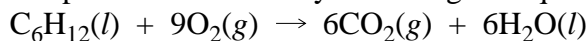


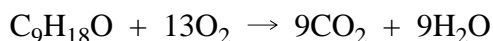
There are very many chemical reactions known. Fortunately, there are not that many *types* of reaction. Hence, we can classify chemical reactions according to the type of reaction. There are many classifications possible and one reaction may be classified in more than one way. All classifications are useful because they help in understanding and remembering many of the known reactions and in predicting some new reactions. Many reactions fall into one of the following **five** classes: (1) **Combustion**, (2) **Combination**, (3) **Decomposition**, (4) **Single Replacement**, and (5) **Double Replacement**.

### COMBUSTION REACTIONS

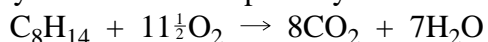
Many compounds are flammable; they can burn when reacted with oxygen. The reaction is called combustion; it gives off a lot of heat but heat must be supplied to start the reaction. We will consider only combustion reactions of compounds containing C and H, and those containing C, H and O. The **only** products of complete combustion of such compounds are CO<sub>2</sub> and H<sub>2</sub>O. Hence, each molecule of the compound produces a number of CO<sub>2</sub> molecules equal to the number of C atoms per molecule and a number of H<sub>2</sub>O molecules equal to *half* the number of H atoms per molecule. The number of O<sub>2</sub> molecules needed per molecule of compound is obtained by balancing the equation.



The physical state of the compound would have to be given; both O<sub>2</sub> and CO<sub>2</sub> exist as gases (except at very low temperatures), and H<sub>2</sub>O would exist as a liquid after cooling (it can also be written as a gas). When the molecule contains oxygen, the number of O atoms in the molecule must be counted when balancing the equation.



The equation may have to be multiplied by 2 to make the number of O<sub>2</sub> molecules a whole number.



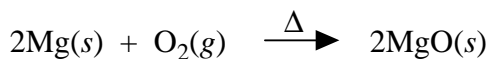
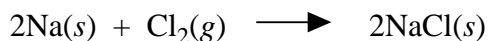
becomes 
$$2\text{C}_8\text{H}_{14} + 23\text{O}_2 \rightarrow 16\text{CO}_2 + 14\text{H}_2\text{O}$$

### COMBINATION REACTIONS (A + B → C)

A combination reaction is one in which two or more reactants form one product. Combination reactions can be assigned to several sub-classes, including the following.

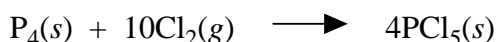
#### (a) **Metal + non-metal → ionic compound**

The combination of a metal and a non-metal produces a binary ionic compound. For the **A** group metals, the charge on the derived cation is known and hence the formula of the ionic product is *predictable*. Some reactions, such as that between Na and Cl<sub>2</sub>, are spontaneous while others, such as that between Mg and O<sub>2</sub>, require heating.



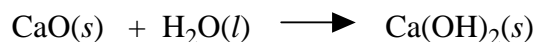
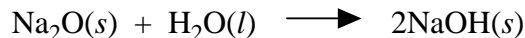
#### (b) **Non-metal + non-metal → molecular compound**

The combination of two non-metals produces molecular compounds but the product is *not predictable* in most cases. Thus, the reaction of N<sub>2</sub> with O<sub>2</sub> can form NO, N<sub>2</sub>O, NO<sub>2</sub>, N<sub>2</sub>O<sub>3</sub>, N<sub>2</sub>O<sub>5</sub> and N<sub>2</sub>O<sub>4</sub>. Similarly, P can form two different compounds on reaction with Cl<sub>2</sub>.

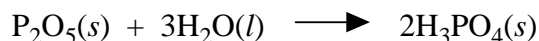
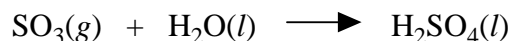


(c) **Metal oxide + water** → **metal hydroxide**

Some examples are:

(d) **Non-metal oxide + water** → **oxyacid**

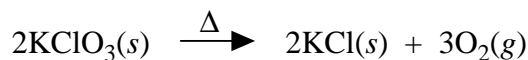
Some examples are:

**DECOMPOSITION REACTIONS** ( $A \rightarrow B + C$ )

Most decomposition reactions are not easily predictable but many carbonates decompose on heating to the oxide and  $\text{CO}_2$ .



Other examples are:

**SINGLE REPLACEMENT REACTIONS** ( $A + BC \rightarrow AC + B$ )

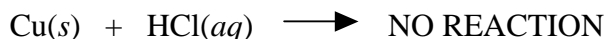
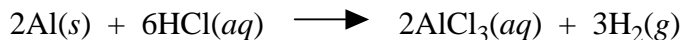
Many single replacement reactions involve the reaction of a metal with  $\text{H}_2\text{O}$ , an acid, or the salt of another metal [(a) to (c), below]. The **Activity Series** (see below) can be used to predict if a reaction occurs.

(a) **Active metal + water** → **metal hydroxide +  $\text{H}_2$** 

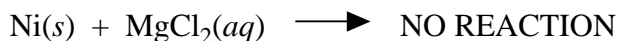
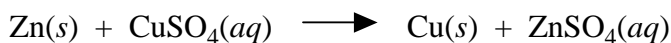
The first five metals in the activity series (Li to Na) are reactive enough to displace H from  $\text{H}_2\text{O}$ .

(b) **Metal + acid** → **metal salt +  $\text{H}_2$** 

Metals **above** H in the activity series will displace H from an acid.

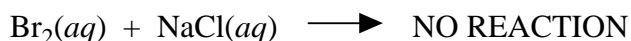
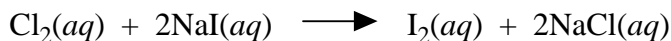
(c) **Metal A + metal B salt** → **metal B + metal A salt**

A metal will displace a **lower** (in the activity series) metal.



(d) **Halogen A + halide B → halogen B + halide A**

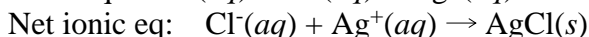
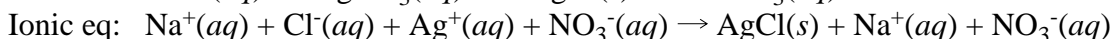
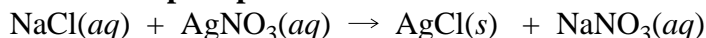
A halogen will displace a **less reactive** halogen (lower in the Periodic Table) from its halide salt.

**DOUBLE REPLACEMENT REACTIONS** ( $\text{AB} + \text{CD} \rightarrow \text{AD} + \text{CB}$ )

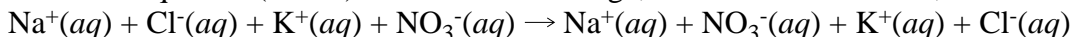
When two ionic compounds are mixed in aqueous solution, a reaction occurs if there is (a) an insoluble product, or (b) a molecular product. In other cases, there is simply the same mixture of ions and no change has occurred. Ionic and net ionic equations (see below) help to show this. The **Solubility Rules** (see below) are used to predict if any of the "products" is insoluble.

(a) **Insoluble product (Precipitation reaction)**

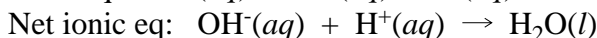
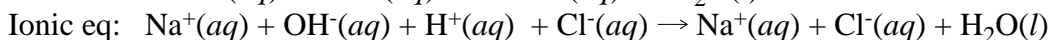
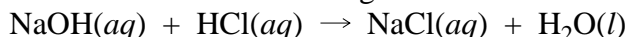
The solid that appears when an insoluble product is formed from mixing two aqueous solutions of ionic compounds is called a **precipitate** and hence such reactions are also called **precipitation reactions**.



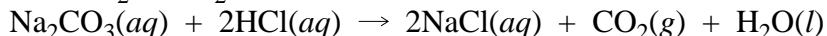
The ionic equation (below) shows that no change, and hence no reaction, occurs.

(b) **Molecular product (Acid-Base reaction)**

The most common type of double replacement reaction in which a molecular compound is formed is the reaction of an acid with a base forming an ionic salt and the molecular compound water.



The reaction of an acid with a carbonate or a hydrogen carbonate is also an acid-base reaction. In these cases,  $\text{CO}_2$ , as well as water, is formed. The  $\text{H}_2\text{CO}_3$ , which is expected from the double replacement, decomposes to  $\text{CO}_2$  and  $\text{H}_2\text{O}$ .

**REDOX REACTIONS (Oxidation-reduction reactions)**

A redox reaction is one in which there is a transfer of electrons from one reactant to another. Many of the reactions discussed above are redox reactions. When ionic compounds are involved, it is easy to see the electron transfer. Many reactions of molecular compounds are also redox reactions but it is more difficult to see this and hence such redox reactions are outside the scope of this course.

(a) **Metal + non-metal → ionic compound**

In the reaction  $2\text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\text{NaCl}(s)$ , Na changes from a neutral atom to  $\text{Na}^+$  by transferring an electron to Cl which becomes  $\text{Cl}^-$ .

(b) **Metal + acid  $\rightarrow$  metal salt + H<sub>2</sub>**

In the reaction  $\text{Zn}(s) + 2\text{HCl}(aq) \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g)$ , Zn changes from a neutral atom to  $\text{Zn}^{2+}$  by transferring two electrons to two  $\text{H}^+$  ions which become two H atoms and then couple to form a  $\text{H}_2$  molecule.

(c) **Metal A + metal B salt  $\rightarrow$  metal B + metal A salt**

In the reaction  $\text{Zn}(s) + \text{CuSO}_4(aq) \rightarrow \text{Cu}(s) + \text{ZnSO}_4(aq)$ , Zn changes from a neutral atom to  $\text{Zn}^{2+}$  by transferring two electrons to  $\text{Cu}^{2+}$  which then becomes a neutral Cu atom.

(d) **Halogen A + halide B  $\rightarrow$  halogen B + halide A**

In the reaction  $\text{Cl}_2(aq) + 2\text{NaI}(aq) \rightarrow \text{I}_2(aq) + 2\text{NaCl}(aq)$ , the two  $\text{I}^-$  ions change to two I atoms (which couple to form the  $\text{I}_2$  molecule) by transferring two electrons to the two Cl atoms (in the  $\text{Cl}_2$  molecule) which then become two  $\text{Cl}^-$  ions.

Note that there is no electron transfer in the double replacement reactions; in precipitation reactions, two ions form an insoluble ionic compound, and in acid-base reactions, a  $\text{H}^+$  ion joins with a  $\text{OH}^-$  ion to form a molecular compound, water.

### ACTIVITY SERIES OF THE METALS

Li K Ba Ca Na Mg Al Zn Fe Cd Ni Sn Pb (H) Cu Hg Ag Au  
 Most reactive Least reactive

### SOLUBILITY RULES

ION	RULES
Group IA $\text{NH}_4^+$	1. All ionic compounds in which the cation is a Group IA element or $\text{NH}_4^+$ are <i>soluble</i> .
$\text{NO}_3^-$	2. All nitrates are <i>soluble</i> .
$\text{Cl}^-$ , $\text{Br}^-$ , $\text{I}^-$	3. All chlorides, bromides and iodides are <i>soluble</i> , except those of $\text{Ag}^+$ and $\text{Pb}^{2+}$ .
$\text{SO}_4^{2-}$	4. All sulphates are <i>soluble</i> except for $\text{CaSO}_4$ , $\text{SrSO}_4$ , $\text{BaSO}_4$ , $\text{PbSO}_4$ and $\text{AgSO}_4$ .
$\text{S}^{2-}$	5. All sulphides are <i>insoluble</i> except those of the Group IA or IIA elements or $\text{NH}_4^+$ .
	6. All other compounds are <i>insoluble</i> .