**Chem 1045** **Chemistry & Chemical Reactivity**
**Lecture Notes** **Kotz/Treichel/Townsend, 8th Ed**

**These Notes are to SUPPLEMENT the Text, They do NOT Replace reading the Text Material. Additional material that is in the Text will be on your tests! To get the most information, READ THE CHAPTER prior to the Lecture, bring in these lecture notes and make comments on these notes. These notes alone are NOT enough to pass any test! The author is not responsible for typos in these notes.**

## Chapter 3, Chemical Reactions

**Start Studying these tables early**

**Figure 3.10 Guidelines to predict the solubility of ionic compounds**
**Table 3.1 Common Acids and Basis**
**Table 3.2 Gas-Forming Reactions**
**Table 3.3 Common Oxidizing & Reducing Agents**

Metal Sulfides are black and metal sulfides come from the center of the earth. Sulfides are insoluble in water so they form a black mass in the deep ocean floor cracks.

Chemical Reactions are the heart of Chemistry. This chapter is an introduction to symbols and chemical reactions.

### 3.1 Intro to Chemical Equations

In the late 1770's Oxygen was discovered by Joseph Priestley coming from heating mercury (II) oxide

$$2\ HgO_{(s)} \xrightarrow{\Delta} 2\ Hg_{(l)} + O2_{(g)}$$

They also determined that Oxygen also comes from water and burning involved a reaction with Oxygen. The heat generated by a guinea pig exhaling Carbon Dioxide ($CO_2$) is the same amount as produced by burning Carbon to Carbon Dioxide. Respiration is slow combustion

**A Balanced Chemical Equation** $P_{4\,(s)} + 6\ Cl_{2\,(g)} \rightarrow 4\ PCl_{3\,(l)}$

In a balanced Chemical Equation you have the same number of individual elements on the left and right side of the reaction arrow.

**Reactants** are on the left of the arrow **Products** are on the right of the arrow

Physical States are represented by: **(s)** = Solid, **(g)** = Gas, **(l)** = Liquid, **(aq)** = aqueous

A Solid can sometimes be shown as ↓ and a gas as ↑

A substance dissolved in water is an **Aqueous Solution (aq)**

**Law of Conservation of Matter** = matter can neither be created or destroyed. Atoms are conserved in Chemical Reactions. The same elements and number of elements on the left side (the reactants) equals those on the right side (the products).

If the total weight of 100.0 g of reactants, there will be a total weight of 100.0 g of products

For the reaction above, there are 4 atoms of P on the left and right.
There are 6 * 2 or 12 atoms of Cl on the left and 4 * 3 atoms or 12 on the right.
The number 6 before the Cl2 and 4 before the PCl3 are called **Stoichiometric Coefficients.**

**Stoichiometric Coefficients** are the coefficients used to balance an equation

$P_{4\,(s)} + 6\, Cl_{2\,(g)} \rightarrow 4\, NaCl_{3\,(l)}$ THIS CANNONT HAPPEN IN CHEMISTRY, Why?

$2\, Fe_{(s)} + 3\, Cl_2 \rightarrow 2\, FeCl_{3\,(s)}$ 2, 3 and 2 are the Stoichiometric Coefficients

$2\, Al_{(s)} + 3\, Br_{2\,(l)} \rightarrow Al_2Br_{6\,(s)}$ What are the Stoichiometric Coefficients?

If we start with 8000 atoms of Al, how many molecules of $Br_2$ are required to consume all of the Al?

8000 atoms of Al * 3 $Br_2$ / 2 Al = 8000 * 3/2 atoms (really molecules) of $Br_2$

**Balancing Equations**

YOU must have the same number of atoms of each element on each side of the equation.
You CANNOT change the subscripts as this changes the identity of the substance

Changing $CO_2$ to CO changes from Carbon Dioxide to Carbon Monoxide

Chemical equations are balanced using stoichiometric coefficients.

Write the balanced equation:

$C_3H_{8\,(g)} + O_{2\,(g)} \rightarrow CO_{2\,(g)} + H_2O_{(g)}$ UNBALANCED EQUATION

Balance the Carbons, then the Hydrogen, then the Oxygen, verify all is correct

$C_3H_{8\,(g)} + O_{2\,(g)} \rightarrow 3\, CO_{2\,(g)} + H_2O_{(g)}$ Carbon Balanced

$C_3H_{8\,(g)} + O_{2\,(g)} \rightarrow 3\, CO_{2\,(g)} + 4\, H_2O_{(g)}$ Hydrogen Balanced

$C_3H_{8\,(g)} + 5\, O_{2\,(g)} \rightarrow 3\, CO_{2\,(g)} + 4\, H_2O_{(g)}$ BALANCED EQUATION

## Use The Ping-Pong Method of Balancing a Chemical Equation

$FeCl_3 + AgNO_3 \rightarrow Fe(NO_3)_3 + AgCl$

1. Start on the left side – pic one cation. Pick one that is large or strange – take **Fe**
2. There is one **Fe** on the left side, there is one **Fe** on the right side, Fe is balanced
3. Look at the **Fe** on the right side. It is attached to Nitrate – **$NO_3$.** Balance this next.
4. There are 3 **$NO_3$** on the right side, but only 1 on the left side
5. Make it so there are 3 **$NO_3$** on the left side:

   $FeCl_3 + \underline{\mathbf{3}}\, Ag\mathbf{NO_3} \rightarrow Fe\mathbf{(NO_3)_3} + AgCl$

6. There are now 3 **$NO_3$** on both the left and right side
7. Attached to the **$NO_3$** on the left side are 3 **Ag.**
8. There is only 1 **Ag** on the right side. Make it 3!

   $FeCl_3 + \underline{\mathbf{3}}\, \mathbf{Ag}NO_3 \rightarrow Fe(NO_3)_3 + \underline{\mathbf{3}}\, \mathbf{Ag}Cl$

9. Connected to the **Ag** on the right is **Cl**. There are 3 **Cl** on the right
10. There are also 3 **Cl** on the left
11. THE EQUATION IS NOW BALANCED! But you should verify it by counting the atoms on each side of the equation!

Write a Balanced Chemical Equation for the following:

$Fe^{2+} + SO_4^{2-} \rightarrow Fe_2SO_4$

$2\, Bi^{3+} + 3\, SO_4^{2-} \rightarrow Be_2(SO_4)_3$

$H_2S + SO_4^{2-} \rightarrow H_2SO_4 + S^{2-}$

$Ca^{2+} + SO_4^{2-} \rightarrow Ca_2SO_4$

Metals and nonmetals react with Oxygen to yield **Oxides**: (Balance the following)

$Fe_{(s)} + O_{2(g)} \rightarrow Fe_2O_{3(s)}$ $\qquad$ $S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$

$P_{4(s)} + O_{2(g)}$ -> $P_4O_{10(s)}$

Burning a hydrocarbon (contains C and H) yields $CO_2$ and $H_2O$ and energy (**Δ** = heat)

**Example 3.1** Write the balanced equation for the combustion of Ammonia Gas ( $NH_3$ ) to give water vapour and Nitrogen Monoxide gas (You should be able to do this on your own by now).

**Combustion is** burning with oxygen and evolves heat. Products are all Oxides

Octane in Gas $\qquad C_8H_{18} + O_2 \rightarrow CO_2 + H_2O_{)} + \Delta$

When a hydrocarbon (contains only H & C) is combusted he products are always $CO_2$, HOH and energy (**Δ** = heat)

Balance: $\quad NH_{3\,(g)} + O_{2\,(g)} \rightarrow NO_{(g)} + H_2O_{(g)}$

$C_4H_{10\,(g)} + O_{2\,(g)} \rightarrow CO_{2\,(g)} + H_2O_{(g)}$

$Pb(C_2H_5)_{4\,(g)} + O_{2\,(g)} \rightarrow PbO_{(s)} + H_2O_{(g)} + CO_{2\,(g)}$

**Chemical Equilibrium:** Chemical Reactions are [ may be as you will learn in 1046 ] reversible.

**Stalagmites** are Calcium Carbonate:

$CaCO_{3(s)} + CO_{2\,(aq)} + H_2O_{(l)} \rightarrow Ca(HCO_3)_{2\,(aq)}$ $\quad$ Water, with dissolved CO2 from the air, goes through the rock dissolving $CaCO_{3(s)}$

$Ca(HCO_3)_{2\,(aq)} \rightarrow CaCO_{3(s)} + CO_{2(g)} + H_2O_{(l)}$ $\quad$ Calcium Bicarbonate gives up water and $CO_2$ and Forms Stalagmite – reverse of above reaction

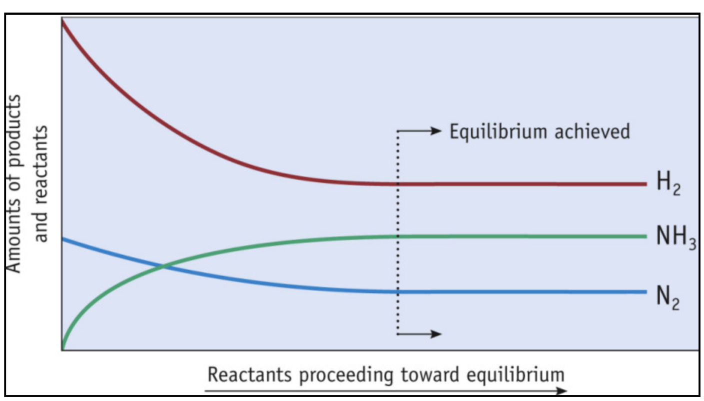
**THUS** $Ca(HCO_3)_{2\,(aq)} \leftarrow\rightarrow CaCO_{3(s)} + CO_{2(g)} + H_2O_{(l)}$ $\quad$ Is a reversible reaction

Adding CO2 forces the reverse of this reaction

**Xcs $CO_2$** $+ H_2O \rightarrow \quad H_2CO_3 \quad \leftarrow\rightarrow H^+ + HCO_3^-$

Fertilizer is made from Ammonia. Ammonia is made from Hydrogen and Nitrogen:

$H_{2(g)} + N_{2(g)} \leftarrow\rightarrow NH_{3(g)}$ $\qquad$ You Balance this equation




At time = infinity, the system has reached Chemical Equilibrium. No further **Macroscopic change** is observed. Also called **Dynamic Equilibrium** = the rate of the forward reaction equals the rate of the reverse reaction. Chemical Reactions always proceed spontaneously toward equilibrium

**Product favored reactions**: reactants are completely or largely converted to products when at equilibrium. Combustion is an example of Product favored reactions.

**Reactant favored reaction**: Only a small amount of products are formed at equilibrium
The ionization of Acetic Acid in water solution only proceeds to a small percent; this is why Acetic Acid is a weak acid:

$$\underset{\text{Acetic Acid}}{CH_3COOH_{(aq)}} + \underset{\text{Water}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{Acetate Ion}}{CH_3CO_2^-{}_{(aq)}} + \underset{\text{Hydronium Ion}}{H_3O^+{}_{(aq)}}$$

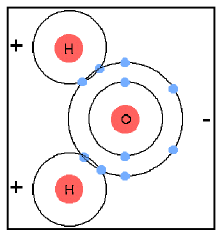
**3.4 Aqueous Solutions.** Most General Chem reactions are carried out in water solutions.

**Solution:** a homogeneous mixture of two or more elements

**Solvent:** the medium in which the solute is dissolved in, usually the item in the largest amount

**Solute:** the item in the smaller amount

**Aqueous Solutions:** solutions in which water is the solvent. Water is good a dissolving ionic compounds because water is polar, has a positive and a negative end. Ionic compounds are usually polar – like dissolves like.




**Electrolysis:** We have two charged electrodes, one + and one -. **Positively charged ions (cations)** are attracted to the negative electrode. **Negatively charged ions (anions)** are attracted to the positively charged electrode and electricity flows!

**Electrodes:** conductor of electricity

**Electrolytes:** compounds whose aqueous solutions conduct electricity.

**Strong Electrolytes:** Substances whose solutions are good electrical conductors as they are completely ionized.

$$NaCl_{(s)} \rightarrow H_2O \text{ / NaCl Solution} \rightarrow Na^+{}_{(aq)} + Cl^-{}_{(aq)}$$

Dissolving 1 mole of NaCl in water gives one mole of $Na^+$ and 1 mole of $Cl^-$. It is 100% dissociated

Dissolving 1 mole of BaCl2 in water gives one mole of $Ba^{2+}$ and 2 moles of $Cl^-$

$$BaCl_{2\,(s)} \rightarrow H2O \text{ / } BaCl_2 \text{ Solution} \rightarrow Ba^{2+}{}_{(aq)} + \mathbf{2\ Cl^-{}_{(aq)}}$$

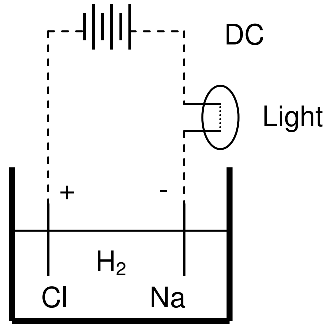
**Weak Electrolytes:** Compounds dissolved in water and only a small fraction of the molecules form ions, such as Acetic Acid

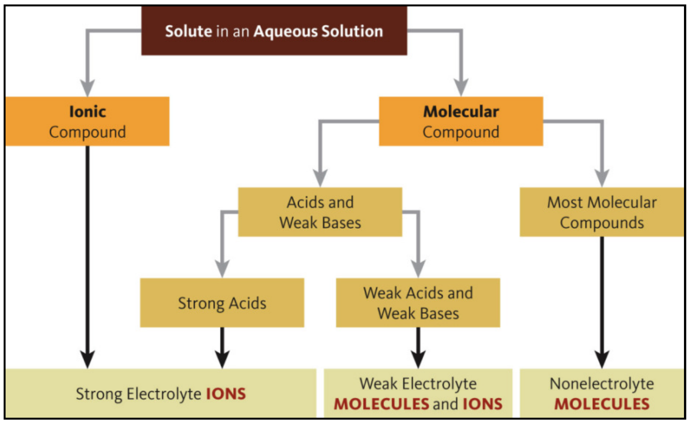
$$\underset{\text{Acetic Acid}}{CH_3COOH_{(aq)}} + \underset{\text{Water}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{Acetate Ion}}{CH_3CO_2^-{}_{(aq)}} + \underset{\text{Hydronium Ion}}{H_3O^+{}_{(aq)}}$$

**Non-Electrolytes:** Compounds whose aqueous solutions do not conduct electricity:

Ethanol $CH_3\text{-}CH_2OH_{(l)} \rightarrow CH_3\text{-}CH_2OH_{(aq)}$

**Experiment to show the conduction of electricity**. Put 2 electrodes into water and attach to a battery and to a light bulb. Bulb will light if electricity is flowing.




Note: even though acids (HCl) and bases (NaOH) may seem like an ionic compound, this book classifies them as Molecular Compounds.

Solubility of Ionic Compounds in Water **MEMORIZE THIS CHART. See also end of these notes for a different approach to solubility.**

| SOLUBLE COMPOUNDS | EXCEPTIONS |
|---|---|
| Almost all salts of $Na^+$, $K^+$, $NH_4^+$ | |
| Salts of nitrate, $NO_3^-$<br>chlorate, $ClO_3^-$<br>perchlorate, $ClO_4^-$<br>acetate, $CH_3CO_2^-$ | |
| Almost all salts of $Cl^-$, $Br^-$, $I^-$ | Halides of $Ag^+$, $Hg_2^{2+}$, $Pb^{2+}$ |
| Salts containing $F^-$ | Fluorides of $Mg^{2+}$, $Ca^{2+}$, $Sr^{2+}$, $Ba^{2+}$, $Pb^{2+}$ |
| Salts of sulfate, $SO_4^{2-}$ | Sulfates of $Ca^{2+}$, $Sr^{2+}$, $Ba^{2+}$, $Pb^{2+}$, $Ag^+$ |
| **INSOLUBLE COMPOUNDS** | **EXCEPTIONS** |
| Most salts of carbonate, $CO_3^{2-}$<br>phosphate, $PO_4^{3-}$<br>oxalate, $C_2O_4^{2-}$<br>chromate, $CrO_4^{2-}$<br>sulfide, $S^{2-}$ | Salts of $NH_4^+$ and the alkali metal cations |
| Most metal hydroxides and oxides | Alkali metal hydroxides and $Ba(OH)_2$ and $Sr(OH)_2$ |

**Soluble or Insoluble**

**Soluble** are materials that are soluble beyond a certain extent
**Insoluble** are materials that do no dissolve to that extent

| | | | | |
|---|---|---|---|---|
| Predict the solubility of: | $KCl$ | $MgCO_3$ | $Fe_2O_3$ | $Cu(NO_3)_2$ |
| | $LiNO_3$ | $CaCl_2$ | $CuO$ | $NaCH_3CO_2$ (sodium acetate) |
| | $Ba(NO_3)_2$ | $CuS$ | $Fe_3(PO_4)_2$ | $Mg(OH)_2$ |

**There are Four Categories of Reactions in Aqueous Solutions**

### Precipitation, Acid-Base, Gas forming, REDOX

## 3.5 Precipitation Reactions

Exchange Reactions – double displacement – the ions exchange partners

$$A^{+}B^{-} \quad + \quad C^{+}D^{-} \quad \longrightarrow \quad A^{+}D^{-} \quad + \quad C^{+}B^{-}$$

A **Precipitation reaction** produces a water insoluble solid product known as a **precipitate**

$AgNO_{3\,(aq)} + KCl_{\,(ag)} \rightarrow AgCl \downarrow_{(s)} + KNO_{3\,(aq)}$

**Predicting Outcome** of a precipitation reaction:

If the Reactants are insoluble – there will be no reaction
If the Products are insoluble – there will be a reaction and a precipitate.
If the Reactants and Products are soluble, there probably will not be a reaction, just a mixture of ions.

$Pb(NO_3)_{2\,(aq)} + K_2CrO_{4\,(aq)} \rightarrow PbCrO_4 \downarrow_{(s} + 2\ KNO_{3\,(aq)}$

$Pb^{2+(}{}_{aq)} + 2\ NO_3^-{}_{(aq)} \rightarrow PbCrO_4 \downarrow_{(s}$ Insoluble

$2\ K^+{}_{(aq)} + CrO_4^{2-}{}_{(aq)} \rightarrow 2\ K^+{}_{(aq)} + 2\ NO_3^-{}_{(aq)}$

**Students do these:**

$Pb(NO_3)_{2\,(aq)} + (NH_4)_2S_{\,(aq)} \rightarrow$ You figure it out

$FeCl_{2\,(aq)} + NaOH_{\,(aq)} \rightarrow$ You figure it out

**Example 3.3:** Is there a ppt? Write the balanced equation:

An aq solution of Potassium Chromate and Silver Nitrate?

Sodium Carbonate and Copper (II) Chloride

Potassium Carbonate and Sodium Nitrate

Nickel (II) Chloride and Potassium Hydroxide

**Ionic Equations**

$AgNO_{3\,(aq)} + KCl_{(ag)} \rightarrow AgCl \downarrow_{(s)} + KNO_{3\,(aq)}$

**Complete Ionic Equation** – break all **SOLUBLE** molecules down to their ions:

$Ag^{+}_{(aq)} + NO_3^{-}{}_{(aq)} + K^{+}_{(aq)} + Cl^{-}{}_{(ag)} \rightarrow AgCl \downarrow_{(s)} + K^{+}_{(aq)} + NO_3^{-}{}_{(aq)}$

Spectator Ion is the same ion on both sides of the equation

$Ag^{+}_{(aq)} + \mathbf{NO_3^{-}{}_{(aq)}} + \underline{\mathbf{K^{+}_{(aq)}}} + Cl^{-}{}_{(ag)} \rightarrow AgCl \downarrow_{(s)} + \underline{\mathbf{K^{+}_{(aq)}}} + \underline{\mathbf{NO_3^{-}{}_{(aq)}}}$

**Net Ionic Equation** – remove the Spectator Ions – remember there must be charge balance and element/polyatomic balance

$$Ag^{+}_{(aq)} + Cl^{-}{}_{(ag)} \rightarrow AgCl \downarrow_{(s)}$$

**Student Write** the Complete Ionic Equation and determine the Net Ionic Equation for:

The strong acid/base reaction of Hydrochloric Acid and Sodium Hydroxide

The reaction of Barium Chloride and Sodium Sulfate

The reaction of Calcium Chloride and Sodium Phosphate

The reaction of Silver Nitrate and Sodium Carbonate

## 3.6 Acid / Base

**Acids:** Produce $CO_2$ bubbles when added to a metal carbonate $CaCO_3$

React with metals to produce $H_2$ gas

Taste Sour (vinegar, citric acid) – Don't ever do a taste test for an acid!

Turn blue litmus to red

### Arrhenius Definition

**Acid** when dissolved in waer, increases the $H^+$ or Hydronium ion concentration

$HCl_{(g)} + H_2O_{(l)} \rightarrow \underline{\mathbf{H_3O^{+}{}_{(aq)}}} + Cl-_{(aq)}$

**Base** when dissolved in water, increases the OH- concentration

$NaOH_{(s)} + H_2O_{(l)} \rightarrow Na^{+} + H_2O_{(l)} + \underline{\mathbf{OH^{-}{}_{(aq)}}}$

**Reaction of an acid and a base products a salt and water**

$HCl_{(aq)} + NaOH_{(s)} \rightarrow NaCl + H\text{-}OH$

**Strong Acid:** Completely ionize in water, eg HCl

**Weak Acid:** Incompletely ionize in water, eg H2CO3 Carbonic Acid

**Strong Base**: Water soluble compounds that contain hydroxide: NaOH, KOH

**Weak Base**: Water soluble hydroxide that partially ionizes: NH4OH

## Common Acids and Bases: Yes, you need to memorize these

**Table 3.1** Common Acids and Bases

| Strong Acids (Strong Electrolytes)* | | Soluble Strong Bases | |
|---|---|---|---|
| HCl | Hydrochloric acid | LiOH | Lithium hydroxide |
| HBr | Hydrobromic acid | NaOH | Sodium hydroxide |
| HI | Hydroiodic acid | KOH | Potassium hydroxide |
| $HNO_3$ | Nitric acid | $Ba(OH)_2$ | Barium hydroxide |
| $HClO_4$ | Perchloric acid | $Sr(OH)_2$ | Strontium hydroxide |
| $H_2SO_4$ | Sulfuric acid | | |
| **Weak Acids (Weak Electrolytes)*** | | **Weak Base (Weak Electrolyte)*** | |
| HF | Hydrofluoric acid | $NH_3$ | Ammonia |
| $H_3PO_4$ | Phosphoric acid | | |
| $H_2CO_3$ | Carbonic acid | | |
| $CH_3CO_2H$ | Acetic acid | | |
| $H_2C_2O_4$ | Oxalic acid | | |
| $H_2C_4H_4O_6$ | Tartaric acid | | |
| $H_3C_6H_5O_7$ | Citric acid | | |
| $HC_9H_7O_4$ | Aspirin | | |

*The electrolytic behavior refers to aqueous solutions of these acids and bases.

**<u>Bronsted-Lowry Definition</u>** **Acid** is a proton donor
**Base** is a proton acceptor

**Acid Base reaction** involves the transfer of a proton from an acid to a base to form a new base and a new acid, the equilibrium favors the weaker acid and base:

**Strong Acid** – HCl Completely Ionized

$HCl_{(aq)} + H_2O_{(l)} \leftarrow \rightarrow$ $\mathbf{H_3O^+_{(aq)}} + Cl^-_{(aq)}$ HCl is a strong acid, this reaction goes 100%
Hydronium Ion

**Weak Acid** – Acetic Acid Partially Ionized

$CH_3COOH_{(aq)} + H_2O_{(l)} \leftarrow \rightarrow H_3O^+_{(aq)} + CH3COO^-_{(aq)}$ Reaction only goes abour 1%
Acetic Acid Acetate Ion

Note: $H_3O^+_{(aq)}$ is a stronger acid than $CH_3COOH$, so the reaction favors the weaker or this is Reactant Favored (to the left)

**Diprotic Acid:** can give up two $H^+$, eg: Sulfuric Acid

1. $H2SO_{4\,(aq)} + H_2O_{(l)} \leftarrow \rightarrow H_3O^+_{(aq)} + HSO_4^-{}_{(aq)}$ Reaction goes 100%
   Hydrogen Sulfate Ion
2. $HSO_4^-{}_{(aq)} + H_2O_{(l)} \leftarrow \rightarrow H_3O^+_{(aq)} + SO_4^{2-}{}_{(aq)}$ Reaction goes < 1%

**Weak Base:** reacts with water to produce OH-, but at less than 100%, e.g.Ammonia

$NH_{3\,(aq)} + H_2O_{(l)} \leftarrow \rightarrow NH_4^+{}_{(aq)} + OH^-_{(aq)}$

**Amphiprotic** – function as an acid or a base ( Water is amphoteric / amphiprotic)
**Is a Base** see Diprotic Acid acid above – water accepts a proton
**Is an Acid** see Weak Base above – water donates a proton

**Exampe 3.5** Discuss reacting cyanide with a proton, is it a Bronsted Acid or Base?
React phosphoric acid and water to form dihydrogen phosphate ion

**Reactions of Acid & Base:**

Acids and Bases react to form water and a salt.

What is the complete ionic equation? What is the net ionic equation?

$HCl_{(aq)} + NaOH_{(aq)} \rightarrow H2O_{(l)} + NaCl_{(aq)}$

**Sulfuric Acid** is produced from sulfur:

$S_{(s)} + O_{2\,(g)} \rightarrow SO_{2\,(g)}$ $2\,SO_{2\,(g)} + O_{2\,(g)} \rightarrow 2\,SO_{3\,(g)}$

$SO_{3\,(g)} + H_2O_{(l)} \rightarrow H_2SO_{4\,(aq)}$

Sulfuric Acid is a colorless syrupy liquid, den 1.84 g/ml

Less expensive to produce than other acids

Reacts with many organic compounds

Reacts with Lime (Calcium Oxide CaO) to produce CaSO4 (Calcium Sulfate) used in wall board

Used to produce fertilizer, pigments, explosive, pulp and paper, detergents and in storage batteries.

**Neutralization Reactions** are reactions between strong acids and strong bases which produce water and a salt:

| | | | |
|---|---|---|---|
| $HNO_{3\,(aq)}$ + | $NaOH_{(aq)} \rightarrow$ | $H_2O_{(l)}$ + | $NaNO_{3\,(aq)}$ |
| Acid | Base | Water | Salt |

$H^+_{(aq)} + NO3^-_{(aq)} + Na^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O\,(l) + Na^+_{(aq)} + NO3^-_{(aq)}$

**Remove the spectator ions and you get this Net Ionic Equation**

$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H2O\,(l)$

$CH_3COOH_{(aq)}$ + $NaOH_{(aq)} \rightarrow H_2O_{(l)} + CH_3COO^-Na^+$

| NaOH is a strong base

Acidic Acid (vinegar) is a weak acid,

**Oxides of Non Metals and Metals** have no H atoms, but react with water to produce $H3O^+$

**Acidic Oxides** are oxides that react with water to produce the Hydronium Ion

$CO_{2\,(g)} + H_2O_{(l)} \leftrightarrow H_2CO_{3\,(aq)}$ Carbonic Acid

$H_2CO_{3\,(aq)} + H_2O_{(l)} \leftrightarrow H_3O^+_{(aq)} + HCO_3^-{}_{(aq)}$ 1st Proton Ionization

$HCO_3^-{}_{(aq)} + H_2O_{(l)} \leftrightarrow H_3O^+_{(aq)} + CO_3^{2-}{}_{(aq)}$ 2nd Proton (Di Protic)

Rainwater contains dissolved CO2, thus is it slightly acidic

$2\,SO_{2\,(g)} + O_{2\,(g)} \rightarrow 2\,SO_{3\,(g)} + H_2O_{(l)} \rightarrow H_2SO_{4\,(aq)}$ Sulfuric Acid

**Basic Oxides** are oxides of metals that give basic aqueous solutions

$CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2\,(s)}$

## 3.7 Gas Forming Reactions – See Table 3.2

**Table 3.2 Gas-Forming Reactions**

| |
|---|
| **Metal carbonate or hydrogen carbonate + acid → metal salt + $CO_2(g)$ + $H_2O(\ell)$** |
| $Na_2CO_3(aq) + 2\ HCl(aq) \rightarrow 2\ NaCl(aq) + CO_2(g) + H_2O(\ell)$ |
| $NaHCO_3(aq) + HCl(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(\ell)$ |
| **Metal sulfide + acid → metal salt + $H_2S(g)$** |
| $Na_2S(aq) + 2\ HCl(aq) \rightarrow 2\ NaCl(aq) + H_2S(g)$ |
| **Metal sulfite + acid → metal salt + $SO_2(g)$ + $H_2O(\ell)$** |
| $Na_2SO_3(aq) + 2\ HCl(aq) \rightarrow 2\ NaCl(aq) + SO_2(g) + H_2O(\ell)$ |
| **Ammonium salt + strong base → metal salt + $NH_3(g)$ + $H_2O(\ell)$** |
| $NH_4Cl(aq) + NaOH(aq) \rightarrow NaCl(aq) + NH_3(g) + H_2O(\ell)$ |

All metal carbonates ( $CO_3^{2-}$ ) and bicarbonates ( $HCO_3^-$ ) react with acids to produce carbonic acid which can decompose to carbon dioxide:

$$CaCO_{3\,(s)} + 2\ HCl_{(aq)} \rightarrow CaCl_{2\,(aq)} + H_2CO_{3\,(aq)}$$

$$H_2CO_{3\,(aq)} \rightarrow H_2O_{(l)} + CO_2 \uparrow_{(g)}$$

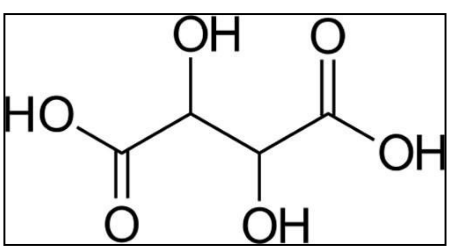
Overall: $CaCO_{3\,(s)} + 2\ HCl_{(aq)} \rightarrow CaCl_{2\,(aq)} + H_2O_{(l)} + CO_2 \uparrow_{(g)}$

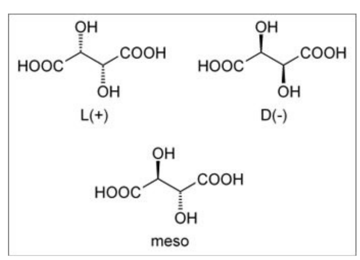
Calcium Carbonate ($CaCO_{3\,(s)}$) is what makes water hard and leaves white marks on cars and other things that hard water dries on. It will react with vinegar (dilute acetic acid) to form soluble acetate:

$$CaCO_{3\,(s)} + 2\ CH_3COOH_{(aq)} \rightarrow Ca(CH_3COO)_{2\,(aq)} + H_2O_{(l)} + CO_2 \uparrow_{(g)}$$

How does bread rise and have holes – by the formation of carbon dioxide from the bicarbonate of soda (baking soda) which reacts with the small amount of tartaric acid also present in baking soda:

| | | | |
|---|---|---|---|
| $H_2C_4H_4O_{6\,(aq)}$ + | $HCO_3^-{}_{(aq)}$ | → | $HC_4H_4O_6^-{}_{(aq)} + H_2O_{(l)} + CO_{2\,(g)}$ |
| Tartaric Acid | Bicarbonate Ion | | Hydrogen Tartrate Ion |

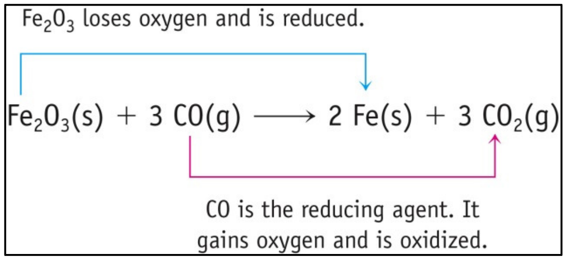





**Example 3.7** Write the balanced equation for the reaction of nickel (II) carbonate with sulfuric acid.

# 3.8 REDOX – Oxidation Reduction Reactions

**Oxidation:** Loss of electrons **Reduction:** Gain of electrons




**Oxidation Numbers** (ON) are the charge an element has or appears to have:

| | | |
|---|---|---|
| 1. Pure Element Oxidation Number | = 0 | ON for Cu is zero |
| 2. Monoatomic ions OxNum | = charge for that ion | ON for $Mg^{2+}$ is +2 |
| 3. Halogens | = -1 | $F^-$ is -1 |
| 4. Oxygen is -2, oxide | = -2 | $H_2O$, Oxygen is -2 |
| Peroxide | = -1 | $H_2O_2$, Oxygen is -1 |
| 5. H is +1, | = +1 | $H_2O$, Hydrogen is +1 |
| hydride is -1 | = -1 | NaH, Hydrogen is -1 |

The algebraic sum of the OxNum for a molecule **must equal Zero**.

**Oxidation** is a Loss of Electrons Increase in Oxidation Number

**Reduction** is a Gain of Electrons Reduction in Oxidation Number

**Oxidation** is a process in which oxygen is added to another substance

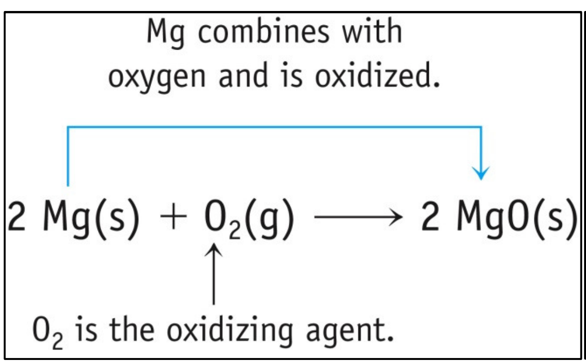
**Oxidation Agent** – a compound that oxidizes another compound, the oxidation agent is reduced

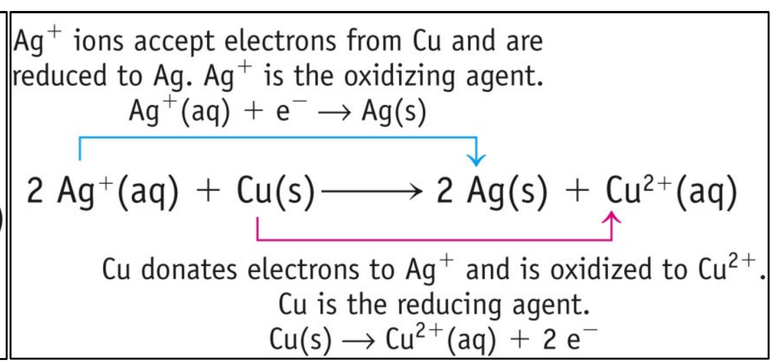
**Reducing Agent** – a compound that reduces another compound, the reducing agent is oxidized

|<- Oxidation -> | Iron goes from Zero to +2

**Fe** + $\mathbf{Cu^{+2}} \rightarrow$ $\mathbf{Fe^{+2}}$ + Cu

|<- Reduction -> | Copper (II) goes from +2 to Zero



Mg releases 2 $e^-$ per atom. Mg is oxidized to $Mg^{2+}$ and is the reducing agent.

$$2\ Mg(s) + O_2(g) \longrightarrow 2\ MgO(s)$$

$O_2$ gains 4 $e^-$ per molecule to form 2 $O^{2-}$. $O_2$ is reduced and is the oxidizing agent.

Na releases 1 $e^-$ per atom.
Oxidation number increases.
Na is oxidized to $Na^+$ and is the reducing agent.

$$2\ Na(s) + Cl_2(g) \longrightarrow 2\ NaCl(s)$$

$Cl_2$ gains 2 $e^-$ per molecule.
Oxidation number decreases by 1 per Cl.
$Cl_2$ is reduced to $Cl^-$ and is the oxidizing agent.

Nitric Acid ($HNO_3$) is a strong oxidizing agent in water dissociates to $H^+$ and $NO_3^-$

Oxidation number of Cu changes from 0 to +2. Cu is oxidized to $Cu^{2+}$ and is the reducing agent.

$$Cu(s) + 2\ NO_3^-(aq) + 4\ H_3O^+(aq) \longrightarrow Cu^{2+}(aq) + 2\ NO_2(g) + 6\ H_2O(\ell)$$

N in $NO_3^-$ changes from +5 to +4 in $NO_2$. $NO_3^-$ is reduced to $NO_2$ and is the oxidizing agent.

***** The easiest way to spot a **REDOX Reaction** is there is a **PURE ELEMENT** on one side of the equation. *****

**Example 3.8** Determine the oxidation number for: Aluminum Oxide, Phosphoric Acid, Sulfur in Sulfate Ion, each Cr in Dichromate ion.

## Recognizing REDOX Reactions

**Table 3.3** Common Oxidizing and Reducing Agents

| Oxidizing Agent | Reaction Product | Reducing Agent | Reaction Product |
|---|---|---|---|
| $O_2$, oxygen | $O^{2-}$, oxide ion or O combined in $H_2O$ or other molecule | $H_2$, hydrogen | $H^+(aq)$, hydrogen ion or H combined in $H_2O$ or other molecule |
| Halogen, $F_2$, $Cl_2$, $Br_2$, or $I_2$ | Halide ion, $F^-$, $Cl^-$, $Br^-$, or $I^-$ | M, metals such as Na, K, Fe, and Al | $M^{n+}$, metal ions such as $Na^+$, $K^+$, $Fe^{2+}$ or $Fe^{3+}$, and $Al^{3+}$ |
| $HNO_3$, nitric acid | Nitrogen oxides* such as NO and $NO_2$ | C, carbon (used to reduce metal oxides) | CO and $CO_2$ |
| $Cr_2O_7^{2-}$, dichromate ion | $Cr^{3+}$, chromium(III) ion (in acid solution) | | |
| $MnO_4^-$, permanganate ion | $Mn^{2+}$, manganese(II) ion (in acid solution) | | |

*NO is produced with dilute $HNO_3$, whereas $NO_2$ is a product of concentrated acid.

1. Determine the oxidation number and see if it changes in a reaction
2. If there is a "Pure Element" on either side of the arrow, it is a Redox
3. If any of the above are involved, it is a Redox

**Table 3.4 Recognizing Oxidation–Reduction Reactions**

| | Oxidation | Reduction |
|---|---|---|
| In terms of oxidation number | Increase in oxidation number of an atom | Decrease in oxidation number of an atom |
| In terms of electrons | Loss of electrons by an atom | Gain of electrons by an atom |
| In terms of oxygen | Gain of one or more O atoms | Loss of one or more O atoms |

Oxidation number of Cu changes from 0 to +2. Cu is oxidized to $Cu^{2+}$ and is the reducing agent.

$$Cu(s) + 2\ NO_3^-(aq) + 4\ H_3O^+(aq) \longrightarrow Cu^{2+}(aq) + 2\ NO_2(g) + 6\ H_2O(\ell)$$

N in $NO_3^-$ changes from +5 to +4 in $NO_2$. $NO_3^-$ is reduced to $NO_2$ and is the oxidizing agent.

Metals usually loose electrons in a chemical reaction (except for Thermite below)

$$\underbrace{Fe_2O_3(s)}_{\text{oxidizing agent}} + \underbrace{2\ Al(s)}_{\text{reducing agent}} \longrightarrow 2\ Fe(\ell) + 2\ Al_2O_3(s)$$

Thermite Reaction gives off lots of heat

**Reactions in Aqueous Solutions**

Precipitation, Acid Base, Gas Forming are EXCHANGE REACTIONS

**Precipitation Reactions:** Reactant ions form an insoluble product

Overall Reaction: $Pb(NO_3)_2\,_{(aq)} + 2\ KI_{(aq)} \rightarrow PbI_2 \downarrow_{(s)} + 2\ KNO_3\,_{(aq)}$

Net Ionic Equation $Pb^{2+}_{(aq)} + 2\ I^- \rightarrow PbI2 \downarrow_{(s)}$

**Acid-Base Reactions:**

Reaction of a strong acid and a strong base usually results in water and a salt products

Overall Reaction: $HNO_3\,_{(aq)} + KOH_{(aq)} \rightarrow H\text{-}OH_{(l)} + KNO_3\,_{(aq)}$

Net Ionic Equation $H_3O^+_{(aq)} + OH^-_{(aq)} \rightarrow 2\ H\text{-}OH_{(l)}$

Reaction of a weak acid and a strong base

Overall Reaction: $CH_3COOH_{(aq)} + NaOH_{(aq)} \rightarrow Na{+}CH_3COO^-_{(aq)} + H\text{-}OH_{(l)}$

Net Ionic Equation $CH_3COOH_{(aq)} + OH^-_{(aq)} \rightarrow CH_3COO^-_{(aq)} + H\text{-}OH_{(l)}$

**Gas-Forming Reaction**: Usually a metal carbonate and an acid

Overall Reaction $CuCO_{3\,(s)} + 2\,HNO_{3\,(aq)} \rightarrow Cu(NO_3)_{2\,(aq)} + CO_{2\,(g)} + \text{H-OH}_{(l)}$

Net Ionic Equation $CuCO_{3\,(s)} + 2\,H_3O^+_{(aq)} \rightarrow Cu^{2+}_{(aq)} + CO_{2\,(g)} + 3\,\text{H-OH}_{(l)}$

Alka-Seltzer and water:

$$\underset{\text{citric acid}}{H_3C_6H_5O_7(aq)} + \underset{\text{hydrogen carbonate ion}}{HCO_3^-(aq)} \longrightarrow \underset{\text{dihydrogen citrate ion}}{H_2C_6H_5O_7^-(aq)} + H_2O(\ell) + CO_2(g)$$

**REDOX Reaction**: These are NOT exchange reactions, but involve electron transfer

Overall Equation $Cu\downarrow_{(s)} + 2\,AgNO_{3\,(aq)} \rightarrow Cu(NO_3)_{2\,(aq)} + 2\,Ag\downarrow_{(s)}$

Net Ionic Equation $Cu\downarrow_{(s)} + 2\,Ag^+_{(aq)} \rightarrow Cu^{2+}_{(aq)} + 2\,Ag\downarrow_{(s}$

## **From my other Lecture Notes (Different Text Book):**

**Solubility – ability do dissolve in water. Solubility Rules for Ionic Compounds [ Table 4.2 ]**

| # | Applies to | Statement | Exceptions |
|---|---|---|---|
| 1. | $Li^+$, $Na^+$, $K^+$, $NH_4^+$ | Group 1A and Ammonium cpds **are soluble** | |
| 2. | $C_2H_3O_2^-$, $NO_3^-$ | Acetates & Nitrates **are soluble** | |
| 3. | $Cl^-$, $Br^-$, $I^-$ | Most Chloride, Bromide & Iodides **are soluble** | $AgX$, $Hg_2X_2$, $PbX_2$ X = Cl, Br, I |
| 4. | $SO_4^{-2}$ | Most Sulfates **are soluble** | $CaSO_4$, $SrSO_4$, $BaSO_4$ $Ag_2SO_4$, $Hg_2SO_4$, $PbSO_4$ |
| 5. | $CO_3^{-2}$ | Most carbonates **are INSOLUBLE** | Grp 1A, $(NH_4)_2CO_3$ |
| 6. | $PO_4^{-3}$ | Most phosphates **are INSOLUBLE** | Grp 1A, $(NH_4)_3PO_4$ |
| 7. | $S^{-2}$ | Most sulfides **are INSOLUBLE** | Grp 1A, $(NH_4)_2S$ |
| 8. | $OH^-$ | Most hydroxides **are INSOLUBLE** | Grp 1A, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$, $NH_4OH$ |

Compounds that dissolve in water **are soluble**.
Compounds that dissolve only a little **are INSOLUBLE**
Soluble compounds are Electrolytes or Non-Electrolytes
Electrolytes can be Strong or Weak
Non-Electrolytes form non electrical conducting solutions.

## Common Acids and Bases Table 4.2

| **Name** | | **Formulae** | **Remarks** |
|---|---|---|---|
| **Acid** | Acetic Acid | $HC_2H_3O_2$ | Vinegar |
| | Acetylsalicylic Acid | $HC_9H_7O_4$ | Aspirin |
| | Ascorbic Acid | $H_2C_6H_6O_6$ | Vitamin C |
| | Citric Acid | $H_3C_6H_5O_7$ | In Lemon Juice |
| | Hydrochloric Acid | $HCl$ | Stomach Acid |
| | Sulfuric Acid | $H_2SO_4$ | Battery Acid |
| **Base** | Ammonia | $NH_3$ [ $NH_4OH$ ] | Water solution is a household cleaner |
| | Calcium Hydroxide | $Ca(OH)_2$ | Lime use in construction mortar |
| | Magnesium Hydroxide | $Mg(OH)_2$ | Mild of magnesia – antacid |
| | Sodium Hydroxide | $NaOH$ | Drain and oven cleaner |

**<u>3. Oxidation Reduction Reactions [ Redox ]</u>** are reactions that involve transfer of electrons form one species to another or in which the oxidation number changes.

An Iron nail in Copper (II) Sulfate: $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$

**The Net Ionic is** $Fe + Cu^{+2} \rightarrow Fe^{+2} + Cu$

$Fe^{o} \rightarrow Fe^{+2} + 2e^{-}$

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

**Oxidation Number** is the actual charge of the atom if it exists as a monoatomic ion – or hypothetical charge.
The Oxidation Number: **<u>Rules 4 Assigning Oxidation Numbers – Table 4.5</u>**

1. an atom / element is ZERO. Na = Metallic Sodium = 0
2. of an atom that exists in a compound as a monoatomic ion equals the charge on that ion.
   NaCl Na = +1, Cl = -1
3. Oxygen in a compound has an Oxidation Number of -2. e.g. In $SO_2$, O = -2 each, S = +4
   Exception is $H_2O_2$ where H = +1 and O = -1 each
4. Hydrogen in a compound has an Oxidation Number of +1
   Exception is when combined with a metal to form a Hydride NaH Na = +1, H = -1
5. Halogens in a compound have an Oxidation Number of -1.
   Except when combined with a halogen above it in the PT. [ Never saw one yet thought! ]
   Or when combine with Oxygen.
6. The sum of the Oxidation Numbers in a compound is ZERO.
   The sum of the Oxidation Numbers in a polyatomic ion equals it's charge.

Oxidation Numbers > +6 or < -4 are probably in error.

| | $2\ Ca + O_2 \rightarrow 2\ CaO$ | |
|---|---|---|
| | 2 Ca + $O_2$ → 2 CaO | **Calcium is Oxidized** |
| Ox Number | 0 0 +2 -2 | **Oxygen is Reduced** |

Calcium goes from an Oxidation Number of 0 to +2
Oxygen goes from an Oxidation Number of 0 to -2

**Problem:** Determine the Oxidation Number of Chlorine in:

| | | | |
|---|---|---|---|
| A. Perchloric Acid | $HClO_4$ | H = +1, O = 4 * -2 | Cl = +7 |
| B. Chlorate Ion | $ClO_3^-$ | O = 3 * -2, Net Charge = -1 | Cl = +5 |

**<u>Half Reactions i</u>**s one of the two parts of a Redox Reaction.
One part has loss of e- or gain of oxidation number, one gain of e- or decrease of oxidation number.

An Iron nail in Copper (II) Sulfate: $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$

**The Net Ionic is** $Fe + Cu^{+2} \rightarrow Fe^{+2} + Cu$

**OXIDATION** $Fe^{o} \rightarrow Fe^{+2} + 2e^{-}$ **electrons lost by Fe**

**REDUCTION** $Cu^{+2} + 2e^{-} \rightarrow Cu^{o}$ **gained by Cu**

**Oxidation** is a **LOSS OF ELECTRONS.** **Reduction** is a **GAIN OF ELECTRONS**

Oxidation Agent – a compound that oxidizes another compound

Reducing Agent – a compound that reduces another compound

**|<- Oxidation ->|**

$Fe + Cu^{+2} \rightarrow Fe^{+2} + Cu$

|<- **Reduction** ->|

## Common Oxidation – Reduction Reactions

1. Combination
2. Decomposition
3. Displacement
4. Combustion

**1. Combination Reaction** is one in which two substances combine to form a third compound

$2\,Na + Cl_2 \rightarrow 2\,NaCl$ Sodium and Chlorine

$2\,Sb + 3\,Cl_2 \rightarrow 2\,SbCl_3$ Antimony and Chlorine

$CaO + SO_2 \rightarrow CaSO_3$ There is no change in Oxidation Numbers
But this is still a Combination Reaction

**2. Decomposition Reaction** is one in which a single compound reacts to give two or more substances. Check Oxidation Number to see if they are Redox – some are not!

$2\,HgO \xrightarrow{Heat} 2\,Hg + O_2$ Heat Mercury (II) Oxide

$2\,KClO_3 \xrightarrow[MnO_2\ Cat]{Heat} 2\,KCl + 3\,O_2$ Heat Potassium Chlorate with MnO2 Cat
Redox

$CaCO_3 \xrightarrow{Heat} CaO + CO_2$ Head Calcium Carbonate not Redox

**3. Displacement or Single Displacement** is were an Element reacts with Compound, displacing an element from the compound. If an element is involved, the reaction is a Redox.

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

$Cu + 2\,Ag+ \rightarrow Cu^{+2} + 2\,Ag$ Net Ionic shows electron transfer

$Zn + 2\,HCl \rightarrow ZnCl_2 + H_2$ Zinc and HCl yields Hyrdorgen Gas

$Zn + 2\,H^+ \rightarrow Zn^{+2} + H_2$ Net Ionic

**Activity Series of the Elements [ Table 4.6 ]**

Li > K > Ba > Ca > Na > Reacts violently with water to give $H_2$

Mg > Al > Zn > Cr > Fe > Cd > Reacts slowly with water to give $H_2$

Co > Ni > Sn > Pb

H2 > Cu > Hg > Ag > Au Do not react with acids to give $H_2$

**4. Combustion Reaction**s a substance reacts with oxygen usually with the rapid release of heat to produce a flame. Butane

$2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$ + Heat

$4\ Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$ Metals burn in air, iron rusts in oxygen

## Balancing Redox Equations

1st Glance $Zn + Ag^+ \rightarrow Zn^{+2} + Ag$

But the charge is not balanced. Do it by Half Reactions

ON 0 +1 +2 0

$Zn + Ag^+ \rightarrow Zn^{+2} + Ag$

$Zn \rightarrow Zn^{+2} + 2e^-$ Oxidation

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

Balance the electrons

$Zn \rightarrow Zn^{+2}\ 2e^-$ Oxidation

$2\ Ag^+ + 2\ e^- \rightarrow 2\ Ag$ Reduction

$\mathbf{Zn + 2\ Ag^+ \rightarrow 2\ Ag + Zn^{+2}}$ **BALANCE ELECTRONS**

0 0 +2 -3

$Mg + N_2 \rightarrow Mg_3N_2$ Magnesium metal and Nitrogen Gas react

$Mg \rightarrow Mg^{+2} + 2\ e^-$ Oxidation – Need 3 of these

$N_2 + 6\ e^- \rightarrow 2\ N^{-3}$ Reduction

$\mathbf{3\ Mg + N_2 \rightarrow 3\ Mg^{+2} + 2\ N^{-3}}$ **[ + / - 6 e⁻ ] BALANCE ELECTRONS**