# Classifying Chemical Reactions

Chemical reactions may be divided into six classifications based on the nature of the particular reaction.

***Synthesis reactions*** *occur when 2 or more substances combine to form a compound.*

$$A + B \text{ ----------> } AB$$

The formation of water from hydrogen and oxygen is a typical synthesis reaction.

$$2H_{2(g)} + O_{2(g)} \text{ ------------> } 2H_2O_{(l)}$$

Two other examples are the formation of calcium oxide and the formation of iron oxide as illustrated below.

$$Ca_{(s)} + H_2O_{(l)} \text{ ------------> } Ca(OH)_2(s)$$
$$2\,Fe_{(s)} + O_{2(g)} \text{ ------------> } 2\,FeO(s)$$

***Decomposition reactions*** *are the opposite of synthesis reactions and occur when a compound breaks down into 2 or more simpler substances.*

$$AB \text{ ----------> } A + B$$

The decomposition of water into hydrogen and oxygen is aan example of theis type of reaction.

$$2H_2O_{(l)} \text{ ------------> } 2H_{2(g)} + O_{2(g)}$$

Two other examples are the decomposition of of calcium oxide and the decomposition of iron oxide as illustrated below.

$$Ca(OH)_2(s) \text{ ------------> } Ca_{(s)} + H_2O_{(l)}$$
$$2FeO(s) \text{ ------------> } 2Fe_{(s)} + O_{2(g)}$$

***Combustion*** *involves the reaction of a substance with oxygen releasing large amounts of light and heat energy.* A familiar form of this type of reaction involves the burning of carbon-based fuels (hydrocarbons).

$$C_xH_y + O_2 \text{ ---------> } CO_2 + H_2O$$

Consider the following examples:

$C_3H_8 + 5\,O_2$ ----------> $3\,CO_2 + 4\,H_2O$ [combustion of propane]
$C_7H_{16} + 11O_2$ ----------> $7\,CO_2 + 8\,H_2O$ [combustion of heptane]

Other materials will also burn in the presence of oxygen. The burning of magnesium is a typical example.

$$Mg + O_2 \text{ ---------> } 2\,MgO$$

(Note that this equation is also an example of a synthesis reaction)

***Single replacement reactions*** *occur when one element is replaced by another in a compound.* This usually occurs when a compound and an element are combined.

$$A + BC \text{ ---------> } AC + B$$

Note that A will replace B in the above reaction only if A is chemically more active than B. A chart showing the relative activity of metals and nonmetals ((Table 5-2) may be found on page 136 of *Heath Chemistry.*

Consider the following examples:

$$Cl_2 + 2\,KBr \text{ ---------> } 2\,KCl + Br_2$$

(Chlorine is more active than bromine and replaces it in the compound)

$$Fe + CuSO_4 \text{ ---------> } FeSO_4 + Cu$$

(Iron is more active than copper and replaces it in the compound)

$$Al + MgCl_2 \text{ ---------> no reaction}$$

(Aluminum is less active than magnesium and **does not** replace it in the compound)

***Double replacement reactions*** *occur when the elements in solutions of two reacting compounds exchange places.*

$$AB + XY \text{ ---------> } AY + XB$$

Consider the following examples:

$$ZnBr_{2(aq)} + 2\,AgNO_{3(aq)} \text{ ---------> } 2\,AgBr_{(s)} + Zn(NO_3)_{2(aq)}$$

(Silver bromide forms a solid [precipitate] while the zinc nitrate remains in solution)

$$Pb(NO_3)_{2(aq)} + 2\,KI_{(aq)} \text{ ---------> } PbI_{2(s)} + 2\,KNO_{3(aq)}$$

(Lead iodide forms a solid [precipitate] while the potassium nitrate remains in solution)

***Water forming reactions*** *are double replacement reactions where water is one of the products formed.*

$$HB + XOH \text{ ---------> } XB + HOH$$

(Note that HOH may also be written as $H_2O$)

Consider the following examples:

$$HCl_{(aq)} + NaOH_{(aq)} \text{ ---------> } NaCl_{(aq)} + H_2O_{(l)}$$

$$2\ H_3PO_{4(aq)} + 3\ Ba(OH)_{2(aq)} \text{ ---------> } Ba_3(PO_4)_{2(aq)} + 6\ H_2O_{(l)}$$

Awareness of the above classifications allows us to more easily predict the products of chemical reactions. Practice this skill by completing, balancing and classifying the following reactions:

1. $Mg + I_2$ --------------->
2. $C_6H_{14} + O_2$ --------------->
3. $Mg + AlCl_3$ --------------->
4. $HCl_{(aq)} + Sr(OH)_{2(aq)}$ --------------->
5. $2\ Fe_2O_3$ --------------->
6. $Pb(NO_3)_{2(aq)} + Na_2Cr_2O_{2(aq)}$ --------------->

**Assignment**: Complete questions 1 to 4 on page 139 of Heath Chemistry.

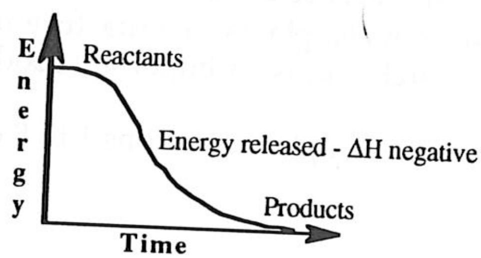
# Energy Changes in Chemical Reactions

During a chemical reaction chemical bonds are broken and new bonds are formed. Energy is required to break chemical bonds and energy is released when new chemical bonds are formed. A chemical reaction will not take place unless there is sufficient energy available to break the chemical bonds holding the atoms of the reactants together. At room temperature, there is often enough energy in the ordinary jostling of the molecules to allow a chemical reaction to proceed. In other cases, however, energy must be added to the system (usually as heat) to permit the reaction to occur.

*When a chemical reaction releases energy into its surroundings we describe it as* ***exothermic.*** In an exothermic reaction, more energy is released when new bonds form than was required to break the original bonds. The excess energy is usually released as heat, light or both. The standard unit for measuring energy exchange in chemical reactions is the joule. *A* ***joule*** *(J)is the amount of energy produced when a force of 1 newton acts over a distance of 1 metre.* This amount of energy is quite small, therefore we usually express energy exchange in units of 1000 joules (kilojoules or KJ). When writing exothermic chemical reactions we usually show the release of energy as a product of the reaction. Consider the following example:

$$2\,C_8H_{18} + 25\,O_2 \text{ ----------> } 16\,CO_2 + 18\,H_2O + \text{energy}$$

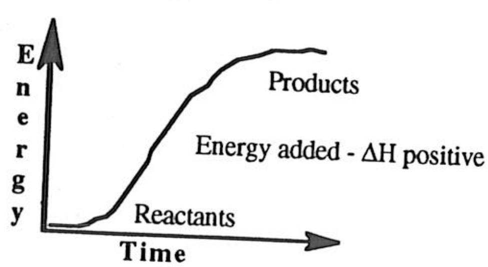
Since heat is released, the products of the reaction are at a lower energy level *(or enthalpy)* than the reactants. Thus, the energy exchange *(ΔH [read delta-H])* has a negative value.




*When a chemical reaction requires more energy to break the bonds of the reactants than is supplied by the formation of new bonds in the products, we describe the reaction as* ***endothermic.*** The additional energy required is usually supplied as heat, but light or electrical energy may also be used. When writing endothermic chemical reactions we usually show the addition of energy on the reactant side of the equation. Consider the following example:

$$2H_2O + \text{(electrical) energy} \text{ ----------> } 2\,H_2 + O_2$$

Since energy is added, the products of the reaction are at a higher energy level than the reactants. Thus, the energy exchange *(ΔH )* has a positive value.



## Some Everyday Examples:

The combustion of gasoline in an automobile engine is a good example of an *exothermic* reaction. An important component of gasoline is *octane* ($C_8H_{18}$). The combustion of the gasoline is initiated by a spark and the reaction proceeds as follows:

$$2\,C_8H_{18} + 25\,O_2 \text{ ------------> } 16\,CO_2 + 18\,H_2O + \text{energy}$$

Most of the energy used to move the pistons in an internal combustion engine comes from the hot expanding exhaust gases inside the cylinders. The exhaust gases are heated by the energy released in the above reaction.

Photosynthesis in green plants is a good example of an *endothermic* process. During photosynthesis, light energy is used to assist the formation of glucose and oxygen from carbon dioxide and water. The following equation summarizes the reaction:

$$6\,CO_2 + 6\,H_2O + \text{(light) energy} \text{ ------------> } C_6H_{12}O_2 + 6\,O_2$$

The glucose ($C_6H_{12}O_2$) formed has a higher energy content than the reactants ($CO_2$ and $H_2O$) and acts as a form of stored energy for the plant. Fortunately for humans and other animals, plants manufacture more glucose than they need. The excess is stored as starch and is an important food source for most living organisms.

**Assignment:** Complete questions 1 to 8 on page 143 of Heath Chemistry.