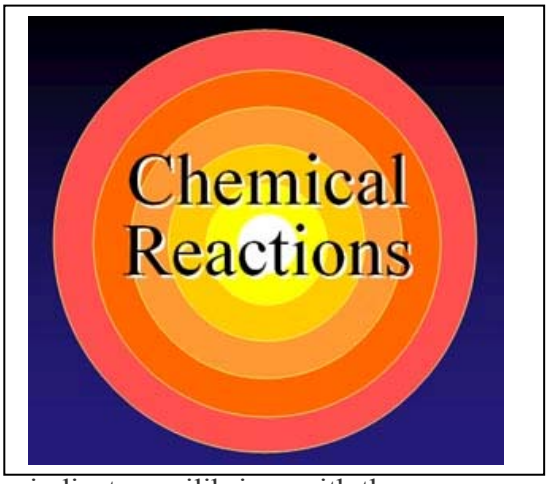

Unit 7
Chemical Equations and Reactions

What is a Chemical Equation?

A **Chemical Equation** is a written representation of the process that occurs in a chemical reaction. A chemical equation is written with the **Reactants** on the left side of an arrow and the **Products** of the chemical reaction on the right side of the equation. The head of the arrow typically points toward the right or toward the product side of the equation, although reactions may indicate equilibrium with the reaction proceeding in **both** directions simultaneously.

The elements in an equation are denoted using their symbols. **Coefficients** next to the symbols indicate the **stoichiometric** numbers. Subscripts are used to indicate the number of atoms of an element present in a chemical species.

An example of a chemical equation may be seen in the combustion of methane:

$$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$$

Balancing Equations Notes

An equation for a chemical reaction in which the number of atoms for each element in the reaction and the total charge are the same for both the reactants and the products. In other words, the mass and the charge are balanced on both sides of the reaction.

Symbol	Meaning
+	used to separate one reactant or product from another
	used to separate the reactants from the products - it is pronounced "yields" or "produces" when the equation is read
	used when the reaction can proceed in both directions - this is called an equilibrium arrow and will be used later in the course
(g)	indicates that the substance is in a gaseous state
	an alternative way of representing a substance in a gaseous state
(s)	indicates that the substance is in a solid state
	an alternative way of representing a substance in a solid state
(aq)	indicates that the substance is dissolved in water - the aq comes from aqueous
(ℓ)	Identifies a phase state as pure liquid
	indicates that heat is applied to make the reaction proceed

LAW OF CONSERVATION OF MASS

In all chemical equations the **LAW OF CONSERVATION OF MASS** must be met. Matter can NOT be created nor destroyed in a chemical reaction.

Remember, in a chemical reaction, the atoms/ions are simply **rearranged** to form new substances.

Therefore, chemical equations **MUST be balanced**.

WHAT IS A "BALANCED" CHEMICAL EQUATION?
A balanced chemical equation is one in which **each side** of the equation has the **same number** of atoms/ions of each element.

Example: $Al\ (s) + O_2\ (g) \rightarrow Al_2O_3\ (s)$

Not Balanced

1 Al	2 Al
2 O	3 O

$\mathbf{4}\ Al\ (s) + \mathbf{3}\ O_2\ (g) \rightarrow \mathbf{2}\ Al_2O_3\ (s)$

Balanced

4 Al	4 Al
6 O	6 O

RULES FOR BALANCING CHEMICAL EQUATIONS

1. Write the **correct** chemical formulas for **all** of the reactants and the products.

2. Write the formulas of the **reactants** on the **LEFT** of the reaction arrow; write the formulas of the **products** on the **RIGHT** of the reaction arrow.

3. COUNT the total number of atoms/ions of each element in the reactants and the total number of atoms/ions of each element in the products.
   ** A **polyatomic ion** that appears ***unchanged*** on both sides of the equation is counted as a single unit.

4. Balance the elements one at a time using **coefficients**.
   - A coefficient is a small WHOLE number that is written ***in front*** of a chemical formula in a chemical equation.

- When **no** coefficient is written, the coefficient is assumed to be **1**.
- It is best to begin with elements **OTHER THAN** hydrogen and oxygen. These elements often occur more than twice in equations.
- ** You must **NOT** attempt to balance the equation by changing subscripts in chemical formulas!!!!!!

5. Check each atom/ion, or polyatomic ion to be sure that the equation is **correctly balanced**.

6. Finally, make sure that all of the coefficients are in the **LOWEST** possible whole number ratios. (At least one of the coefficients must be a prime number!)

Use coefficients to make sure the number of atoms is the same on both sides of the equation.

1. _2_ H_2 + ___ O_2 → _2_ H_2O

2. _2_ HCl + ___ Zn →___ $ZnCl_2$ + ___ H_2

3. _2_ Al + _3_ CaS → ___ Al_2S_3 + _3_ Ca

4. Write the skeleton equation for the reaction of **solid Iron and gaseous chlorine react to produce a solid iron (III) chloride**

Write unbalanced equation

$Fe\ (s) + Cl_2\ (g) \rightarrow FeCl_3\ (s)$

Write balanced equation

$2\ Fe\ (s) + 3\ Cl_2\ (g) \rightarrow 2\ FeCl_3\ (s)$

Diatomic Elements

Diatomic Elements are always diatomic (written with a subscribe of 2) when they are in their elemental form

1. Hydrogen	H_2	5. Chlorine	Cl_2
2. Nitrogen	N_2	6. Iodine	I_2
3. Oxygen	O_2	7. Bromine	Br_2
4. Fluorine	F_2		

Types of Chemical Reactions Notes

- **Synthesis**- **two or more** elements or compounds combine to form one compound.
- **Decomposition**- a **single** compound decomposes into two or more elements or smaller compounds.
- **Single Replacement**- a metal will **replace** a less active metal in an ionic compound OR a nonmetal will replace a less active nonmetal.
- **Double Replacement**- the metals in ionic compounds **switch** places.
- **Combustion**- an **organic** compound containing carbon, hydrogen and sometimes oxygen reacts with oxygen gas to form carbon dioxide and water.

Examples

$$Cu(s) + AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$$
Single Replacement

$$C_6H_{12}O_6(s) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$
Combustion

$$Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$$
Synthesis

$$CuCl_2(aq) + Pb(NO_3)_2(aq) \rightarrow Cu(NO_3)_2(aq) + PbCl_2(s)$$
Double Replacement

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$
Decomposition

- **Synthesis**: Definition - two or more substances react to form 1 product. Usually releases energy, ***EXOTHERMIC***. Combination reactions that contain oxygen as a reactant can also be considered combustion.

$$A + X \rightarrow AX$$
$$4\ Fe\ (s) + 3\ O_2\ (g) \rightarrow 2\ Fe_2O_3\ (s)$$
$$CaO\ (s) + H_2O\ (l) \rightarrow Ca(OH)_2\ (s)$$

One example of a synthesis reaction is the combination of iron and sulfur to form iron (II) sulfide:

$$8\ Fe + S_8 \text{ ---> } 8\ FeS$$

- **DECOMPOSITION**: Definition - A single compound breaks down into 2 or more elements or compounds

$$AX \rightarrow A + X$$
$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2\ (g)$$
$$2KClO_3\ (s) \rightarrow 2KCl\ (s) + 3O_2\ (g)$$

$$CaCO_3\ (s) \rightarrow CaO\ (s) + CO_2\ (g)$$

These reactions often require an energy source as an initiator. Energy sources can be heat, light, or electricity. They are usually ***ENDOTHERMIC, .
One example of a decomposition reaction is the electrolysis of water to make oxygen and hydrogen gas:

$$2H_2O \text{ ---> } 2\ H_2 + O_2$$

- **Combustion** : Definition - Oxygen gas combines with a substance and releases energy in the form of light or heat. So combustion reactions are usually *exothermic*. Combination reactions that contain oxygen as a reactant can also be considered combustion.

$$A + O_2 \rightarrow$$
$$C(s) + O_2(g) \rightarrow CO_2(g) + \text{energy}$$
$$4\ Fe\ (s) + 3O_2\ (g) \rightarrow 2Fe_2O_3\ (s) + \text{energy}$$

For hydrocarbons:

$$C_xH_y + [x + (y/4)]\ O_2 \rightarrow xCO_2 + (y/2)H_2O$$
$$C_3H_8(g) + O_2(g) \rightarrow CO_2\ (g) + H_2O\ (g) + light + heat$$

- **Single Replacement** : Definition - one ion replaces another in a compound.

$$AB + C \text{ --> } AC + B$$

One example of a single displacement reaction is when magnesium replaces hydrogen in water to make magnesium hydroxide and hydrogen gas:

$$Mg + 2\ H_2O \text{ ---> } Mg(OH)_2 + H_2$$

- **Double Replacement** : Definition - two ions replace each other or switch places in compounds.

$$AB + CD \text{ --> } AC + BD$$

One example of a double displacement reaction is the reaction of lead (II) nitrate with potassium iodide to form lead (II) iodide and potassium nitrate:

$$Pb(NO_3)_2 + 2\ KI \text{ ---> } PbI_2 + 2\ KNO_3$$

Sample Problems (the solutions are in the next section)

List the type of the following reactions.	Solutions
1) $NaOH + KNO_3 \rightarrow NaNO_3 + KOH$	1) double replacement
2) $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$	2) combustion
3) $2\ Fe + 6\ NaBr \rightarrow 2\ FeBr_3 + 6\ Na$	3) single replacement
4) $CaSO_4 + Mg(OH)_2 \rightarrow Ca(OH)_2 + MgSO_4$	4) double replacement
5) $NH_4OH + HBr \rightarrow H_2O + NH_4Br$	5) acid-base (double replacement)
6) $Pb + O_2 \rightarrow PbO_2$	6) synthesis
7) $Na_2CO_3 \rightarrow Na_2O + CO_2$	7) decomposition

Practice
Balancing Chemical Equations Worksheet

Balance the following chemical equations using coefficients

1 **1**$Al(OH)_3(s)$ + **3**$HCl\ (aq) \rightarrow$ **1**$AlCl_3\ (aq)$ + **3**$H_2O\ (l)$

2. **3**$Fe_2O_3\ (s)$ + **1**$CO\ (g) \rightarrow$ **2**$Fe_3O_4(s)$ + **1**$CO_2\ (g)$

3. **4**$FeO\ (s)$ + **1**$O_2\ (g) \rightarrow$ **2**$Fe_2O_3\ (s)$

4. **2**$C_6H_6\ (l)$ + **15**$O_2\ (g) \rightarrow$ **12**$CO_2\ (g)$ + **6**$H_2O\ (g)$

5. **3**$Ca(OH)_2\ (aq)$ + **2**$H_3PO_4\ (aq) \rightarrow$ **6**$H_2O\ (l)$ + **1**$Ca_3(PO_4)_2\ (s)$

6. **2**$I_4O_9\ (s) \rightarrow$ **1**$I_2O_6(s)$ + **3**$I_2\ (s)$ + **6**$O_2\ (g)$

there's another way to balance this equation…can you figure it out?

7. **2**$Eu\ (s)$ + **6**$HF\ (g) \rightarrow$ **2**$EuF_3\ (s)$ + **3**$H_2\ (g)$

8. **3**$NaHCO_3\ (aq)$ + **1**$C_6H_8O_7\ (aq) \rightarrow$ **3**$CO_2\ (g)$ + **3**$H_2O\ (l)$ + **1**$Na_3C_6H_5O_7\ (aq)$

9. **1**$Ni\ (s)$ + **4**$CO\ (g) \rightarrow$ **1**$Ni(CO)_4\ (g)$

10. **1**$K_2PtCl_4\ (aq)$ + **2**$NH_3\ (aq) \rightarrow$ **1**$Pt(NH_3)_2Cl_2\ (s)$ + **2**$KCl\ (aq)$