# Ch 100: Fundamentals for Chemistry

Chapter 8: Chemical Equations
Lecture Notes

## Chemical Equations (Intro)

1. Chemical equations are used to symbolically describe chemical reactions
2. In a chemical equation (or reaction for that matter) the substances that undergo chemical change(s) are called the reactants
3. The resulting substances formed are called the products
4. The standard representation of a chemical equation:

**Reactant(s) → Product(s)**

**Example**: The production of water

$$\underline{2}H_2\,(g) + \underline{1}O_2\,(g) \rightarrow \underline{2}H_2O\,(g)$$

The underlined numbers are called coefficients.

a. The number of each molecule for each reactant & product in the chemical reaction
b. They are always whole numbers

## Chemical Equations (cont.)

Balanced chemical equations indicate the ____

1. identity of each reactant & product involved in the reaction
2. phase of each reactant and product involved in the reaction (i.e. solid (s), liquid (l) or gas (g))
3. relative quantity of each reactant and product involved in the reaction (the coefficients!)
4. relative **molar** quantity of each reactant and product involved in the reaction (the coefficients!)

## Balancing Chemical Equations

According to the Law of Mass Conservation (& John Dalton!) matter is never created nor destroyed during chemical reactions

- All of the atoms in the reactants of a chemical reaction must be accounted for in the products

The Basic Process of Balancing Chemical Equations:

1. Identify all reactants & products in the reaction & write out their formulas *(this is the unbalanced chemical equation)*
2. Count the number of each atom for each compound for each reactant & product

   (these values must be the same for both reactants & products when the reaction is balanced!)
4. Starting with the most "complicated" molecule, systematically adjust the coefficients to balance # of the atoms on each side of the reaction *(balance one atom at a time)*
5. Repeat until all atoms are balanced for the reaction
6. Now you should have a balanced chemical equation!

## Balancing Chemical Equations (example)

When sodium metal is added to water a violent reaction takes place producing aqueous sodium hydroxide and releasing hydrogen gas.

1. Write out the unbalanced chemical reaction:

2. Now, balance the chemical reaction:

## Balancing Chemical Reactions (Hint)

When a polyatomic ion(s) appears on both the reactant & product side of the reaction unchanged, treat the whole ion as a "unit" when balancing the reaction

Example:

$$AgNO_3(aq) + CaCl_2(aq) \rightarrow AgCl(s) + Ca(NO_3)_2(aq)$$

1. Note the nitrate ion ($NO_3^-$) gets swapped between the $Ag^+$ and the $Ca^{2+}$ ions in this reaction
2. So $NO_3^-$ can be treated as a whole unit when balancing this reaction
3. Balance it!

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## Common Classifications for Chemical Reactions

1. **Combination (or Synthesis):** reactions in which reactants combine to make *one product*
2. **Decomposition:** reactions in which *one reactant* breaks down into smaller products
3. **Single Displacement:** reactions where a part of one reactant is displaced and combined with another reactant

   $$2Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$
4. **Double Displacement:** reactions where a part of two reactants is displaced and exchanged

   $$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

   *Examples:*
   a. *Acid-base neutralization*
   b. *Formation of insoluble products (Precipitation reactions)*
   c. *Metal oxide + acid*
   d. *Gas formation*

5. **Oxidation-Reduction Reactions:** reactions involving the transfer or rearrangement of electrons

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## Combination & Decomposition Reactions

1. Reactions in which chemicals combine to make *one product* are called **Combination or Synthesis Reactions**
   a. Metal + Nonmetal reactions can be classified as Combination Reactions

   $$2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$$

   b. Reactions between Metals or Nonmetals with $O_2$ can be classified as Combination Reactions

   $$N_2(g) + O_2(g) \rightarrow 2\ NO(g)$$

   **Note:** these two types of Combination Reactions are also subclasses of *Oxidation-Reduction Reactions*

2. Reactions in which *one reactant* breaks down into smaller molecules are called **Decomposition Reactions**
   a. Decomposition reactions are generally initiated by the addition of energy (via electric current or heat)
   b. Decomposition reactions are the opposite of Combination Reactions:

   $$2\ NaCl(l) \rightarrow 2\ Na(l) + Cl_2(g)$$

## Single Displacement Reactions

Single displacement reactions involve one part of a reactant being transferred to another

The basic pattern of the single displacement reaction:

$$XY + A \rightarrow X + AY$$

*Example 1:* Metal + Acid → Salt + Hydrogen

$$Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

*Example 2:* Metal + Water → Hydrogen + Metal Oxide (or metal hydroxide)

$$3\ Fe(s) + 4\ H_2O(l) \rightarrow 4\ H_2(g) + Fe_2O_3(s)$$

*Example 3:* Metal + Salt → Metal + Salt

$$2\ Al(s) + Fe_2O_3(s) \rightarrow 2\ Fe(s) + Al_2O_3(s)$$

*Example 4:* Halogen + Halide Salt → Halogen + Halide Salt

$$Cl_2(g) + 2\ NaBr(s) \rightarrow Br_2(g) + 2\ NaCl(s)$$

## Double Displacement Reactions

Double Displacement Reactions involve the double exchange of a component (such as ions) between two reactants

The basic form of double displacement reactions is:

$$XY + AB \rightarrow XB + AY$$

where X, Y, A, and B are the components of the reactants

Example 1: Acid Base Neutralization

$$H_2SO_4(aq) + Ca(OH)_2(aq) \rightarrow CaSO_4(aq) + 2\ H_2O(l)$$

*or 2HOH*

Example 2: Metal Oxide + Acid

$$CaO(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l)$$

*or HOH*

Example 3: Formation of an Insoluble Precipitate (Precipitation)

$$KCl(aq) + AgNO_3(aq) \rightarrow KNO_3(aq) + AgCl(s)$$

Example 4: Formation of a Gas

$$HCl\ (aq) + ZnS\ (s) \rightarrow ZnCl_2(aq) + H_2S\ (g)$$

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# Solubility & Precipitation Reactions

1. When 2 solutions are combined and result in the formation of an insoluble product:
   a. The product will not dissolve in the solvent
   b. The product will form a precipitate
2. Solubility is an intrinsic physical property and a measure of how well a substance (solute) will dissolve in another substance (solvent)
   a. Solubility is temperature dependent
   b. Solid solubility increases with increased temperature (i.e. you can dissolve more sugar in hot water than in cold water)
   c. Gas solubility increases with decreased temperature (i.e. you can dissolve more $CO_2$ in cold water than hot water)
3. A solute is soluble if any of it will dissolve in a solvent
   Eg. NaCl is soluble in water
4. A solute is insoluble if no appreciable amount of it will dissolve in solvent
   Eg. AgCl is insoluble in water
5. Precipitation (formation of an insoluble solid) is one indication that a chemical change has occurred!

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# General Rules for Solubility

1. Most compounds that contain $NO_3^-$ ions are soluble
2. Most compounds that contain $Na^+$, $K^+$, or $NH_4^+$ ions are soluble
3. Most compounds that contain $Cl^-$ ions are soluble, except AgCl, $PbCl_2$, and $Hg_2Cl_2$
4. Most compounds that contain $SO_4^{2-}$ ions are soluble, except $BaSO_4$, $PbSO_4$, and $CaSO_4$
5. Most compounds that contain $OH^-$ ions are slightly soluble (will precipitate), except NaOH, KOH, are soluble and $Ba(OH)_2$, $Ca(OH)_2$ are moderately soluble
6. Most compounds that contain $S^{2-}$, $CO_3^{2-}$, or $PO_4^{3-}$ ions are slightly soluble (will precipitate)

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## Oxidation-Reduction Reactions

Reactions that involve transfer or rearrangement of electrons are called **oxidation-reduction reactions.**

*Examples of oxidation-reduction reactions:*

1. Metal + Nonmetal: $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
   a. The metal loses an electron(s) and becomes a cation (oxidation → metal gets oxidized: $Na \rightarrow Na^+ + e^-$)
   b. The nonmetal gains an electron(s) and becomes an anion (reduction → nonmetal gets reduced: $Cl + e^- \rightarrow Cl^-$)
   c. In this reaction, electrons are transferred from the metal to the nonmetal
2. $O_2$ as a reactant or product: $CH_4(s) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
   a. In this reaction, it is not obvious that electron transfer has taken place. *In this case, oxidation states are altered.*
   b. Often this type of reaction involves the release of large amounts of energy, even combustion

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## Rates of Chemical Reactions

How quickly a chemical reaction occurs is indicated by its reaction rate

1. How quickly the concentration of products increases
2. How quickly the concentration of reactants decreases

The Factors that influence reaction rates:

1. Reactants must be in contact
   - Reactions occur due to collisions
   - Without contact between reactants there can be no reaction
2. Concentration of reactants
   - The more reactant molecules packed into a given space the more likely a collision (& reaction) will occur
3. Temperature
   - the average KE of each reactant affects how much energy will be transferred between reactants during a molecular collision
   - Molecules must transfer enough KE to break the existing bonds

# The Role of Energy in Chemical Reactions

Energy transformations always accompany chemical reactions:

1. Energy is required to break bonds (energy absorbed or activation energy)
2. Energy is released when bonds are formed

***Note:*** *The amount of energy required to break a chemical bond equals the energy released when that type of bond is formed, this is called the* ***Bond Energy***

For a chemical reaction to occur:

1. Energy must be absorbed in order to break chemical bonds in the reactants
2. Energy is released as new bonds are formed in the products

i. Endothermic reactions absorb more energy than they release

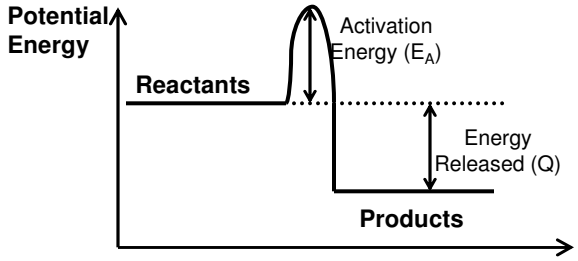
$$N_2(g) + O_2(g) + 393\ kJ \rightarrow 2NO(g)$$

ii. Exothermic reactions release more energy than they absorb

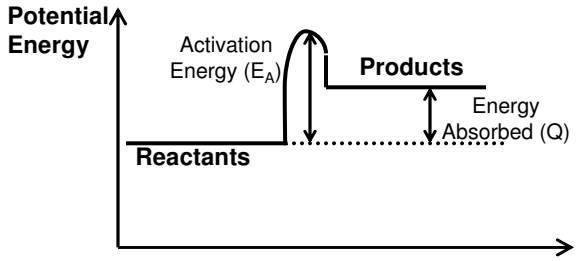
$$H_2(g) + Cl_2(g) \rightarrow 2NO(g) + 185\ kJ$$

# Energy in Chemical Reactions

**Exothermic Reactions**

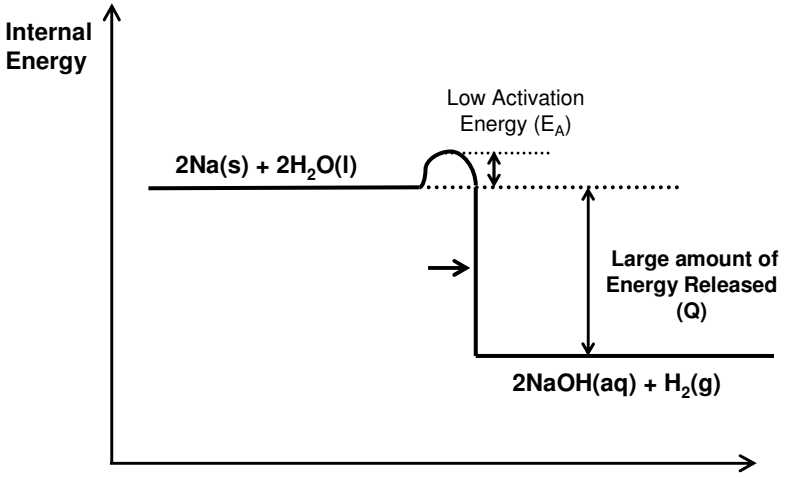



**Endothermic Reactions**



## Energy in Reactions (cont.)

**Example: Sodium Water Reaction**




## Catalysts

1. Catalysts are <u>substances that speed up chemical reactions</u>
   a. Allow reactions to occur that might not otherwise take place (due to low temperature for example)
   b. Lower activation energy for a chemical reaction
2. Participation of catalysts in a chemical reaction
   a. They may undergo a chemical change as a reactant but they are always recycled as a product (so there is no net change in the catalyst molecule)
3. Catalysts are indicated in a chemical reaction by placing the chemical formula over/under the reaction arrow.

$$\textbf{Reactants} \xrightarrow{\text{catalyst}} \textbf{Products}$$

**Example:** The breakdown of hydrogen peroxide

$$2H_2O_2\ (aq) \xrightarrow{catalase} H_2O\ (l) + O_2(g)$$

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# Catalysts & Energy in Reactions

## Catalysts lower Activation Energy