms by adding water molecules
* Balance any hydrogen atoms by adding hydrogen ions
* Balance electron charges
* Add half cell reactions
* For Basic solutions add hydroxide ions to both sides to equal the number of hydrogen ions present
* Simplify the half reaction by combining hydrogen ion and hydroxide to form water
* Cancel out water molecule on both sides
* Balance the charges by adding electrons

See page 669-670 for an example problem

Predicting REDOX Reactions

Oxidizng Agent – a substance that gains or removes electrons from another substance in a REDOX reaction.

Reducing Agent – a substance that loses or gives up electrons to another substance in a REDOX reaction

In single displacement reactions, the positively charged metal ion will react and gain electrons (oxidizing agent) to become a metal atom. The metal will react to lose electrons (reducing agent) and become a positively charged metal ion.

By observing single displacement reactions evidence shows that some reactions will spontaneously occur but other do not react at all.

We can rank the ability of metal ions to react with the metal atoms. We can also rank the ability of metal atoms to react with the metal ions. This is the Activity Series of Metals.

The table is organized from Most Reactive (easily oxidized) to Least Reactive (not easily oxidized).

 See page 678, 682-683 for an example problems