# Types of Chemical Reactions

Synthesis: Only 1 Product

Example: 2 Na (s) + Cl2 (g) 2 NaCl (s) element + element 🡪 compound

Example: **H2O** (l) + **SO3** (g) 🡪 H2SO4 (aq) small + small 🡪 bigger compound

Decomposition: Only 1 Reactant

Example: 2 H2O (l) 🡪 2 H2 (g) + O2 (g) compound 🡪 element + element

Example: CaCO3 (s) 🡪 CaO (s) + CO2 (g)  bigger compound 🡪 small + small

Single Replacement: Must have elements on both sides

***If a metal element is a more reactive, it will replace the metal in the ionic compound.***

Example: 2 Al (s) + 3 Ag2O (aq)  Al2O3 (aq) + 2 Ag (s) aluminum swaps with silver

***If a nonmetal element is a more reactive, it will replace the nonmetal in the ionic compound.***

Example: Cl2 (g) + 2 NaBr (aq)  2 NaCl (aq) + Br2 (s) chlorine swaps with bromine

Double Replacement: All Compounds NO WATER!!!

Example: Pb(NO3)2 (aq) + 2 KI(aq)  2 KNO3 (aq) + PbI (s)

For Double Replacement the outsides go together and the insides go together, but it

DOES NOT MAKE WATER!!!!

# Types of Chemical Reactions

Neutralization: Acid + Base 🡪 Salt + WATER!!!

Example: H(NO3)2 (aq) + 2 KOH(aq)  2 KNO3 (aq) + H(OH) (l)

Acids have **(H+1)** in front and Bases have **(OH–1)** in back. These combine to make **WATER**.

To help balance this equation, rewrite **WATER** as **H(OH)**.

Notice that water is always a **LIQUID** in this type of reaction.

Combustion: Hydrocarbon + O2 🡪 CO2 + WATER!!!

Example: 1 C5H12 (l) + 8 O2(g)  5 CO2 (g) + 6 H2O (g)

Example: 2 C3H8O (l) + 9 O2(g)  6 CO2 (g) + 8 H2O (g)

Combustion equations always end in **O2 🡪 CO2 + H2O**,

which are all gases due to the high temperature when burning the hydrocarbon.

**DO NOT WRITE WATER AS H(OH)!!!!**

**Follow these steps to balance combustion equations:**

**1.) Swap the hydrogens. Write the product’s subscript as the reactant’s coefficient and the reactant’s subscript as the product’s coefficient.**

**2.) Balance the carbon atoms. Multiply the coefficient by the subscript on the reactant side to get the coefficient on the reactant side.**

**3.) Add up all the oxygen atoms on the product side, and then divide by 2. (Note: If the hydrocarbon has oxygen, you must subtract that before dividing by 2.)**

**4.) Simplify… If all coefficients are even, then divide by 2.**

Incomplete Combustion

When there is a limited supply of oxygen gas, the normal combustion reaction *does not* produce carbon dioxide; instead, it produces carbon or carbon monoxide.

**Example 1: Some oxygen available, but not an excess amount of oxygen.**

**Hydrocarbon + Oxygen**  **Carbon Monoxide + Water**

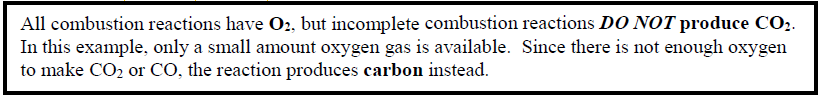
2 C4H10 + 9 O2  8 CO + 10 H2O

All combustion reactions have **O2**, but incomplete combustion reactions ***DO NOT* produce CO2.** In this example, some oxygen gas is available but not enough to make CO2, so the reaction produces **carbon monoxide** instead.

**Example 2: Very little oxygen available. (Think of a sealed room or flask.)**

**Hydrocarbon + Oxygen**  **Carbon + Water**

2 C4H10 + 5 O2  8 C + 10 H2O



## Note: When you compare the 3 different combustion reactions, the incomplete combustion reactions have lower coefficients for the oxygen because it is the limiting reactant.

### 2 C4H10 + 13 O2  8 CO2 + 10 H2O Complete with Carbon Dioxide

2 C4H10 + 9 O2  8 CO + 10 H2O Incomplete with Carbon Monoxide

2 C4H10 + 5 O2  8 C + 10 H2O Incomplete with Carbon