Chapter 3: Stoichiometry

Mass, Formulas, and Reactions

Chapter Objectives:
• Learn how to use the atomic mass of an element and the molecular weight of a compound to relate grams, moles, and the number of formula units.
• Learn how to balance chemical equations.
• Learn how to use the mole concept to relate amounts of chemicals to each other (stoichiometry).
• Learn how to use percent compositions to find empirical and molecular formulas.
• Learn how to find the theoretical yield in limiting reactant problems.

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Chapter 3
Stoichiometry:

Counting By Weighing

• If the total mass of a sample of small objects is known, and the average mass of each small object is known, the number of objects in the sample can be determined.
• The same logic works for counting the number of atoms or molecules in a sample, but first we have to figure out how to weigh an atom.

The Mole Concept

• It is not possible to count the number of atoms or molecules involved in chemical reactions, since the molecules are so small, and so many are involved, even in a very small-scale reaction.
• Instead, it is necessary to measure amounts of molecules by using their mass.
• The relationship between sub-microscopic quantities like atoms and molecules, and macroscopic quantities like grams, is made using the mole concept.
• Using moles allows us to count particles by weighing them.

The Mole

• The mole (abbreviated mol) is the SI unit for amount of substance.
• A mole is defined as the amount of a substance that contains the same number of entities as there are atoms in exactly 12 g of carbon-12.
• 12 g of carbon-12 contains 6.022×10^23 atoms. This number is known as Avogadro’s number, N_A, in honor of Amedeo Avogadro (1776-1856, who first proposed the concept, and who also coined the word "molecule").

1 mole = 6.022×10^23 units (Avogadro’s number, N_A)

1 mol carbon-12 contains 6.022×10^23 atoms
1 mol H_2O contains 6.022×10^23 molecules
1 mol NaCl contains 6.022×10^23 formula units

The Molar Mass of an Element

• The molar mass (M or MM) of an element is the mass in grams of one mole of atoms of the element. It is numerically equal to the atomic mass of the element in amu’s:

molar mass in g/mol = atomic mass in amu

• 1 Fe atom has a mass of 55.847 amu.
1 mole of Fe atoms has a mass of 55.847 grams.

• 1 O atom has a mass of 15.9944 amu.
1 mole of O atoms has a mass of 15.9944 grams.

• 1 mole Al = 26.98 g Al = 6.022×10^23 atoms Al
• 1 mole He = 6.022×10^23 atoms He = 4.003 g He
**The Molar Mass of a Compound**

- The **molar mass of a compound** is the mass in grams of one mole of molecules or formula units of the compound. It is numerically equal to the **mass of the compound in amu’s**:  
  \[
  \text{molar mass in g/mol} = \text{sum of the atomic masses of the atoms in the molecule/formula unit}
  \]

  - For molecular compounds, this is often referred to as the **molecular mass**, and for ionic compounds, it is sometimes referred to as the **formula mass**.

  Molar mass \( \text{H}_2\text{O} \) = \(2 \times \text{atomic mass H} + (1 \times \text{atomic mass O})\)  
  \[= (2 \times 1.00794) + (1 \times 15.9994)\]  
  \[= 18.02\]  
- 1 \( \text{H}_2\text{O} \) molecule has a mass of 18.02 amu.  
- 1 mole of \( \text{H}_2\text{O} \) molecules has a mass of 18.02 grams.

**Relating Moles, amu’s and Grams**

- 1 \( \text{O}_2 \) molecule has a mass of 32.00 amu  
- 1 mole of \( \text{O}_2 \) has a mass of 32.00 g

- 1 \( \text{NaCl} \) formula unit has a mass of 58.44 amu  
- 1 mole of \( \text{NaCl} \) has a mass of 58.44 g

- 1 mole of \( \text{C}_6\text{H}_12\text{O}_6 \) = 180.16 g

- 1 mole of \( \text{Mg(C}_2\text{H}_3\text{O}_2)_2 \) = 83.35 g

**Just How Large is Avogadro’s Number?**

- How much is a mole of water molecules?

**Examples: Gram-Mole Conversions**

1. How many moles are present in 4.60 g of silicon?

   **Answer:** 0.164 mol Si

2. How many g of Si are present in 9.0 mol of Si?

   **Answer:** 250 g Si
Examples: Gram-Mole Conversions

3. How many atoms are in a sample of uranium with a mass of 1.000 µg?

Answer: $2.530 \times 10^{15}$ atoms U

4. How many atoms of carbon are in 2.5 mol of $C_2H_6O$?

Answer: $3.0 \times 10^{24}$ atoms C

5. A pure silver ring contains $2.80 \times 10^{22}$ silver atoms. How many grams of silver atoms does it contain?

Answer: 5.02 g Ag

6. How many moles of sucrose, $C_{12}H_{22}O_{11}$, are in a tablespoon of sugar that contains 2.85 g?

Answer: 0.00833 mol $C_{12}H_{22}O_{11}$

7. How many grams are in 0.0626 mol of NaHCO$_3$, the main ingredient in Alka-Seltzer tablets?

Answer: 5.26 g NaHCO$_3$

8. A sample of glucose, $C_6H_{12}O_6$, contains $1.52 \times 10^{25}$ molecules. How many kilograms of glucose is this?

Answer: 4.55 kg
Chapter 3: Stoichiometry

Chemical Equations

A Chemical Reaction Illustrated

Sodium (Na) + Chlorine (Cl₂) → Sodium Chloride (NaCl)

Balancing Chemical Reactions

• Equations are balanced by placing a stoichiometric coefficient in front of each species, indicating how many units of each compound participate in the reaction.
  – If no coefficient is present, it is assumed to be 1.
  – Usually, we use the smallest whole-number ratios for the coefficients.
  – **Never balance equations by changing subscripts!** This changes the identity of the species involved in the reaction!
  – In general, it’s a good idea to balance the atoms in the most complex substances first, and the atoms in the simpler substances last.

Examples: Balancing Reactions

1. \[ \_ \text{C(s)} + \_ \text{O}_2(g) \rightarrow \_ \text{CO}_2(g) \]
   \[ \_ \text{SO}_2(g) + \_ \text{O}_2(g) \rightarrow \_ \text{SO}_3(g) \]
   \[ \_ \text{Fe}_2\text{O}_3(s) + \_ \text{C(s)} \rightarrow \_ \text{Fe(s)} + \_ \text{CO}_2(g) \]
   \[ \_ \text{HCl(aq)} + \_ \text{CaCO}_3(s) \rightarrow \_ \text{CaCl}_2(aq) + \_ \text{H}_2\text{O(l)} + \_ \text{CO}_2(g) \]
   \[ \_ \text{N}_2(g) + \_ \text{O}_2(g) \rightarrow \_ \text{N}_2\text{O}_5(g) \]
   \[ \_ \text{Al(NO}_3)_3 + \_ \text{CaSO}_4 \rightarrow \_ \text{Al}_2(\text{SO}_4)_3 + \_ \text{Ca(NO}_3)_2 \]
Examples: Balancing Combustion Reactions

• In a combustion reaction, hydrocarbons (containing only H and C) react with molecular oxygen (O\textsubscript{2}) to produce carbon dioxide and water. (Incomplete combustion can result in other products, such as carbon monoxide and atomic carbon, or soot.)

2. \text{C}_4\text{H}_{10} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}

\text{C}_2\text{H}_6 + \text{O}_2 \rightarrow

\text{C}_3\text{H}_8 + \text{O}_2 \rightarrow

\text{C}_2\text{H}_5\text{OH} + \text{O}_2 \rightarrow

What Do the Coefficients Mean?

• Since moles combine in the same ratio that atoms or molecules do, the coefficients in a balanced chemical reaction specify the relative amounts in moles of the substances involved in the reaction.

Stoichiometry: Chemical Arithmetic

Greek: \textit{stoicheion} element or part + \textit{metron} measure

• Stoichiometry is the study of the numerical relationships in chemical formulas and reactions.

  – Knowing the stoichiometry of a formula allows us to relate moles and grams for particular reactants or products (e.g., that 1 mole of H\textsubscript{2}O weighs 18.02 g).

  – Knowing the stoichiometry of a reaction allows us to relate amounts of different substances to each other, using the mole ratios in the balanced equation, and allows us to predict how much of the products will be formed or how much of the reactants will be needed.

Reaction Stoichiometry: An Example

2 H\textsubscript{2} (g) + 1 O\textsubscript{2}(g) \rightarrow 2 H\textsubscript{2}O(g)

• Suppose we have 25.0 g of O\textsubscript{2}. How many grams of H\textsubscript{2} will be needed for this reaction? How many grams of H\textsubscript{2}O will be produced?

  – We can’t convert g O\textsubscript{2} \textit{directly} into g H\textsubscript{2}, but if we convert g O\textsubscript{2} into moles, we can use the \textit{coefficients of the balanced equation} to obtain moles of H\textsubscript{2}, and then convert to g H\textsubscript{2}.

  \begin{align*}
  \text{Balance chemical equation:} \\
  \text{Find appropriate mole ratio:} \\
  \text{Use mole ratio to convert:} \\
  \text{Convert to grams:}
  \end{align*}

  Mass of known substance \rightarrow Mole of known substance \rightarrow Mole of desired substance \rightarrow Convert to grams \rightarrow Mass of desired substance

  \text{Convert g O\textsubscript{2} to mol O\textsubscript{2}:}
  \begin{align*}
  25.0 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} &= 0.781 \text{ mol O}_2 \\

  \text{Convert mol O\textsubscript{2} to mol H\textsubscript{2}:}
  \begin{align*}
  0.781 \text{ mol } O_2 \times \frac{2 \text{ mol } H_2}{1 \text{ mol } O_2} &= 1.56 \text{ mol } H_2 \\

  \text{Convert mol H\textsubscript{2} to g H\textsubscript{2}:}
  \begin{align*}
  1.56 \text{ mol H}_2 \times \frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} &= 3.15 \text{ g H}_2
  \end{align*}
  \end{align*}
Reactions Stoichiometry: An Example

$$2 \text{H}_2 (g) + 1 \text{O}_2 (g) \rightarrow 2 \text{H}_2\text{O}(g)$$

Or we can put everything together:

$$25.0 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{2 \text{ mol H}_2}{1 \text{ mol O}_2} \times \frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} = 3.15 \text{ g H}_2$$

How many grams of H$_2$O will be formed?

$$25.0 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{2 \text{ mol H}_2}{1 \text{ mol O}_2} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 28.2 \text{ g H}_2\text{O}$$

Examples: Reaction Stoichiometry

2. In 2004, the world burned $3.0 \times 10^{10}$ barrels of petroleum, roughly equivalent to $3.4 \times 10^{15}$ g of gasoline (C$_8$H$_{18}$). How much CO$_2$ is released into the atmosphere from the combustion of this much gasoline?

$$2\text{C}_8\text{H}_{18}(l) + 25\text{O}_2(g) \rightarrow 16\text{CO}_2(g) + 18\text{H}_2\text{O}(g)$$

Answer: $1.0 \times 10^{16}$ g CO$_2$

Examples: Reaction Stoichiometry

3. Aqueous sodium hypochlorite (NaOCl), best known as household bleach, is prepared by reaction of sodium hydroxide with chlorine:

$$2\text{NaOH}(aq) + \text{Cl}_2(g) \rightarrow \text{NaOCl}(aq) + \text{NaCl}(aq) + \text{H}_2\text{O}$$

How many grams of NaOH are needed to react with 25.0 g of Cl$_2$?

Answer: 28.2 g NaOH

Examples: Reaction Stoichiometry

4. How many grams of Cl atoms are needed to combine with 24.4 g of Si atoms to make silicon tetrachloride, SiCl$_4$?

Answer: 123 g Cl

Examples: Reaction Stoichiometry

5. One of the most spectacular reactions of aluminum, the thermite reaction, is with iron(III) oxide, Fe$_2$O$_3$, by which metallic iron is made. So much heat is generated that the iron forms in the liquid state. The equation is:

$$2\text{Al}(s) + \text{Fe}_2\text{O}_3(s) \rightarrow \text{Al}_2\text{O}_3(s) + 2\text{Fe}(l)$$

A certain welding operation, used over and over, requires that each time at least 86.0 g of Fe be produced. (a) What is the minimum mass in grams of Fe$_2$O$_3$ that must be used for each operation? (b) How many grams of aluminum are also needed?

Answer: (a) 123 g Fe$_2$O$_3$; (b) 41.5 g Al
Chapter 3: Stoichiometry

Yields of Chemical Reactions

• In the examples we’ve seen, we have assumed that all of the reactions “go to completion” — that is, that all reactant molecules are converted into product molecules.

• In real life, some product is almost always lost due to:
  – too much heating
  – too little heating
  – klutzes
  – gremlins
  – evil spirits
  – evil co-workers
  – etc.

  – contamination in the glassware
  – impurities in the reactants
  – incomplete reactions
  – reactants evaporating into the air
  – side reactions that form other products

Yields of Chemical Reactions

• The theoretical yield is the amount that would be obtained if the reaction goes to completion (i.e., the maximum amount that could be made).

• The actual yield of a reaction is the amount that is actually obtained. (You could’ve guessed that.)

• The percent yield (% yield) is the actual yield expressed as a percentage of the theoretical yield:

\[
% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%
\]

Examples: Percent Yield

6. Methyl tert-butyl ether (MTBE, \(\text{C}_5\text{H}_{12}\text{O}\)), a substance used as an octane booster in gasoline, can be made by reaction of isobutylene (\(\text{C}_4\text{H}_8\)), with methanol (\(\text{CH}_3\text{OH}\)). What is the percent yield of the reaction if 32.8 g of MTBE is obtained from reaction of 26.3 g of isobutylene with sufficient methanol?

\[
\text{C}_4\text{H}_8(g) + \text{CH}_3\text{OH}(l) \rightarrow \text{C}_5\text{H}_{12}\text{O}(l)
\]

Answer: 79.4%

Limiting Reactants

• Whenever there is a reaction between more than one reactant, we can run out of one reactant before we run out of the other one.
  – The reactant we run out of first, which limits the yield of the entire reaction, is the limiting reactant (or limiting reagent).
  – The excess reactant is any reactant that is present in a larger amount than what is required to react completely with the limiting reactant.

• When we are given a reaction between two or more reactants, one may be completely consumed before the other(s). The reaction must stop at this point, leaving us with the remaining reactants in excess.

• The amount of this reactant, then, determines the maximum amount of the product(s) that can form, and is known as the limiting reactant.

• For example, suppose we were making standard 4-door cars, and we had the following (incomplete) list of “ingredients.” How many cars could we make?

<table>
<thead>
<tr>
<th>Engines</th>
<th>4</th>
</tr>
</thead>
<tbody>
<tr>
<td>Steer. wheels</td>
<td>4</td>
</tr>
<tr>
<td>Doors</td>
<td>15</td>
</tr>
<tr>
<td>Headlights</td>
<td>8</td>
</tr>
<tr>
<td>Seats</td>
<td>4</td>
</tr>
<tr>
<td>Rear-view mirrors</td>
<td>4</td>
</tr>
<tr>
<td>Windshield wipers</td>
<td>8</td>
</tr>
<tr>
<td>Wheels</td>
<td>11</td>
</tr>
</tbody>
</table>
**Limiting Reactants and Sundaes**

**Pizza recipe:**
1 crust + 5 oz. tomato sauce + 2 cups cheese → 1 pizza

If we have 4 crusts, 10 cups of cheese, and 15 oz. of tomato sauce, how many pizzas can we make?

- Tomato sauce is the limiting reagent, and the theoretical yield is 3 pizzas.

Next page:

**Limiting Reactants**

N\(_2\)(g) + 3H\(_2\)(g) → 2NH\(_3\)(g)  [Haber process]

• Suppose we mix 1.00 mol of N\(_2\) and 5.00 mol of H\(_2\). What is the maximum amount of NH\(_3\) that can be produced? How much H\(_2\) will be left over?

• Now suppose we mix 2.15 mol of N\(_2\) and 6.15 mol of H\(_2\). What is the theoretical yield of NH\(_3\)?

  **Assuming the N\(_2\) reacts completely, how much NH\(_3\) can be made?**

  \[
  2.15 \text{ mol } N_2 \times \frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2} = 4.30 \text{ mol } NH_3
  \]

  **Assuming the H\(_2\) reacts completely, how much NH\(_3\) can be made?**

  \[
  6.15 \text{ mol } H_2 \times \frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2} = 4.10 \text{ mol } NH_3
  \]

H\(_2\) is the limiting reactant; the theoretical yield of NH\(_3\) is 4.10 mol.

**Examples: Limiting Reactants**

1. Butane, C\(_4\)H\(_{10}\), undergoes combustion with oxygen, O\(_2\), to form carbon dioxide and water:

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

   Starting with 86.3 g C\(_4\)H\(_{10}\) and 100. g of O\(_2\) are mixed.
   a. Which of the two reactants is the limiting reagent, and how many grams of CO\(_2\) will be formed?
   b. How many grams of H\(_2\)O will be formed?
   c. How many grams of excess reagent are left over?
   d. If the actual yield of CO\(_2\) had been 75.0 g, what would be the percent yield of the reaction?

   **Answer:** (a) O\(_2\) limiting; 84.6 g CO\(_2\); (b) 43.3 g H\(_2\)O; (c) 72 g C\(_4\)H\(_{10}\); (d) 88.6%

2. Ammonia, NH\(_3\), can be synthesized by the following reaction:

   \[
   2\text{NO}(g) + 5\text{H}_2(g) \rightarrow 2\text{NH}_3(g) + 2\text{H}_2\text{O}(g)
   \]

   Starting with 86.3 g NO and 25.6 g H\(_2\), find the theoretical yield of ammonia in grams.

   **Answer:** NO limiting; 49.0 g NH\(_3\)

3. In a synthesis of phosphorus trichloride, a chemist mixed 12.0 g P with 35.0 g Cl\(_2\); she obtained 42.4 g of PCl\(_3\). What is the % yield of PCl\(_3\)?

   \[
   2\text{P}(s) + 3\text{Cl}_2(g) \rightarrow 2\text{PCl}_3(l)
   \]

   **Answer:** 93.8%
Percent Composition and Mass Percentage

- The **percent composition** of a compound is a list of the elements present in a substance listed by **mass percent**. Knowing the percent composition is often a first step to determining the formula of an unknown compound.

- The **mass percentage** (mass %) of an element in the compound is the portion of the compound’s mass contributed by that element, expressed as a percentage:

\[
\text{Mass\%\ of\ element\ } X = \frac{\text{atoms\ of\ } X\ \text{in\ formula} \times \text{molar mass of } X}{\text{molar mass of compound}} \times 100
\]

Examples: Mass Percentage

1. Glucose, or blood sugar, has the molecular formula \( \text{C}_6\text{H}_{12}\text{O}_6 \).
   a. What is the percent composition of glucose?
   b. How many grams of carbon are in 39.0 g of glucose (the amount of sugar in a typical soft drink)?

   **Answer:**
   a) 40.00% C, 6.714% H, 53.29% O
   b) 15.6 g C

Examples: Mass Percentage

2. The U.S. Food and Drug Administration (FDA) recommends that you consume less than 2.4 g of sodium per day. What mass of sodium chloride in grams can you consume and still be within the FDA guidelines?

   **Answer:** 6.1 g NaCl

Empirical Formula from Mass Percentage

- We can use the percent composition of a substance to find its **empirical and molecular formula**.

- If by some process we determine the percent composition of an unknown compound, we can convert this into a gram ratio by assuming that we have 100 g of the compound, and then to a mole ratio by using the atomic weights:

\[
\text{Sample: } 84.1\%\ C, 15.9\%\ H
\]

Assume 100 g of sample:

\[
\begin{align*}
84.1\text{g C} \times \frac{1\text{mol C}}{12.01115\text{g C}} &= 7.00\text{ mol C} \\
15.9\text{g H} \times \frac{1\text{mol H}}{1.00797\text{g H}} &= 15.8\text{ mol H}
\end{align*}
\]
Empirical Formula from Mass Percentage

- Since atoms combine in the same ratio that moles do, we divide all of the numbers of moles by the smallest number to put everything into lowest terms:

\[
\text{C}_2.26\text{H}_{1.00}\to \text{C}_1.00\text{H}_{2.26}
\]

• If the mole ratio is not all whole numbers, we multiply through by the smallest integer which will turn all of the numbers into integers. These numbers are the subscripts of the elements in the empirical formula.

\[
\left(\text{C}_1.00\text{H}_{2.26}\right)_4 \to \text{C}_4.00\text{H}_{8.94}
\]

Molecular Formula from Empirical Formula

- If we know the molar mass of the compound, we can obtain the molecular formula by dividing the weight of the empirical formula into the molar mass; this will determine the number of empirical formula units in the molecule.

\[
\frac{\text{molecular weight}}{\text{empirical formula weight}} = \frac{228.48\text{ g/mol}}{57.12\text{ g/mol}} = 4.00
\]

\[
\left(\text{C}_4\text{H}_{12}\right)_4 \to \text{C}_{16}\text{H}_{48}
\]

Examples: Empirical & Molecular Formulas

3. Vitamin C (ascorbic acid) contains 40.92% C, 4.58% H, and 54.50% O by mass. What is the empirical formula of ascorbic acid?

Answer: \(\text{C}_3\text{H}_4\text{O}_3\)

4. Spodumene, lithium aluminium inosilicate, is one of the most common lithium-containing minerals. It consists of 3.730% Li, 14.50% Al, 30.18% Si, and 51.59% O. What is the empirical formula of spodumene?

Answer: \(\text{LiAl}_3\text{Si}_2\text{O}_6\)

5. Black iron oxide is an ore containing iron and oxygen that occurs in magnetite. A 2.4480 g sample of the ore is found to contain 1.7714 g of iron. Calculate the empirical formula of black iron oxide.

Answer: \(\text{Fe}_3\text{O}_4\)

6. Styrofoam is a polymer made from the monomer styrene. Elemental analysis of styrene shows its percent composition to be 92.26% C and 7.75% H. Its molecular mass is found to be 104.15 g/mol. What are the empirical and molecular formulas of styrene?

Answer: empirical = CH, molecular = \(\text{C}_8\text{H}_8\)
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Examples: Empirical & Molecular Formulas
7. Butanedione is a main component in the smell and taste of butter and cheese. The empirical formula of butanedione is \( \text{C}_2\text{H}_3\text{O} \) and its molar mass is 86.09 g/mol. What is its molecular formula?

Answer: \( \text{C}_4\text{H}_6\text{O}_2 \)

Elemental / Combustion Analysis
• One common way of obtaining a chemical formula is by performing a combustion analysis (a specific type of elemental analysis).
• In this technique, an unknown sample is burned in pure \( \text{O}_2 \) (a combustion reaction), which converts all of the carbon atoms in the sample into \( \text{CO}_2 \) and all of the hydrogen atoms into \( \text{H}_2\text{O} \).

\[
\text{C, H, O} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}
\]

Figure 3.33

Elemental / Combustion Analysis
• The masses of \( \text{CO}_2 \) and \( \text{H}_2\text{O} \) are measured after the process is complete, and from this data, the amount of carbon and hydrogen in the original sample can be determined.
• Elements besides C and H must be determined by other methods; O is usually found by difference.

\[
g\text{CO}_2 \rightarrow \text{mol CO}_2 \rightarrow \text{mol C} \rightarrow g \text{ C in sample} \rightarrow \% C
g\text{H}_2\text{O} \rightarrow \text{mol H}_2\text{O} \rightarrow \text{mol H} \rightarrow g \text{ H in sample} \rightarrow \% H
\]

\[
\% \text{ O} = 100\% - (\% \text{ C} + \% \text{ H})
\]

Examples: Combustion Analysis
8. A sample of an unknown compound with a mass of 0.5438 g is burned in a combustion analysis. The mass of \( \text{CO}_2 \) produced was 1.039 g and the mass of \( \text{H}_2\text{O} \) was 0.6369 g. What is the empirical formula of the compound?

Answer: \( \text{C}_3\text{H}_6\text{O} \)

Molecules and Isomers
• Even knowing the empirical or molecular formulas of a compound does not necessarily tell us what that compound actually is.
• We’ve already seen that the empirical formula only tells us about the relative numbers of atoms present within the formula unit or molecule.
• Many different compounds can have the same empirical formula. For instance, there are dozens of different compounds that have the empirical formula \( \text{CH}_2\text{O} \).

– Notice that in on the following slide, there is no relationship between the structure and how many ‘\( \text{CH}_2\text{O} \)’ units the molecule contains.

Some Compounds with Empirical Formula \( \text{CH}_2\text{O} \)
• Composition by mass

\begin{tabular}{|c|c|c|c|c|}
\hline
Name & Molecular Formula & No. of \( \text{CH}_2\text{O} \) Units & Molar Mass (g/mol) & Function \\
\hline
Formaldehyde & \( \text{CH}_2\text{O} \) & 1 & 30.05 & Disinfectant; biological preservative \\
Acetic acid & \( \text{C}_2\text{H}_4\text{O}_2 \) & 2 & 60.05 & Vinegar (5% solution); acetic polymers \\
Lactic acid & \( \text{C}_3\text{H}_6\text{O}_3 \) & 3 & 90.05 & Found in sour milk and sourdough bread; forms in muscles during exercise \\
Erythrose & \( \text{C}_4\text{H}_8\text{O}_4 \) & 4 & 120.10 & Forms during sugar metabolism \\
Ribose & \( \text{C}_5\text{H}_10\text{O}_5 \) & 5 & 150.15 & Component of ribonucleic acid (RNA); found in vitamin \( \text{B}_1 \) \\
Glucose & \( \text{C}_6\text{H}_12\text{O}_6 \) & 6 & 180.16 & Major nutrient for energy in cells \\
\hline
\end{tabular}
**Structural Isomers**

- Even compounds that have the same molecular formula can have the atoms connected in a different order — these are **structural isomers**.

<table>
<thead>
<tr>
<th></th>
<th>Ethanol</th>
<th>Dimethyl Ether</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Molecular Formula</strong></td>
<td>C₂H₆O</td>
<td>C₂H₆O</td>
</tr>
<tr>
<td><strong>Molar Mass (g/mol)</strong></td>
<td>46.07</td>
<td>46.07</td>
</tr>
<tr>
<td><strong>Appearance</strong></td>
<td>Colorless liquid</td>
<td>Colorless gas</td>
</tr>
<tr>
<td><strong>Melting point</strong></td>
<td>-117°C</td>
<td>-139°C</td>
</tr>
<tr>
<td><strong>Boiling point</strong></td>
<td>78.5°C</td>
<td>-25°C</td>
</tr>
<tr>
<td><strong>Density (at 20°C)</strong></td>
<td>0.789 g/mL</td>
<td>0.00195 g/mL</td>
</tr>
<tr>
<td><strong>Function</strong></td>
<td>Intoxicant</td>
<td>Refrigerant</td>
</tr>
</tbody>
</table>