Ch 6. Chemical Composition

Chapter Goals

1. To understand the mole concept
2. Understanding Avogadro’s number and using it to calculate moles
3. Calculating Molar mass
4. Calculating Mass percents
5. Determining Empirical Formula’s from mass percent information
6. Distinguish between Empirical and Molecular Formula’s

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Chemical Packages—Moles

1 mole (of something) = 6.022 X 10^{23} particles

1 mole of pennies = 6.022 X 10^{23} pennies

How Chemists use moles:
1 mole = mass in grams of one smallest unit of that element OR compound

Example:
Molecular weight of water = 18 g (rounded)
Q: How is that related to moles?
A: 18 g of water has 6.022 X 10^{23} of water molecules.

The number 6.022 X 10^{23} is Avogadro’s number!
Moles, Definition

• Mole = Number of things equal to the number of atoms in 12 g of C-12.
  – 1 atom of C-12 weighs exactly 12 amu.
  – 1 mole of C-12 weighs exactly 12 g.

• In 12 g of C-12 there are $6.022 \times 10^{23}$ C-12 atoms.

Moles, Contd.

Same amount but different masses.

The two types of nails will have two different masses even if both are 1 dozen in number.

Similarly, 1 mole (same amount) of different elements or compounds will have different masses.

How many grams are in 1.000 mol of iron?
Relationship Between Moles and Mass

The mass of one mole of atoms is called the **molar mass**.
The molar mass of an element, in grams, is numerically equal to the element’s atomic mass, in amu.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Pieces in 1 mole</th>
<th>Weight of 1 mole</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>$6.022 \times 10^{23}$ atoms</td>
<td>1.008 g</td>
</tr>
<tr>
<td>Carbon</td>
<td>$6.022 \times 10^{23}$ atoms</td>
<td>12.01 g</td>
</tr>
<tr>
<td>Oxygen</td>
<td>$6.022 \times 10^{23}$ atoms</td>
<td>16.00 g</td>
</tr>
<tr>
<td>Sulfur</td>
<td>$6.022 \times 10^{23}$ atoms</td>
<td>32.06 g</td>
</tr>
<tr>
<td>Calcium</td>
<td>$6.022 \times 10^{23}$ atoms</td>
<td>40.08 g</td>
</tr>
<tr>
<td>Chlorine</td>
<td>$6.022 \times 10^{23}$ atoms</td>
<td>35.45 g</td>
</tr>
<tr>
<td>Copper</td>
<td>$6.022 \times 10^{23}$ atoms</td>
<td>63.55 g</td>
</tr>
</tbody>
</table>

What is the molar mass of sucrose, $C_{12}H_{22}O_{11}$?

What is the mass of $4.55 \times 10^{23}$ copper atoms?
Mole Relationships in Chemical Formulas

<table>
<thead>
<tr>
<th>Moles of compound</th>
<th>Moles of constituents</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 mol NaCl</td>
<td>1 mol Na, 1 mol Cl</td>
</tr>
<tr>
<td>1 mol H₂O</td>
<td>2 mol H, 1 mol O</td>
</tr>
<tr>
<td>1 mol CaCO₃</td>
<td>1 mol Ca, 1 mol C, 3 mol O</td>
</tr>
<tr>
<td>1 mol C₆H₁₂O₆</td>
<td>6 mol C, 12 mol H, 6 mol O</td>
</tr>
</tbody>
</table>

*How many grams of oxygen are in 40.1 g of calcium hydroxide?

Moles, Contd.

Practice: A pure raindrop contains 0.05 g of water.

Calculate:

How many moles(mol) of water?

How many molecules of water?

How many atoms of Hydrogen?

How many atoms of Oxygen In the raindrop?
Percent Composition

- Means percentage of each element in a compound. By mass.
- Can be determined from:
  - The formula of the compound.
  - An experimental mass analysis of the compound.
- The percentages may not always total to 100% due to rounding.

\[
\text{Percentage} = \frac{\text{mass of element } X \text{ in } 1 \text{ mol}}{\text{mass of } 1 \text{ mol of the compound}} \times 100\%
\]

Mass Percent as a Conversion Factor

- The mass percent tells you the mass of a constituent element in 100 g of the compound.
  - The fact that NaCl is 39% Na by mass means that 100 g of NaCl contains 39 g Na.
- This can be used as a conversion factor.
  - 100 g NaCl \equiv 39 g Na

*What is the mass percent of oxygen in caffeine, \(C_8H_{10}N_4O_2\)?*
Empirical Formulas

- The simplest, whole-number ratio of atoms in a molecule is called the **empirical formula**.
  - Can be determined from percent composition or combining masses.
- The molecular formula is a multiple of the empirical formula.

Empirical Formulas, Continued

**Hydrogen Peroxide**
Molecular formula = \( \text{H}_2\text{O}_2 \)
Empirical formula = \( \text{HO} \)

**Benzene**
Molecular formula = \( \text{C}_6\text{H}_6 \)
Empirical formula = \( \text{CH} \)

**Glucose**
Molecular formula = \( \text{C}_6\text{H}_{12}\text{O}_6 \)
Empirical formula = \( \text{CH}_2\text{O} \)
Finding an Empirical Formula

1. Convert the percentages to grams.
   a. Skip if already grams.
2. Convert grams to moles.
   a. Use molar mass of each element.
3. Write a pseudoformula using moles as subscripts.
4. Divide all by smallest number of moles.
5. Multiply all mole ratios by number to make all whole numbers, if necessary.

What is the empirical formula of ascorbic acid, C₆H₈O₆, a compound found in citrus fruits?

All These Molecules Have the Same Empirical Formula. How Are the Molecules Different?

<table>
<thead>
<tr>
<th>Name</th>
<th>Molecular Formula</th>
<th>Empirical Formula</th>
<th>Molar mass?</th>
</tr>
</thead>
<tbody>
<tr>
<td>Glyceraldehyde</td>
<td>C₃H₆O₃</td>
<td>CH₂O</td>
<td></td>
</tr>
<tr>
<td>Erythrose</td>
<td>C₄H₈O₄</td>
<td>CH₂O</td>
<td></td>
</tr>
<tr>
<td>Arabinose</td>
<td>C₅H₁₀O₅</td>
<td>CH₂O</td>
<td></td>
</tr>
<tr>
<td>Glucose</td>
<td>C₆H₁₂O₆</td>
<td>CH₂O</td>
<td></td>
</tr>
</tbody>
</table>
Calculating Empirical Formula’s from Analyzed and Reaction Data

From analyzed data:
6.90 Determine the empirical formula of Naproxen (Aleve) from the following data obtained from a lab analysis,

\[ \text{C } 73.03\%, \text{ H } 6.13\%, \text{ O } 20.84\% \]

\begin{itemize}
  \item C = black
  \item H = grey
  \item O = red
\end{itemize}

6.91 from reaction data
A 1.45 g sample of phosphorus burns in air to give 2.57 g of the Phosphorus oxide. Determine the empirical formula of the oxide.

Molecular Formulas for Compounds

- The molecular formula is a multiple of the empirical formula.
- To determine the molecular formula, you need to know the empirical formula and the molar mass of the compound.

A bar of Hershey’s Special Dark chocolate contains 31 mg of caffeine per serving. Caffeine has the following mass% composition:

\[ \text{C } 49.48\%, \text{ H } 5.19\%, \text{ N } 28.85\%, \text{ O } 16.48\% \]

molar mass is 194.19 g/mol. Find the molecular formula of caffeine.

\begin{itemize}
  \item C = black, N = blue, O = red
  \item H = grey
\end{itemize}
Because of increasing evidence of damage in the ozone layer, CFC Production was banned in 1996.

A leak in the air conditioning system of an older car releases 55 g of CF₂Cl₂ per month. How much Cl is emitted each year into the atmosphere?

Use:

\[
g \text{CF}_2\text{Cl}_2 \rightarrow \text{mol} \text{ CF}_2\text{Cl}_2 \rightarrow \text{mol} \text{ Cl} \rightarrow g \text{ Cl}
\]

There are still about 100 million air conditioners being used that need to be recharged using CFC-12.

Hydrogen, a possible future fuel can be obtained from sources such as water, ethanol etc. How much hydrogen in grams, can be obtained from 1.0 L of water? (density of water = 1.0 g/cc)

\[
g \text{H}_2\text{O} \rightarrow \text{mole} \text{ H}_2\text{O} \rightarrow \text{mole} \text{ H} \rightarrow g \text{ H}
\]

- Use molar mass
- mol compound to mol element
- Use molar mass
The Chapter in Summary

1. Calculating g to moles and moles to g using molar mass
2. Calculating number atoms or molecules using Na
3. Calculating mass percent composition

\[
\text{Mass percent of } X = \frac{\text{Mass of } X \text{ in 1 mol of compound}}{\text{Mass of 1 mol of the compound}} \times 100\%
\]

4. Calculating empirical formula’s from experimental data

5. Calculating molecular formula’s using empirical formula and molar mass