Chapter 5
Chemical Reactions and Equations

• What is a chemical reaction?
• How do we know a chemical reaction occurs?
• Writing chemical equations
• Predicting chemical reactions
• Representing reactions in aqueous solution

What is a chemical reaction?

- Chemical reaction
  - The conversion of one substance or set of substances into another
- Reactant
  - Any substance that is started with is called a reactant
- Product
  - New substances that form during the course of the reaction
  - Products differ from reactants only in the arrangement of their component atoms

A Microscopic View
Chemical Reaction
How do we know a chemical reaction occurs?

- Do any of the pictures below show a chemical reaction? How can you tell?

Physical Clues of a Chemical Reaction

- The most common evidence of a chemical reaction is:
  - Change in color
  - Production of light
  - Formation of a solid (such as a precipitate in solution, or smoke in air, or a metal coating)
  - Formation of a gas (bubbles in solution or fumes in the gaseous state)
  - Absorption or release of heat (sometimes appearing as a flame)

Writing Chemical Equations

- Chemical equation
  - A symbolic representation of a chemical reaction

- Balanced equation
  - The number of atoms of each element is the same in the products as in that reactants
  - Conservation of mass is always maintained
### Writing Chemical Equations

**Aluminum + iron(III) oxide $\rightarrow$ aluminum oxide + iron**

\[
\text{Al}(s) + \text{Fe}_2\text{O}_3(s) \rightarrow \text{Al}_2\text{O}_3(s) + \text{Fe}(s)
\]

<table>
<thead>
<tr>
<th># of atoms (reactants)</th>
<th># of atoms (products)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 Al</td>
<td>2 Al</td>
</tr>
<tr>
<td>2 Fe</td>
<td>1 Fe</td>
</tr>
<tr>
<td>3 O</td>
<td>3 O</td>
</tr>
</tbody>
</table>

Therefore, we need to balance the equation with coefficients:

\[
2 \text{Al}(s) + \text{Fe}_2\text{O}_3(s) \rightarrow \text{Al}_2\text{O}_3(s) + 2 \text{Fe}(s)
\]

### Writing Chemical Equations

**Methane + oxygen $\rightarrow$ carbon dioxide + water**

\[
\text{CH}_4(g) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g)
\]

Currently, the number of atoms of each element is shown below. These numbers were obtained by multiplying the subscript to the right of the element’s symbol by the stoichiometric coefficient.

<table>
<thead>
<tr>
<th># of atoms (reactants)</th>
<th># of atoms (products)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 C</td>
<td>1 C</td>
</tr>
<tr>
<td>4 H</td>
<td>2 H</td>
</tr>
<tr>
<td>2 O</td>
<td>3 O</td>
</tr>
</tbody>
</table>

### Writing Chemical Equations

**CH\(_4\)(g) + O\(_2\)(g) \rightarrow CO\(_2\)(g) + H\(_2\)O(g)**

First, we look at the carbon atoms. Since the number of carbon atoms on the reactant side is already equal to the number of carbon atoms on the product side, we don’t need to add coefficients.

Next, we look at the hydrogen atoms. Currently, there are four hydrogen atoms on the reactant side and 2 hydrogen atoms on the product side. Thus, we need to add a coefficient of 2 in front of water to make the hydrogen atoms equal.

\[
\text{CH}_4(g) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(g)
\]
Writing Chemical Equations

\[ \text{CH}_4(g) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(g) \]

Finally, we look at the oxygen atoms. Currently, there are 2 oxygen atoms on the reactant side and 4 oxygen atoms (combined from carbon dioxide and water) on the product side. Thus, we add a coefficient of 2 in front of the oxygen gas.

\[ \text{CH}_4(g) + 2 \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(g) \]

Now the reaction is balanced!

A General Approach to Balancing Equations

1. Identify the reactants and products and write their correct formulas.
2. Write a skeletal equation including physical states.
3. Change coefficients one at a time until the atoms of each element are balanced. (Start with the elements that occur least often in the equation)
4. Make a final check by counting the atoms of each element on both sides of the equation.

Practice – Balancing Equations

- Balance the following equations:
  1. Potassium chlorate $\rightarrow$ potassium chloride + oxygen
  2. Aluminum acetate reacts with potassium sulfate to form potassium acetate and aluminum sulfate
  3. Hexane ($\text{C}_6\text{H}_{14}$) reacts with oxygen gas to form carbon dioxide and water
  4. Zinc reacts with hydrochloric acid to form zinc chloride and hydrogen gas
1. Potassium chlorate → potassium chloride + oxygen

First, translate the chemical names into chemical formulas:

Potassium chlorate → potassium chloride + oxygen

K\(^{+1}\) ClO\(^{3-}\) → K\(^{+1}\) Cl\(^{-1}\) O\(_2\) (diatomic)

KClO\(_3\) (s) → KCl(aq) + O\(_2\) (g)

Next, balance the chemical reaction:

2 KClO\(_3\) (s) → 2 KCl(aq) + 3 O\(_2\) (g)

2. Aluminum acetate reacts with potassium sulfate to form potassium acetate and aluminum sulfate

Aluminum acetate + potassium sulfate →

Al\(^{+3}\) C\(_2\)H\(_3\)O\(_2\)\(^{3+}\) + K\(^{+1}\) SO\(_4^{2-}\)

Potassium acetate + aluminum sulfate

K\(^{+1}\) C\(_2\)H\(_3\)O\(_2\)\(^{-1}\) Al\(^{+3}\) SO\(_4^{2-}\)

Al(C\(_2\)H\(_3\)O\(_2\))\(_3\) (aq) + K\(_2\)SO\(_4\) (aq) → KC\(_2\)H\(_3\)O\(_2\) (aq) + Al\(_2\)(SO\(_4\))\(_3\) (s)

Balanced:

2 Al(C\(_2\)H\(_3\)O\(_2\))\(_3\) (aq) + 3 K\(_2\)SO\(_4\) (aq) → 6 KC\(_2\)H\(_3\)O\(_2\) (aq) + Al\(_2\)(SO\(_4\))\(_3\) (s)

3. Hexane (C\(_6\)H\(_{14}\)) reacts with oxygen gas to form carbon dioxide and water

Hexane was given. Oxygen gas is diatomic. Carbon dioxide has 1 carbon atom and 2 oxygen atoms (using the Greek prefixes) and water has 2 hydrogen atoms and 1 oxygen atom.

C\(_6\)H\(_{14}\) (l) + O\(_2\) (g) → CO\(_2\) (g) + H\(_2\)O (g)

Balanced:

2 C\(_6\)H\(_{14}\) (l) + 9.5 O\(_2\) (g) → 6 CO\(_2\) (g) + 7 H\(_2\)O (g)

OR

2 C\(_6\)H\(_{14}\) (l) + 19 O\(_2\) (g) → 12 CO\(_2\) (g) + 14 H\(_2\)O (g)
Practice – Balancing Equations

3. Zinc reacts with hydrochloric acid to form zinc chloride and hydrogen gas

Zinc + Hydrochloric acid → Zinc chloride + hydrogen

\[ \text{Zn} + \text{H}^+ \text{Cl}^- \rightarrow \text{Zn}^{2+} \text{Cl}^- + \text{H}_2 \]  

(diatom)

\[ \text{Zn} + \text{HCl}(aq) \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g) \]

Balanced:
\[ \text{Zn}(s) + 2 \text{HCl}(aq) \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g) \]

Predicting Classes of Reactions

• If we depict elements and compounds with letters and spheres:

If A and B are elements (monatomic) and C, D, E, and F are possible atoms, monatomic ions, or polyatomic ions arranged to form molecules, then how many different types of chemical reactions can you identify by rearranging these atoms or groups of atoms?

(NOTE: Do not react more than two elements or molecules)

Table 5.1 The Classes of Chemical Reactions

<table>
<thead>
<tr>
<th>Class</th>
<th>Reactants</th>
<th>Products</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Decomposition</td>
<td>1 compound</td>
<td>2 elements (or smaller compounds)</td>
<td>CD → C + D</td>
</tr>
<tr>
<td>Combination</td>
<td>2 elements or compounds 1 compound</td>
<td>A + B → AB</td>
<td></td>
</tr>
<tr>
<td>Single-Replacement</td>
<td>1 element &amp; 1 compound</td>
<td>1 element &amp; 1 compound</td>
<td>A + CD → C + AD</td>
</tr>
<tr>
<td>Double-Replacement</td>
<td>2 compounds</td>
<td>2 compounds</td>
<td>CD + EF → CF + ED</td>
</tr>
</tbody>
</table>
Practicing – Classes of Reactions

• Classify each of the following as a decomposition, combination, single-replacement, or double-replacement reaction.

1. \( \text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s) \)
2. \( \text{CuCl}_2(aq) + \text{Na}_2\text{S}(aq) \rightarrow \text{CuS}(s) + 2 \text{NaCl}(aq) \)
3. \( \text{NiSO}_3(s) \rightarrow \text{NiO}(s) + \text{SO}_2(g) \)
4. \( \text{Ca}(s) + \text{PbCl}_2(aq) \rightarrow \text{CaCl}_2(aq) + \text{Pb}(s) \)

Practicing Solutions – Classes of Reactions

• Classify each of the following as a decomposition, combination, single-replacement, or double-replacement reaction.

1. \( \text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s) \) **Combination**
2. \( \text{CuCl}_2(aq) + \text{Na}_2\text{S}(aq) \rightarrow \text{CuS}(s) + 2 \text{NaCl}(aq) \) **Double-replacement**
3. \( \text{NiSO}_3(s) \rightarrow \text{NiO}(s) + \text{SO}_2(g) \) **Decomposition**
4. \( \text{Ca}(s) + \text{PbCl}_2(aq) \rightarrow \text{CaCl}_2(aq) + \text{Pb}(s) \) **Single-replacement**

Decomposition Reactions

• A *compound breaks down into its component elements.*

• Example:
  \( 2 \text{HgO}(s) \text{ heat} \rightarrow 2 \text{Hg}(l) + \text{O}_2(g) \)
Decomposition Reactions

- Ammonium compounds lose ammonia gas.
  \[(\text{NH}_4)^+\text{SO}_4(s) \rightarrow \text{NH}_3(g) + \text{H}_2\text{SO}_4(l)\]

- Oxoacids decompose in a similar way to form nonmetal oxides and water.
  \[\text{H}_2\text{CO}_3(aq) \rightarrow \text{H}_2\text{O}(l) + \text{CO}_2(g)\]

- Metal carbonates, except those of group 1A metals, decompose to metal oxides and carbon dioxide gas.
  \[\text{NiCO}_3(s) \rightarrow \text{NiO}(s) + \text{CO}_2(g)\]

- Peroxides decompose to oxides and oxygen gas.
  \[2\text{H}_2\text{O}_2(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{O}_2(g)\]

- Oxides and halides of the metals Au, Pt, and Hg decompose to the elements.
  \[2\text{HgO}(s) \rightarrow 2\text{Hg}(l) + \text{O}_2(g)\]

- Oxoacids decompose in a similar way to form nonmetal oxides and water.
  \[\text{H}_3\text{PO}_4(aq) \rightarrow \text{H}_3\text{O}(l) + \text{PO}_4^{3-}\]

- Ammonium compounds lose ammonia gas.
  \[\text{(NH}_4\text{)}_2\text{SO}_4(s) \rightarrow \text{NH}_3(g) + \text{H}_2\text{SO}_4(l)\]

Table 5.2: Decomposition Reactions That Occur When Compounds Are Heated

<table>
<thead>
<tr>
<th>Reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Oxides and halides of the metals Au, Pt, and Hg decompose to the elements.</td>
</tr>
<tr>
<td>Peroxides decompose to oxides and oxygen gas.</td>
</tr>
<tr>
<td>Metal carbonates, except those of group 1A metals, decompose to metal oxides and carbon dioxide gas.</td>
</tr>
<tr>
<td>Oxoacids decompose in a similar way to form nonmetal oxides and water.</td>
</tr>
<tr>
<td>Ammonium compounds lose ammonia gas.</td>
</tr>
</tbody>
</table>

Combination Reactions

- Two elements, an element and a compound, or two compounds react to produce a single compound.
  - Most metals react with most nonmetals to form ionic compounds.
  - A nonmetal may react with a more reactive nonmetal to form a molecular compound.
  - A compound and an element may combine to form another compound if one exists with a higher atom: atom ratio.
  - Two compounds may react to form a new compound.
- Example: \[2\text{Al}(s) + 3\text{Br}_2(l) \rightarrow 2\text{AlBr}_3(s)\]
Combination Reactions

Single-Displacement Reactions

- A free element displaces another element from a compound to form another compound and a different free element.

- Example:
  \[ 2 \text{Al}(s) + \text{Fe}_3\text{O}_4(s) \rightarrow \text{Al}_2\text{O}_3(s) + 2 \text{Fe}(s) \]
Activity Series

- This is a list of metals in order of their reactivity
- A more active element displaces a less active element from its compounds.

Single-Displacement Reactions: Copper and Silver Nitrate

Single-Displacement Reactions: Copper and Silver Nitrate
Single-Displacement Reactions: Copper and Silver Nitrate

Double-Displacement Reactions

• *Two compounds exchange ions or elements to form new compounds.*
  – Precipitation reactions
  – Gas-forming reactions
  – Acid-Base Neutralizations

Precipitation Reactions

• *When one product separates from the reaction solution, it is insoluble.*
  – An insoluble compound formed in a reaction is a precipitate.
  – Example:
  \[ \text{BaCl}_2(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq}) \rightarrow 2 \text{NaCl}(\text{aq}) + \text{BaSO}_4(\text{s}) \]
Table 5.3 Solubility Rules

<table>
<thead>
<tr>
<th>Rule</th>
<th>Solubility Rules</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na⁺, K⁺, NH₄⁺, and other alkali metal ions</td>
<td>Most compounds of these cations and amphoteric oxides are soluble.</td>
</tr>
<tr>
<td>SO₄²⁻, PO₄³⁻, SO₃²⁻</td>
<td>Most sulfates and phosphates are soluble. Exceptions are BaSO₄, Hg₂SO₄, PbSO₄, BaCO₃, Hg₂CO₃, and Ag₂SO₄.</td>
</tr>
<tr>
<td>Cl⁻, Br⁻, I⁻</td>
<td>Most halides, bromides, and iodides are soluble. Exceptions are AgCl, AgBr, AgI, and Hg₂Cl₂, Br₂, I₂.</td>
</tr>
<tr>
<td>Ag⁺</td>
<td>Silver compounds, except Ag₂O, and Ag₂CO₃ are insoluble. Ag₂(SO₄)₃ is slightly soluble.</td>
</tr>
<tr>
<td>OH⁻, CO₃²⁻</td>
<td>Oxides and hydroxides are insoluble. Exceptions are alkali metal hydroxides, Ba(OH)₂, Sr(OH)₂, and Ca(OH)₂ (somewhat soluble).</td>
</tr>
<tr>
<td>NO₃⁻</td>
<td>Sulfides are insoluble. Exceptions are compounds of Na⁺, K⁺, NH₄⁺, and the alkali metal ions.</td>
</tr>
<tr>
<td>CO₃²⁻, PO₄³⁻, SO₄²⁻, NO₃⁻</td>
<td>Most carbonates, phosphates, sulfates, and nitrates are insoluble. Exceptions are compounds of Na⁺, K⁺, and NH₄⁺.</td>
</tr>
</tbody>
</table>

Practice - Precipitation Reactions

- Jennifer mixes solutions of cadmium(II) nitrate and sodium sulfide and obtains an orange precipitate that is the pigment cadmium orange. Identify the precipitate, and write a balanced chemical equation for the reaction she carried out.

Practice Solutions - Precipitation Reactions

- The starting solutions contain Cd²⁺(aq), NO₃⁻(aq), Na⁺(aq), and S²⁻(aq).
- The solubility rules say that most compounds of sodium are soluble and all nitrates are soluble, so the precipitate must be formed between cadmium(II) ions and sulfide ions.
- By matching the positive and negative charges to get electrical neutrality, we can determine the formula of the precipitate to be CdS(s).
Practice Solutions - Precipitation Reactions

• The solution contains Na⁺(aq) and NO₃⁻(aq) ions. We can write the skeletal equation for this reaction as:

\[ \text{Cd(NO}_3\text{)}_2(aq) + \text{Na}_2\text{S}(aq) \rightarrow \text{CdS(s) + NaNO}_3(aq) \]

• Balanced, the equation is:

\[ \text{Cd(NO}_3\text{)}_2(aq) + \text{Na}_2\text{S}(aq) \rightarrow \text{CdS(s) + 2 NaNO}_3(aq) \]

Gas-forming Reactions

• One product compound separates from the reaction mixture because it forms a gas.

• Example:

\[ \text{CaCO}_3(s) + 2 \text{HCl(aq)} \rightarrow \text{CaCl}_2(aq) + \text{CO}_2(g) + \text{H}_2\text{O(l)} \]

Acid-Base Neutralizations

• A double-displacement reaction involving an acid and a base.
  – An acid reacts with a base to form an ionic compound and water.
  – Example:

\[ \text{HCl(aq) + NaOH(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)} \]

Acid Base Salt Water
Practice – Acid Base Neutralizations

- Calcium oxide is the white powder, lime. When added to water, it makes slaked lime, which is a solution of the base calcium hydroxide. If sulfuric acid is added to slaked lime, then what products form? Write the balanced chemical equation for this reaction.

Practice Solutions – Acid Base Neutralizations

- The starting solutions contain $\text{H}^+ (aq)$, $\text{SO}_4^{2-} (aq)$, $\text{Ca}^{2+} (aq)$, and $\text{OH} (aq)$.
- The solubility rules say that most sulfate ions are soluble, unless combined with calcium. Thus, calcium sulfate is our precipitate. Hydrogen ions and hydroxide ions form water.
- By matching the positive and negative charges to get electrical neutrality, we can determine the formula of the precipitate to be $\text{CaSO}_4 (s)$.

Practice Solutions - Precipitation Reactions

- The solution contains no spectator ions. We can write the skeletal equation for this reaction as:
  
  $\text{H}_2\text{SO}_4 (aq) + \text{Ca(OH)}_2 (aq) \rightarrow \text{CaSO}_4 (s) + \text{H}_2\text{O} (l)$

- Balanced, the equation is:
  
  $\text{H}_2\text{SO}_4 (aq) + \text{Ca(OH)}_2 (aq) \rightarrow \text{CaSO}_4 (s) + 2 \text{H}_2\text{O} (l)$
Combustion Reactions

- Any reaction involving oxygen as a reactant and that rapidly produces heat and flame.

Example:

\[
\text{CH}_4(g) + 2 \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(g)
\]

Representing Reactions in Aqueous Solution

- When compounds exist in solution, they exist as ions.
  - Insoluble compounds do not exist as ions.
  - \( \text{Pb(NO}_3\text{)}_2(aq) + \text{K}_2\text{CrO}_4(aq) \rightarrow \text{PbCrO}_4(s) + 2 \text{KNO}_3(aq) \)

- Ionic Equation
  - Separate every compound with an (aq) by it into component ions.
  - You cannot separate solids, liquids, or gases.
  - \( \text{Pb}^{2+} + 2 \text{NO}_3^- + 2 \text{K}^+ + \text{CrO}_4^{2-} \rightarrow \text{PbCrO}_4(s) + 2 \text{K}^+ + 2 \text{NO}_3^- \)
  - Spectator ions: ions that exist on both sides of the arrow.

- Net Ionic Equation
  - The remaining equation without the spectator ions
  - \( \text{Pb}^{2+}(aq) + \text{CrO}_4^{2-}(aq) \rightarrow \text{PbCrO}_4(s) \)
Barium sulfate, used in the white pigment lithopone, can be prepared by mixing solutions of barium chloride and sodium sulfate. Write balanced molecular, ionic, and net ionic equations for this reaction.

Both reactants are strong electrolytes, so they dissociate into the ions $\text{Ba}^{2+}(aq)$, $\text{Cl}^-(aq)$, $\text{Na}^+(aq)$ and $\text{SO}_4^{2-}(aq)$ in solution. Possible formulas for the precipitate are $\text{BaCl}_2$, $\text{Na}_2\text{SO}_4$, $\text{NaCl}$, or $\text{BaSO}_4$. The first two choices can be eliminated because they are starting reactants. The solubility rules (Table 5.3) indicate that $\text{NaCl}$ is soluble and $\text{BaSO}_4$ is insoluble. Thus, we can write the molecular equation:

$\text{BaCl}_2(aq) + \text{Na}_2\text{SO}_4(aq) \rightarrow \text{BaSO}_4(s) + 2\text{NaCl}(aq)$
Since $\text{BaCl}_2$, $\text{Na}_2\text{SO}_4$, and NaCl are all strong electrolytes in solution, we can write their formulas as separate ionic formulas, giving the following ionic equation:

$$\text{Ba}^{2+}(aq) + 2 \text{Cl}^{-}(aq) + 2 \text{Na}^{+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{BaSO}_4(s) + 2 \text{Na}^{+}(aq) + 2 \text{Cl}^{-}(aq)$$

Finally, we note that $\text{Na}^+$ and $\text{Cl}^-$ occur on both sides of the equation, so they are spectator ions. They can be eliminated, giving us the following net ionic equation:

$$\text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{BaSO}_4(s)$$