Chapter 7

Stoichiometry

Mass Relationships and Chemical Reactions
Chapter Outline

• 7.1 Chemical Reactions and the Conservation of Mass
• 7.2 Balancing Chemical Equations
• 7.3 Combustion Reactions
• 7.4 Stoichiometric Calculations and the Carbon Cycle
• 7.5 Percent Composition and Empirical Formulas
• 7.6 Empirical and Molecular Formulas
• 7.7 Combustion Analysis
• 7.8 Limiting Reagents and Percent Yield
Chemical Reactions

Combination Reaction: two or more substances combine to form one product.

\[
\text{N}_2\text{O}_5(g) + \text{H}_2\text{O}(\ell) \rightarrow 2 \text{HNO}_3(aq)
\]

Reactants \rightarrow Products
Chemical Equation

• Chemical equation:
  • Describes proportions of reactants (the substances that are consumed) and products (the substances that are formed) during a chemical reaction.
  • Describes the changes on the atomic level.
    • $\text{SO}_3(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_4(l)$
    • $\text{Fe}_2\text{O}_3(s) + 3\text{H}_2\text{SO}_4(aq) \rightarrow 3\text{H}_2\text{O}(l) + \text{Fe}_2(\text{SO}_4)_3(aq)$
  • Physical state of reactants and products:
    • $(s)$ = solid; $(l)$ = liquid; $(g)$ = gas; $(aq)$ = aqueous soln.
States of Substances

- States are shown by abbreviations in parenthesis after each chemical
  \[ \text{H}_2\text{O (s), H}_2\text{O (l), H}_2\text{O (g)} \]

- Standard phases are:
  - (s) – solid
  - (l) – liquid
  - (g) – gas
  - (aq) – aqueous – dissolved in water
  - (↑) – gas produced from aqueous phase
  - (↓) – solid produced from aqueous phase
Types of Reactions

- **Synthesis** – compound formed from its base elements:
  \[ \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 \]

- **Decomposition** – compound decomposes into its base elements:
  \[ 2\text{NH}_3 \rightarrow \text{N}_2 + 3\text{H}_2 \]
Types of Reactions

- Single replacement – an element replaces another in a compound:
  \[ 2\text{NaBr} + \text{I}_2 \rightarrow 2\text{NaI} + \text{Br}_2 \]

- Double replacement – two elements or polyatomic ions in two separate compounds switch places:
  \[ \text{KNO}_3 + \text{Ca(OH)}_2 \rightarrow \text{KOH} + \text{Ca(NO}_3\text{)}_2 \]
Types of Reactions

Synthesis

\[ X + Y \rightarrow XY \]

Decomposition

\[ X_2Y \rightarrow X + Y \]

Single-replacement

\[ X + YZ \rightarrow Y + XZ \]

Double-replacement

\[ AX + BY \rightarrow AY + BX \]
Types of Reactions

- Complete combustion – fuel and oxygen produce water and carbon dioxide:
  \[
  \text{CH}_4 + \text{O}_2 \rightarrow \text{H}_2\text{O} + \text{CO}_2
  \]

- Incomplete combustion – fuel and oxygen produce water and carbon Monoxide:
  \[
  \text{CH}_4 + \text{O}_2 \rightarrow \text{H}_2\text{O} + \text{CO}
  \]
Law of Conservation of Mass

- **Law of conservation of mass**
  - The sum of the masses of the reactants in a chemical equation is equal to the sum of the masses of the products.

- **Stoichiometry**
  - Quantitative relation between reactants and products in a chemical equation
  - Indicated in chemical equation by coefficients
Moles and Chemical Equations

\[ \text{Fe}_2\text{O}_3(\text{s}) + 3\text{H}_2\text{SO}_4(\text{aq}) \rightarrow 3\text{H}_2\text{O}(\text{l}) + \text{Fe}_2(\text{SO}_4)_3(\text{aq}) \]

- **Chemical Equation**
  - Indicates substances involved (reactants, products)
- **Coefficients**
  - Indicate proportions of reactants and/or products
  - On macroscale, indicate number of moles of each substance.
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Balancing Chemical Equations

- Balanced chemical equations follow the law of conservation of mass.
  - (not balanced)

$$\text{CH}_4(g) + \text{NH}_3(g) \rightarrow \text{HCN}(g) + \text{H}_2(g)$$

reactants \quad \rightarrow \quad \text{products}
Balancing Chemical Equations

• Three Step Approach:
  • Write correct formulas for reactants and products, including physical states.
  • Balance an element that appears in only one reactant and product first.
  • Choose coefficients to balance other elements as needed.
  • Reduce coefficients to lowest whole numbers.
Balancing Chemical Equations

\[ \text{CH}_4(g) + \text{NH}_3(g) \rightarrow \text{HCN}(g) + \text{H}_2(g) \]

**Reactants**

- CH\(_4\)(g)
- NH\(_3\)(g)

**Products**

- HCN(g)
- H\(_2\)(g)

<table>
<thead>
<tr>
<th>Element</th>
<th>Reactant Side</th>
<th>Product Side</th>
<th>Balanced?</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>1</td>
<td>1</td>
<td>✔️</td>
</tr>
<tr>
<td>N</td>
<td>1</td>
<td>1</td>
<td>✔️</td>
</tr>
<tr>
<td>H</td>
<td>4 + 3 = 7</td>
<td>1 + 2 = 3</td>
<td>✗</td>
</tr>
</tbody>
</table>
Balancing Chemical Equations

\[ \text{CH}_4(g) + \text{NH}_3(g) \rightarrow \text{HCN}(g) + 3\text{H}_2(g) \]

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<tr>
<td>C</td>
<td>1</td>
<td>1</td>
<td>✓</td>
</tr>
<tr>
<td>N</td>
<td>1</td>
<td>1</td>
<td>✓</td>
</tr>
<tr>
<td>H</td>
<td>4 + 3 = 7</td>
<td>1 + (3 \times 2) = 7</td>
<td>✓</td>
</tr>
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Hydrocarbons:
- molecular compounds composed of only hydrogen and carbon.
- “organic” compounds.
- Combustion products are $\text{CO}_2$ and $\text{H}_2\text{O}$.

$$\text{CH}_4(g) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g)$$
Combustion Reactions

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

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<tbody>
<tr>
<td>C</td>
<td>1</td>
<td>1</td>
<td>✓</td>
</tr>
<tr>
<td>H</td>
<td>4</td>
<td>(2 \times 2 = 4)</td>
<td>✓</td>
</tr>
<tr>
<td>O</td>
<td>(2 \times 2 = 4)</td>
<td>((2 \times 1) + 2 = 4)</td>
<td>✓</td>
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</table>
Balance the following combustion reaction.

a) $\text{C}_5\text{H}_{12} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
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Photosynthesis and Respiration

• Photosynthesis:
  • Plants convert \( \text{CO}_2 \) and \( \text{H}_2\text{O} \) into grape:  
    \[
    \text{CO}_2(g) + 6\text{H}_2\text{O}(l) \rightarrow \text{C}_6\text{H}_{12}\text{O}_6(aq) + 6\text{O}_2(g)
    \]

• Respiration (reverse of photosynthesis):
  • Living organisms use glucose as a source of energy:
    \[
    \text{C}_6\text{H}_{12}\text{O}_6(aq) + 6\text{O}_2(g) \rightarrow \text{CO}_2(g) + 6\text{H}_2\text{O}(l)
    \]

• Combustion of Hydrocarbons
  • Returns \( 6.8 \times 10^{12} \) kg/yr of C to atmosphere.
Calculating the mass of a product from the mass of a reactant requires:

- The **mole ratio** from the balanced chemical equation.
- Molar mass of the reactant.
- Molar mass of the product.
Stoichiometry Example

• How much CO\(_2\) enters the atmosphere annually from the combustion of 6.8 x 10\(^{12}\) kg of Carbon?

• Balanced Eq’n: C(s) + O\(_2\)(g) → CO\(_2\)(g)

\[
\left(6.8 \times 10^{12}\ \text{kg C}\right) \rightarrow \text{?} \rightarrow 2.5 \times 10^{13}\ \text{kg CO}_2
\]

\[
\left(\frac{1000\ \text{g}}{\text{kg}}\right) \left(\frac{1\ \text{mole C}}{12\ \text{g}}\right) \left(\frac{1\ \text{mole CO}_2}{1\ \text{mole C}}\right) \left(\frac{44.0\ \text{g CO}_2}{1\ \text{mole CO}_2}\right) \left(\frac{1\ \text{kg}}{1000\ \text{g}}\right)
\]
How much carbon dioxide would be formed if 10.0 grams of $\text{C}_3\text{H}_8$ were completely burned in oxygen?

$$\text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O}$$

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:
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Percent Composition

- **Percent Composition:**
  - the composition of a compound expressed in terms of the percentage by mass of each element
  - \[
  \text{mass of element in compound} \times 100\% \over \text{mass of compound}
  \]
Percent Composition

- **Example:** percent C in CH\(_4\) and C\(_{10}\)H\(_{16}\):

\[
\left( \frac{\text{mass C}}{\text{mass CH}_4} \right) \left( \frac{(12.01 \text{ g/mol C})}{16.04 \text{ g/mol CH}_4} \right) \times 100 = 74.88\%
\]

\[
\left( \frac{\text{mass C}}{\text{mass C}_{10}\text{H}_{16}} \right) \left( \frac{(12.01 \text{ g/mol C}) \times 10}{136.23 \text{ g/mol C}_{10}\text{H}_{16}} \right) \times 100 = 88.16\%
\]
Empirical Formula

• **Empirical Formula:**
  - Formula based on the lowest whole number ratio of its component elements
  - $C_4H_8$ reduces to $CH_2$
  - $Na_2O_2$ reduces to $NaO$
Mass % to Empirical Formula

• Approach:
  1. Assume 100 g of substance.
  2. Convert mass of each element to moles.
  3. Compute mole ratios.
  4. If necessary, convert to smallest whole number ratios by dividing by smallest number moles.
Example:
Compound is 74.88% C and 25.12 % H
1. In 100 g sample, 74.88 g C, 25.12 g H
2. 6.23 moles C, 24.92 moles H
3. Ratio of 24.92 mol H to 6.23 mol C
4. Reduces to 4 moles H:1 mole C
5. Empirical formula of CH₄
For thousands of years the mineral chalcocite has been a highly prized source of copper. Its chemical composition is 79.85% Cu and 20.15% S. What is its empirical formula?

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:
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Empirical vs. Molecular Formulas

• **Empirical Formula:**
  • Simplest whole-number molar ratio of elements in a compound

• **Molecular Formula:**
  • *Actual* molar ratio of elements in a compound
  • Equal to whole # multiple of empirical formula
  • Need empirical formula and molecular mass.
  • Both C$_2$H$_2$ and C$_6$H$_6$ have the same empirical formula, CH
    • Glucose: empirical formula = CH$_2$O
      molecular formula = (CH$_2$O) x 6 = C$_6$H$_{12}$O$_6$
Empirical vs. Molecular Formulas

• Glycoaldehyde (60.05 g/mol):
  • Elemental analysis
    • = 40.00% C, 6.71% H, 53.28% O
  • Mole ratios
    • C:H:O = 3.33 : 6.66 : 3.33, which simplifies to 1:2:1
  • Empirical formula = CH$_2$O (30.02 g?)
  • Molecular formula = CH$_2$O x 2 = C$_2$H$_4$O$_2$
Asbestos was used for years as an insulating material in buildings until prolonged exposure to asbestos was demonstrated to cause lung cancer. Asbestos is a mineral containing magnesium, silicon, oxygen, and hydrogen. One form of asbestos, chrysotile (520.27 g/mol), has the composition 28.03% magnesium, 21.60% silicon, 1.16% hydrogen. Determine the molecular formula of chrysotile.

**Practice: Empirical to Molecular Formula**

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:
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Combustion Analysis

• The % of C and H in $C_aH_b$ determined from the mass of $H_2O$ and $CO_2$ produced by combustion:
  
  $$C_aH_b + \text{excess } O_2 \rightarrow a \ CO_2(g) + \frac{b}{2} H_2O$$
Combustion analysis of an unknown compound indicated that it is 92.23% C and 7.82% H. The mass spectrum indicated the molar mass is 78 g/mol. What is the molecular formula of this unknown compound?

• Collect and Organize:
• Analyze:
• Solve:
• Think about It:
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Hydrogen and Oxygen react to form water:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

$H_2$ (g) = white; $O_2$(g) = red. Which runs out first?
Limiting Reactant

- Limiting Reactant:
  - Substance that is completely consumed in the chemical reaction
  - Determines the amount of product that can be formed during the reaction
  - Identified by:
    - # of moles of reactants
    - Stoichiometry of balanced chemical equation
Limiting Reactants

- \( \text{SO}_3(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2\text{SO}_4(\text{l}) \)

- How much product is obtained from reaction of 20.00 g \( \text{SO}_3 \) and 10.00 g \( \text{H}_2\text{O} \)?

- Find limiting reactant:
  - \( \text{SO}_3 \): \( \frac{20.00 \text{ g}}{80.06 \text{ g/mol}} = 0.2498 \text{ moles} \)
  - \( \text{H}_2\text{O} \): \( \frac{10.00 \text{ g}}{18.02 \text{ g/mol}} = 0.5549 \text{ moles} \)

- Stoichiometry: need 1 mol \( \text{SO}_3 \): 1 mol \( \text{H}_2\text{O} \)

- Limiting reactant = \( \text{SO}_3 \)
If 10.0 g of calcium hydroxide are reacted with 10.0 g of carbon dioxide to produce calcium bicarbonate:

a. What is the limiting reactant?

b. How many grams of calcium bicarbonate will be produced?

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:
Percent Yield

- **Theoretical Yield:**
  The maximum amount of product possible in a chemical reaction for given quantities of reactants

- **Actual Yield:**
  The measured amount of product formed

**Percent Yield** = \( \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\% \)
Aluminum burns in bromine liquid, producing aluminum bromide. In one experiment, 6.0 g of aluminum reacted with an excess of bromine to yield 50.3 g aluminum bromide. Calculate the theoretical and percent yields.

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:
This concludes the Lecture PowerPoint presentation for Chapter 7

CHEMISTRY
an atoms-focused approach

GILBERT
KIRSS
FOSTER

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