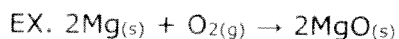


## Synthesis (composition/combination):

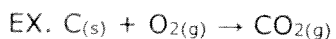
- two or more elements or compounds may combine to form a more complex compound.
- **Basic form:  $A + X \rightarrow AX$**

### Examples of synthesis reactions:

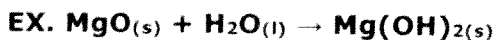
1. Metal + oxygen  $\rightarrow$  metal oxide



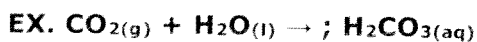
2. Nonmetal + oxygen  $\rightarrow$  nonmetallic oxide



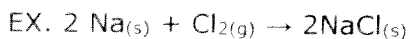
3. **Metal oxide + water  $\rightarrow$  metallic hydroxide**



4. **Nonmetallic oxide + water  $\rightarrow$  acid**



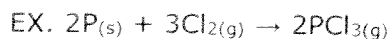
5. Metal + nonmetal  $\rightarrow$  salt



6. **Metal oxide + nonmetal oxide  $\rightarrow$  salt**

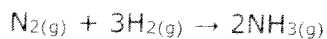


7. A few nonmetals combine with each other. (try to think of some you've seen!)



### This reaction must be remembered:

1. Nitrogen and hydrogen gases combine to form ammonia

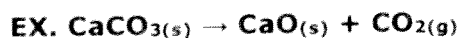


## B. Decomposition:

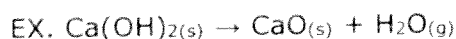
- A single compound breaks down into its component parts or simpler compounds.
- **Basic form:  $AX \rightarrow A + X$**

### Examples of decomposition reactions:

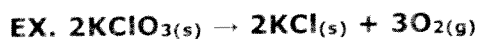
1. **Metallic carbonates, when heated, form metallic oxides and  $CO_2(g)$ .**



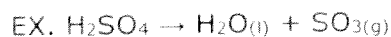
2. Most metallic hydroxides, when heated, decompose into metallic oxides and water.



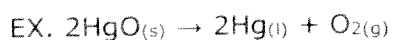
3. **Metallic chlorates, when heated, decompose into metallic chlorides and oxygen.**



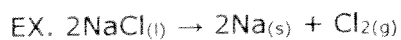
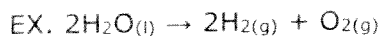
4. Some acids, when heated, decompose into nonmetallic oxides and water.



5. Some oxides, when heated, decompose.



6. Some decomposition reactions are produced by electricity.



### **This reaction must be remembered:**

**Hydrogen peroxide decomposes into water and oxygen gas**

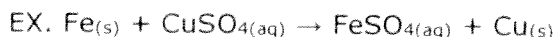


## C. Single Replacement (single displacement):

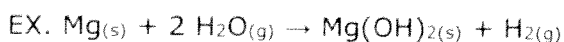
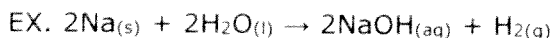
- a more active element takes the place of another element in a compound and sets the less active one free.
- Basic form:  $A + BX \rightarrow AX + B$  or  $AX + Y \rightarrow AY + X$**

### Examples of replacement reactions:

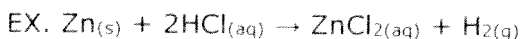
- Replacement of a metal in a compound by a more active metal.



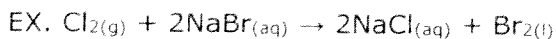
- Replacement of hydrogen in water by an active metal.



- Replacement of hydrogen in acids by active metals.




- Replacement of nonmetals by more active nonmetals.



**NOTE:** Refer to the **activity series for metals** and nonmetals (halogens only – fluorine most active, iodine least active – use the periodic table!) to predict products of replacement reactions. If the free element is above the element to be replaced in the compound, then the reaction will occur. If it is below, then no reaction occurs.

**TABLE 4.3 A Partial Activity Series of the Elements**

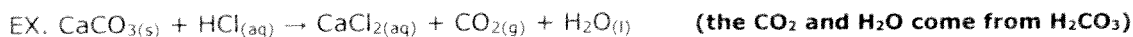
Oxidation Reaction		
Strongly reducing 	$\text{Li} \rightarrow \text{Li}^+ + e^-$	These elements react rapidly with aqueous $\text{H}^+$ ions (acid) or with liquid $\text{H}_2\text{O}$ to release $\text{H}_2$ gas.
	$\text{K} \rightarrow \text{K}^+ + e^-$	
	$\text{Ba} \rightarrow \text{Ba}^{2+} + 2e^-$	
	$\text{Ca} \rightarrow \text{Ca}^{2+} + 2e^-$	
	$\text{Na} \rightarrow \text{Na}^+ + e^-$	
Weakly reducing	$\text{Mg} \rightarrow \text{Mg}^{2+} + 2e^-$	These elements react with aqueous $\text{H}^+$ ions or with steam to release $\text{H}_2$ gas.
	$\text{Al} \rightarrow \text{Al}^{3+} + 3e^-$	
	$\text{Mn} \rightarrow \text{Mn}^{2+} + 2e^-$	
	$\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$	
	$\text{Cr} \rightarrow \text{Cr}^{3+} + 3e^-$	
	$\text{Fe} \rightarrow \text{Fe}^{2+} + 2e^-$	
Weakly reducing	$\text{Co} \rightarrow \text{Co}^{2+} + 2e^-$	These elements react with aqueous $\text{H}^+$ ions to release $\text{H}_2$ gas.
	$\text{Ni} \rightarrow \text{Ni}^{2+} + 2e^-$	
	$\text{Sn} \rightarrow \text{Sn}^{2+} + 2e^-$	
Weakly reducing	$\text{H}_2 \rightarrow 2\text{H}^+ + 2e^-$	These elements do not react with aqueous $\text{H}^+$ ions to release $\text{H}_2$ .
	$\text{Cu} \rightarrow \text{Cu}^{2+} + 2e^-$	
	$\text{Ag} \rightarrow \text{Ag}^+ + e^-$	
	$\text{Hg} \rightarrow \text{Hg}^{2+} + 2e^-$	
	$\text{Pt} \rightarrow \text{Pt}^{2+} + 2e^-$	
Weakly reducing	$\text{Au} \rightarrow \text{Au}^{3+} + 3e^-$	

## D. Double Replacement (double displacement):

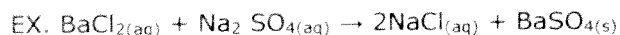
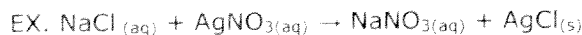
- Occurs between ions in aqueous solution. A reaction will occur when a pair of ions come together to **produce at least one** of the following:
  - A product that decomposes - remember these:
    - $\text{H}_2\text{CO}_3$  formed decomposes into  $\text{H}_2\text{O}$  and  $\text{CO}_2(\text{g})$
    - $\text{H}_2\text{SO}_3$  formed decomposes into  $\text{H}_2\text{O}$  and  $\text{SO}_2(\text{g})$
    - $\text{NH}_4\text{OH}$  formed decomposes into  $\text{H}_2\text{O}$  and  $\text{NH}_3(\text{g})$
  - a precipitate (**insoluble** solid - see solubility rules below)
  - a gas like  $\text{H}_2\text{S}$
  - water or some other non-ionized substance (like a weak acid - see list of strong acids below).
- Basic form:  $\text{AX} + \text{BY} \rightarrow \text{AY} + \text{BX}$**

### Examples of double replacement reactions:

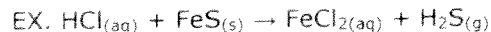
- Formation of a product that decomposes.



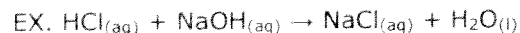
- Formation of precipitate.



- Formation of a gas.



- Formation of water. (If the reaction is between an acid and a base it is called a neutralization reaction.)



## Solubility Rules

The following are the solubility rules for common ionic solids. If there two rules appear to contradict each other, the preceding rule takes precedence.

- Salts containing **Group I elements** are **soluble**. There are few exceptions to this rule. Salts containing the **ammonium ion** ( $\text{NH}_4^+$ ) are also **soluble**.
- Salts containing **nitrate ion** ( $\text{NO}_3^-$ ) are generally **soluble**. Other anions that make **soluble** salts are **acetate** ( $\text{C}_2\text{H}_3\text{O}_2^-$ ), **chlorate**, ( $\text{ClO}_3^-$ ), and **perchlorate** ( $\text{ClO}_4^-$ ).
- Salts containing  $\text{Cl}^-$ ,  $\text{Br}^-$ , or  $\text{I}^-$  are **generally soluble**. Important **exceptions** to this rule are halide salts of  **$\text{Ag}^+$ ,  $\text{Pb}^{2+}$ , and  $(\text{Hg}_2)^{2+}$** . Thus,  $\text{AgCl}$ ,  $\text{PbBr}_2$ , and  $\text{Hg}_2\text{Cl}_2$  are insoluble.
- Most **sulfate** salts are **soluble**. Important **exceptions** to this rule include  **$\text{Ca}^{2+}$ ,  $\text{Sr}^{2+}$ ,  $\text{Ba}^{2+}$ ,  $\text{Pb}^{2+}$ , and  $\text{Hg}_2^{2+}$**  which are all insoluble.
- Most hydroxide salts are only slightly soluble. **Hydroxide** salts of **Group I elements** are **soluble**. Hydroxide salts of **Group II elements** ( **$\text{Ca}^{2+}$ ,  $\text{Sr}^{2+}$ , and  $\text{Ba}^{2+}$** ) are **soluble**. Others are considered insoluble or only slightly soluble.
- Consider that all other salts are insoluble if not "covered" by a previous solubility rule.

### Strong acids

$\text{HCl}$ ,  $\text{HBr}$ ,  $\text{HI}$ ,  $\text{HNO}_3$ ,  
 $\text{HClO}_4$ ,  $\text{H}_2\text{SO}_4$

If produced WILL NOT cause a DR replacement rxn to occur - but all other acids are WEAK and will cause a reaction if produced.

## Combustion of Hydrocarbons:

Another important type of reaction, in addition to the four types above, is that of the combustion of a hydrocarbon. When a hydrocarbon is burned with sufficient oxygen supply, the products are always carbon dioxide and water vapor. If the supply of oxygen is low or restricted, then carbon monoxide will be produced. This is why it is so dangerous to have an automobile engine running inside a closed garage or to use a charcoal grill indoors.

- Hydrocarbon ( $C_xH_y$ ) +  $O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$
- EX.  $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$
- EX.  $2C_4H_{10(g)} + 13O_{2(g)} \rightarrow 8CO_{2(g)} + 10H_2O_{(g)}$

### NOTE:

- **Complete combustion** means the higher oxidation number ( $CO_2$ ) is attained.
- **Incomplete combustion** means the lower oxidation number ( $CO$ ) is attained.
- **The phrase "To burn"** means to add oxygen unless told otherwise.

## Some additional notes:

Know your diatomic free elements: hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine

Ionic compounds NOT dissolved in water are in the solid state

Make sure FORMULAS are correct FIRST, then balance the equation with coefficients!!!!

## The Activity Series of the Elements

Whether a reaction occurs between a given ion and a given element depends on the relative ease with which the various species gain or lose electrons (are oxidized or reduced). An **activity series** ranks the elements in order of their reducing ability in aqueous solution.

TABLE 4.3 A Partial Activity Series of the Elements

SINGLE REPLACEMENT

Oxidation Reaction		
Strongly reducing ↑	Li → Li <sup>+</sup> + e <sup>-</sup>	These elements react rapidly with aqueous H <sup>+</sup> ions (acid) or with liquid H <sub>2</sub> O to release H <sub>2</sub> gas.
	K → K <sup>+</sup> + e <sup>-</sup>	
	Ba → Ba <sup>2+</sup> + 2 e <sup>-</sup>	
	Ca → Ca <sup>2+</sup> + 2 e <sup>-</sup>	
	Na → Na <sup>+</sup> + e <sup>-</sup>	
	Mg → Mg <sup>2+</sup> + 2 e <sup>-</sup>	These elements react with aqueous H <sup>+</sup> ions or with steam to release H <sub>2</sub> gas.
	Al → Al <sup>3+</sup> + 3 e <sup>-</sup>	
	Mn → Mn <sup>2+</sup> + 2 e <sup>-</sup>	
	Zn → Zn <sup>2+</sup> + 2 e <sup>-</sup>	
	Cr → Cr <sup>3+</sup> + 3 e <sup>-</sup>	
	Fe → Fe <sup>2+</sup> + 2 e <sup>-</sup>	
	Co → Co <sup>2+</sup> + 2 e <sup>-</sup>	These elements react with aqueous H <sup>+</sup> ions to release H <sub>2</sub> gas.
	Ni → Ni <sup>2+</sup> + 2 e <sup>-</sup>	
	Sn → Sn <sup>2+</sup> + 2 e <sup>-</sup>	
	H <sub>2</sub> → 2 H <sup>+</sup> + 2 e <sup>-</sup>	
Weakly reducing ↓	Cu → Cu <sup>2+</sup> + 2 e <sup>-</sup>	These elements do not react with aqueous H <sup>+</sup> ions to release H <sub>2</sub> .
	Ag → Ag <sup>+</sup> + e <sup>-</sup>	
	Hg → Hg <sup>2+</sup> + 2 e <sup>-</sup>	
	Pt → Pt <sup>2+</sup> + 2 e <sup>-</sup>	
	Au → Au <sup>3+</sup> + 3 e <sup>-</sup>	

## Solubility Rules

The following are the solubility rules for common ionic solids. If there two rules appear to contradict each other, the preceding rule takes precedence.

DOUBLE REPLACEMENT

- Salts containing **Group I elements** are **soluble**. There are few exceptions to this rule. Salts containing the **ammonium ion** (NH<sub>4</sub><sup>+</sup>) are also **soluble**.
- Salts containing **nitrate ion** (NO<sub>3</sub><sup>-</sup>) are generally **soluble**. Other anions that make **soluble** salts are **acetate** (C<sub>2</sub>H<sub>3</sub>O<sub>2</sub><sup>-</sup>), **chlorate**, (ClO<sub>3</sub><sup>-</sup>), and **perchlorate** (ClO<sub>4</sub><sup>-</sup>).
- Salts containing Cl<sup>-</sup>, Br<sup>-</sup>, or I<sup>-</sup> are **generally soluble**. Important **exceptions** to this rule are halide salts of **Ag<sup>+</sup>, Pb<sup>2+</sup>, and (Hg<sub>2</sub>)<sup>2+</sup>**. Thus, AgCl, PbBr<sub>2</sub>, and Hg<sub>2</sub>Cl<sub>2</sub> are insoluble.
- Most **sulfate** salts are **soluble**. Important **exceptions** to this rule include **Ca<sup>2+</sup>, Sr<sup>2+</sup>, Ba<sup>2+</sup>, Pb<sup>2+</sup>, and Hg<sub>2</sub><sup>2+</sup>** which are all insoluble.
- Most hydroxide salts are only slightly soluble. **Hydroxide** salts of **Group I elements** are **soluble**. Hydroxide salts of **Group II elements** (**Ca<sup>2+</sup>, Sr<sup>2+</sup>, and Ba<sup>2+</sup>**) are **soluble**. Others are considered insoluble or only slightly soluble.
- Consider that all other salts are insoluble if not "covered" by a previous solubility rule.