UNIT 6
SYMBOLS AND BALANCING CHEMICAL EQUATIONS

Chemical Formulas show chemistry at a standstill. Chemical Equations show chemistry in action.

A. Equations show:
   1. the reactants which enter into a reaction.
   2. the products which are formed by the reaction.
   3. the amounts of each substance used and each substance produced.

B. Two important principles to remember:
   1. Every chemical compound has a formula which cannot be altered.
   2. A chemical reaction must account for every atom that is used. This is an application of the Law of Conservation of Matter (or Mass) which states that in a chemical reaction atoms are neither created nor destroyed.
      - # of atoms in reactants = # of atoms in products
      - Total mass of reactants = Total mass of products

C. Some things to remember about writing equations:
   1. The diatomic elements when they stand alone are always written H₂, N₂, O₂, F₂, Cl₂, Br₂, I₂
   2. The sign, \( \rightarrow \), means "yields" and shows the direction of the action.
   3. A small delta, (\( \Delta \)), above the arrow shows that heat has been added.
   4. A double arrow, \( \leftrightarrow \), shows that the reaction is reversible and can go in both directions. (Don’t worry about that now!)
   5. There are symbols to identify the phase of each substance in the reaction (liquid, solid, …)
   6. Before beginning to balance an equation, check each formula to see that it is correct.
      NEVER change a formula during the balancing of an equation.
   7. Balancing is done by placing coefficients in front of the formulas to insure the same number of atoms of each element on both sides of the arrow.
      - All elements in the formula are multiplied by the value of the coefficient

\[
\text{Reactants} \quad \rightarrow \quad \text{Products}
\]

**Symbols**
- (s) solid
- (g) gas
- (aq) aqueous
- (l) liquid
- (v) vapor
- (ppt) a solid precipitate
- \( \Delta \) heat is added or changes

Reactants – starting materials
Products – chemicals produced as a result of the reaction

- Heat can function as a reactant or product. Often it is the Heat Energy that we seek
  1. Power Plants harness the HEAT produced by the combustion of coal or nuclear fission to generate electricity
**Steps used to Balance Equations:**

Write the formula for each reactant and product.

Using the chemical formulas, write a skeleton equation for the reaction. Separate each reactant with a + sign and separate each product with a + sign.

**PUT A BLANK _____ IN FRONT OF EACH CHEMICAL IN THE REACTION**

Count the atoms of each element in the reactants.

Count the atoms of each element in the products.

Use **coefficients** to multiply the chemical to make sure the number of atoms of each element are equal on each side of the equation.

This step is a trial and error process.

**HINT - BALANCE SINGLE ELEMENTS, DIATOMICS, AND / OR HYDROGEN LAST.**

7. Check your work to make sure the numbers of each atom are equal. Double check to see that the COEFFICIENTS are shown in the lowest ratio

**Ex:** Hydrogen gas reacts with oxygen gas to produce water vapor.

(write the formulas based on the rules for nomenclature – ionic, covalent, diatomics, acids, common names for elements)

Step 1. Write Skeleton Equation: _____H₂ (g) + _____O₂(g) → _____H₂O (v)

Step 2,3,4. Use coefficients to balance: 2H₂ (g) + _____O₂(g) → 2H₂O (v)

Step 5. Reduce coefficients, if necessary.

Step 6. Check your work.

**Unit 6 Worksheets – Balancing WS 1, 2, 3, 4**

**Categorization of Reactions Activity**
I. **Synthesis** – two or more elements or compounds combine to form a single compound

Common synthesis reactions occur when substances **burn, oxidize, tarnish, rust**.

**Synthesis**: see Reference Table

a. Formation of binary compound: \( A + B \rightarrow AB \)

\[ 2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O} \]

b. Metal oxide-water reactions: \( \text{MO} + \text{H}_2\text{O} \rightarrow \text{base (with OH}^{-}) \)

\[ \text{Na}_2\text{O} + \text{H}_2\text{O} \rightarrow 2\text{NaOH} \]

c. Nonmetal oxide-water reactions: \( (\text{NM})\text{O} + \text{H}_2\text{O} \rightarrow \text{acid (with H}^{+}) \)

\[ \text{N}_2\text{O}_5 + \text{H}_2\text{O} \rightarrow \text{HNO}_3 \quad \text{Nitric Acid} \]

\[ \text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_4 \quad \text{Sulfuric Acid} \]

\[ \text{P}_2\text{O}_5 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{PO}_4 \quad \text{Phosphoric Acid} \]

**Practice synthesis**: Write the product and balance –

1. magnesium burns in oxygen \( \rightarrow \)

2. calcium oxide added to water \( \rightarrow \)
II. Decomposition - A single compound breaks down into its component parts or simpler compounds. Common decomposition reactions occur when heated or when they undergo electrolysis (breakdown via electricity)

**Decomposition: see Reference Table**

a. Binary compounds: $\text{AB} \rightarrow \text{A} + \text{B}$
   
   $2\text{HgO}(s) \rightarrow 2\text{Hg}(l) + \text{O}_2(g)$

b. Metallic carbonates: $\text{MCO}_3 \rightarrow \text{MO} + \text{CO}_2$
   
   $\text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g)$

c. Metallic hydrogen carbonates: $\text{MHCO}_3 \rightarrow \text{MO} + \text{H}_2\text{O}(l) + \text{CO}_2(g)$
   
   $\text{NaHCO}_3 \rightarrow \text{Na}_2\text{O} + \text{H}_2\text{O} + \text{CO}_2$

d. Metallic hydroxides: $\text{MOH} \rightarrow \text{MO} + \text{H}_2\text{O}$
   
   $\text{Ca(OH)}_2(s) \rightarrow \text{CaO}(s) + \text{H}_2\text{O}(g)$

e. Metallic chlorates: $\text{MClO}_3 \rightarrow \text{MCl} + \text{O}_2$
   
   $2\text{KClO}_3(s) \rightarrow 2\text{KCl}(s) + 3\text{O}_2(g)$

f. Oxyacids decompose to nonmetal oxides and water: $\text{acid} \rightarrow (\text{NM})\text{O} + \text{H}_2\text{O}$
   
   $\text{H}_2\text{SO}_4 \rightarrow \text{H}_2\text{O}(l) + \text{SO}_3(g)$

**Practice: Decomposition**

- sodium carbonate (heated) yields ?
- lithium chlorate (heated) yields ?
- electrolysis of aluminum oxide results in ?
- sulfuric acid heated gently produces ?
III. Single Replacement:

- **Use Activity Series on NC Reference Tables**
- A more active element takes the place of another element in a compound and sets the less active one free.
- **Basic form:** \( AX + BX \rightarrow AX + B \) or \( AX + Y \rightarrow AY + X \)

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

\[
\text{EX. } Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)
\]

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

\[
\begin{align*}
\text{EX. } 2Na(s) + 2H_2O(l) & \rightarrow 2NaOH(aq) + H_2(g) \\
\text{EX. } Mg(s) + H_2O(g) & \rightarrow MgO(s) + H_2(g)
\end{align*}
\]

c. Replacement of hydrogen in acids by active metals.

\[
\text{EX. } Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)
\]

d. Replacement of nonmetals by more active nonmetals.

\[
\text{EX. } Cl_2(g) + 2NaBr(aq) \rightarrow 2NaCl(aq) + Br_2(l)
\]

**Practice – single replacement**

1. iron filings added to copper(II) sulfate in solution

2. aluminum in hydrochloric acid

3. potassium metal added to cold water

4. zinc metal added to mercury(II) nitrate

5. silver metal added to copper(II) sulfate

6. chlorine gas bubbled through a solution of calcium bromide
IV. **Double Replacement** - \( AB + CD \rightarrow AD + CB \): The cations and anions “trade partners” and new compounds are formed. When writing the formulas for the products, you MUST obey the rules for ionic or covalent compounds

**MUST** have one of the following occur
- Formation of a precipitate from solution
- Formation of a gas
- Formation of water as in an Acid-Base neutralization reaction

1. **Formation of precipitate** – **See Reference Tables**

   EX. \( \text{NaCl}_{(aq)} + \text{AgNO}_3_{(aq)} \rightarrow \text{NaNO}_3_{(aq)} + \text{AgCl}_{(s)} \)

   EX. \( \text{BaCl}_2_{(aq)} + \text{Na}_2\text{SO}_4_{(aq)} \rightarrow 2\text{NaCl}_{(aq)} + \text{BaSO}_4_{(s)} \)

2. **Formation of a gas.**

   EX. \( \text{HCl}_{(aq)} + \text{FeS}_{(s)} \rightarrow \text{FeCl}_2_{(aq)} + \text{H}_2\text{S}_{(g)} \)

   List of Common Gases that could be produced
   \( \text{F}_2, \text{Cl}_2, \text{H}_2, \text{N}_2, \text{O}_2, \text{SO}_2, \text{SO}_3, \text{CO}, \text{CO}_2, \text{H}_2\text{S}, \text{NO}, \text{NO}_2, \text{NH}_3 \)

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

   EX. \( \text{HCl}_{(aq)} + \text{NaOH}_{(aq)} \rightarrow \text{NaCl}_{(aq)} + \text{H}_2\text{O}_{(l)} \)

V. **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** – has carbon, hydrogen in the compound

  General Equation: \( (\text{C}_x\text{H}_y) + \text{O}_2_{(g)} \rightarrow \text{CO}_2_{(g)} + \text{H}_2\text{O}_{(g)} \)

  EX. \( \text{CH}_4_{(g)} + 2\text{O}_2_{(g)} \rightarrow \text{CO}_2_{(g)} + 2\text{H}_2\text{O}_{(g)} \)

  EX. \( 2\text{C}_4\text{H}_{10}_{(g)} + 13\text{O}_2_{(g)} \rightarrow 8\text{CO}_2_{(g)} + 10\text{H}_2\text{O}_{(g)} \)