Synthesis (composition/combination):

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

Examples of synthesis reactions:

1. Metal + oxygen → metal oxide

   EX. $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

2. Nonmetal + oxygen → nonmetallic oxide

   EX. $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

3. **Metal oxide + water → metallic hydroxide**

   EX. $MgO_{(s)} + H_2O_{(l)} \rightarrow Mg(OH)_{2(s)}$

4. **Nonmetallic oxide + water → acid**

   EX. $CO_{2(g)} + H_2O_{(l)} \rightarrow$; $H_2CO_{3(aq)}$

5. Metal + nonmetal → salt

   EX. $2\ Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$

6. **Metal oxide + nonmetal oxide → salt**

   EX: $CaO_{(s)} + SO_{3(g)} \rightarrow CaSO_{4(s)}$

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

   EX. $2P_{(s)} + 3Cl_{2(g)} \rightarrow 2PCl_{3(g)}$

This reaction must be remembered:

1. Nitrogen and hydrogen gases combine to form ammonia

   $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$

B. Decomposition:

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

Examples of decomposition reactions:

1. **Metallic carbonates, when heated, form metallic oxides and $CO_{2(g)}$.**

   EX. $CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$

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

   EX. $Ca(OH)_{2(s)} \rightarrow CaO_{(s)} + H_2O_{(g)}$

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

   EX. $2KClO_{3(s)} \rightarrow 2KCl_{(s)} + 3O_{2(g)}$

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

   EX. $H_2SO_4 \rightarrow H_2O_{(l)} + SO_{3(g)}$

5. Some oxides, when heated, decompose.

   EX. $2HgO_{(s)} \rightarrow 2Hg_{(l)} + O_{2(g)}$

6. Some decomposition reactions are produced by electricity.

   EX. $2H_2O_{(l)} \rightarrow 2H_{2(g)} + O_{2(g)}$

   EX. $2NaCl_{(l)} \rightarrow 2Na_{(s)} + Cl_{2(g)}$

This reaction must be remembered:

Hydrogen peroxide decomposes into water and oxygen gas

$2\ H_2O_{2(aq)} \rightarrow 2\ H_2O_{(l)} + O_{2(g)}$

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 → AX + B or AX + Y → AY + X**

Examples of replacement reactions:

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

   EX. $Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$

2. Replacement of hydrogen in water by an active metal.

   EX. $2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$

   EX. $Mg_{(s)} + 2\ H_2O_{(g)} \rightarrow Mg(OH)_{2(s)} + H_{2(g)}$

3. Replacement of hydrogen in acids by active metals.

   EX. $Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$

4. Replacement of nonmetals by more active nonmetals.

   EX. $Cl_{2(g)} + 2NaBr_{(aq)} \rightarrow 2NaCl_{(aq)} + Br_{2(l)}$

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	$Li \rightarrow Li^{+} + e^{-}$	These elements react rapidly with aqueous H^{+} ions (acid) or with liquid H_2O to release H_2 gas.
	$K \rightarrow K^{+} + e^{-}$	
	$Ba \rightarrow Ba^{2+} + 2e^{-}$	
	$Ca \rightarrow Ca^{2+} + 2e^{-}$	
	$Na \rightarrow Na^{+} + e^{-}$	
	$Mg \rightarrow Mg^{2+} + 2e^{-}$	These elements react with aqueous H^{+} ions or with steam to release H_2 gas.
	$Al \rightarrow Al^{3+} + 3e^{-}$	
	$Mn \rightarrow Mn^{2+} + 2e^{-}$	
	$Zn \rightarrow Zn^{2+} + 2e^{-}$	
	$Cr \rightarrow Cr^{3+} + 3e^{-}$	
	$Fe \rightarrow Fe^{2+} + 2e^{-}$	
	$Co \rightarrow Co^{2+} + 2e^{-}$	These elements react with aqueous H^{+} ions to release H_2 gas.
	$Ni \rightarrow Ni^{2+} + 2e^{-}$	
	$Sn \rightarrow Sn^{2+} + 2e^{-}$	
	$H_2 \rightarrow 2H^{+} + 2e^{-}$	
	$Cu \rightarrow Cu^{2+} + 2e^{-}$	These elements do not react with aqueous H^{+} ions to release H_2.
	$Ag \rightarrow Ag^{+} + e^{-}$	
	$Hg \rightarrow Hg^{2+} + 2e^{-}$	
	$Pt \rightarrow Pt^{2+} + 2e^{-}$	
Weakly reducing	$Au \rightarrow Au^{3+} + 3e^{-}$	

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:
    1. A product that decomposes - remember these:
        - H_2CO_3 formed decomposes into H_2O and $CO_{2(g)}$
        - H_2SO_3 formed decomposes into H_2O and $SO_{2(g)}$
        - NH_4OH formed decomposes into H_2O and $NH_{3(g)}$
    2. a precipitate (**insoluble** solid – see solubility rules below)
    3. a gas like H_2S
    4. water or some other non-ionized substance (like a weak acid – see list of strong acids below).
- **Basic form: AX + BY → AY + BX**

Examples of double replacement reactions:

1. Formation of a product that decomposes.

   EX. $CaCO_{3(s)} + HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$ **(the CO_2 and H_2O come from H_2CO_3)**

2. Formation of precipitate.

   EX. $NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow NaNO_{3(aq)} + AgCl_{(s)}$

   EX. $BaCl_{2(aq)} + Na_2SO_{4(aq)} \rightarrow 2NaCl_{(aq)} + BaSO_{4(s)}$

3. Formation of a gas.

   EX. $HCl_{(aq)} + FeS_{(s)} \rightarrow FeCl_{2(aq)} + H_2S_{(g)}$

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

   EX. $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

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.

1. Salts containing **Group I elements** are **soluble**. There are few exceptions to this rule. Salts containing the **ammonium ion** (NH_4^+) are also **soluble**.
2. Salts containing **nitrate ion** (NO_3^-) are generally **soluble**. Other anions that make **soluble** salts are **acetate** ($C_2H_3O_2^-$), **chlorate**, (ClO_3^-), and **perchlorate** (ClO_4^-).
3. Salts containing Cl^-, Br^-, or I^- are **generally soluble**. Important **exceptions** to this rule are halide salts of **Ag^+, Pb^{2+}, and $(Hg_2)^{2+}$**. Thus, $AgCl$, $PbBr_2$, and Hg_2Cl_2 are insoluble.
4. Most **sulfate** salts are **soluble**. Important **exceptions** to this rule include **Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, and Hg_2^{2+}** which are all insoluble.
5. Most hydroxide salts are only slightly soluble. **Hydroxide** salts of **Group I** elements are **soluble**. Hydroxide salts of **Group II** elements (**Ca^{2+}, Sr^{2+}, and Ba^{2+}**) are **soluble**. Others are considered insoluble or only slightly soluble.
6. Consider that all other salts are insoluble if not "covered" by a previous solubility rule.

Strong acids
HCl, HBr, HI, HNO_3, $HClO_4$, H_2SO_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.

SINGLE REPLACEMENT

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	$H_2 \rightarrow 2H^{+} + 2e^{-}$	
	$Cu \rightarrow Cu^{2+} + 2e^{-}$	These elements do not react with aqueous H^{+} ions to release H_2.
	$Ag \rightarrow Ag^{+} + e^{-}$	
	$Hg \rightarrow Hg^{2+} + 2e^{-}$	
Weakly	$Pt \rightarrow Pt^{2+} + 2e^{-}$	
reducing	$Au \rightarrow Au^{3+} + 3e^{-}$	

DOUBLE REPLACEMENT

Solubility Rules

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3. Salts containing Cl^{-}, Br^{-}, or I^{-} are **generally soluble**. Important **exceptions** to this rule are halide salts of **Ag^{+}, Pb^{2+}, and $(Hg_2)^{2+}$**. Thus, $AgCl$, $PbBr_2$, and Hg_2Cl_2 are insoluble.
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5. Most hydroxide salts are only slightly soluble. **Hydroxide** salts of **Group I** elements are **soluble**. Hydroxide salts of **Group II** elements (**Ca^{2+}, Sr^{2+}, and Ba^{2+}**) are **soluble**. Others are considered insoluble or only slightly soluble.
6. Consider that all other salts are insoluble if not "covered" by a previous solubility rule.