Molecular Mass
Molecular Mass

synonymous with molar mass and molecular weight

is the sum of the atomic masses of all the atoms in a molecule

the mass in grams of one mole of a compound
not all compounds are molecular
calculated exactly
the same way as
molecular mass

Solid structure of NaCl
Determine the molar mass of Sr(NO₃)₂.

1 mol Sr = 87.62 g
2 mol N = 2(14.00 g)
6 mol O = 6(16.00 g)

\[
1 \text{ mol Sr(NO}_3\text{)}_2 = 211.62 \text{ g}
\]
pg. 323 problem 27

What is the mass in grams of 3.25 mol sulfuric acid (H\textsubscript{2}SO\textsubscript{4}).

1 mol S = 32.07 g
2 mol H = 2(1.00 g)
4 mol O = 4(16.00 g)

\[
\text{1 mol H}_2\text{SO}_4 = 98.07 \text{ g}
\]

\[
3.25 \text{ mol H}_2\text{SO}_4 \times \frac{98.07 \text{ g H}_2\text{SO}_4}{1 \text{ mol H}_2\text{SO}_4} = 319 \text{ g H}_2\text{SO}_4
\]
Determine the number of moles present in 22.6 g AgNO₃.

\[
\begin{align*}
1 \text{ mol Ag} &= 107.87 \text{ g} \\
1 \text{ mol N} &= 14.00 \text{ g} \\
3 \text{ mol O} &= 3(16.00 \text{ g}) \\
\text{1 mol AgNO}_3 &= 169.87 \text{ g}
\end{align*}
\]

\[
\begin{align*}
22.6 \text{ g AgNO}_3 &\times \frac{1 \text{ mol AgNO}_3}{169.87 \text{ g AgNO}_3} = 0.133 \text{ mol AgNO}_3
\end{align*}
\]
A sample of silver chromate (Ag₂CrO₄) has a mass of 25.8 g.

a. How many Ag⁺ ions are present?

\[
\begin{align*}
25.8 \text{ g Ag}_2\text{CrO}_4 & \times \frac{1 \text{ mol Ag}_2\text{CrO}_4}{331.74 \text{ g Ag}_2\text{CrO}_4} \times \frac{2 \text{ mol Ag}^+}{1 \text{ mol Ag}_2\text{CrO}_4} \\
& \times \frac{6.02 \times 10^{23} \text{ ions Ag}^+}{1 \text{ mol Ag}^+} \\
& = 9.36 \times 10^{22} \text{ ions Ag}^+
\end{align*}
\]
A sample of silver chromate (Ag$_2$CrO$_4$) has a mass of 25.8 g.

b. How many CrO$_4^{2-}$ ions are present?

\[
\frac{25.8 \text{ g Ag}_2\text{CrO}_4}{1 \text{ mol Ag}_2\text{CrO}_4} \times \frac{1 \text{ mol Ag}_2\text{CrO}_4}{331.74 \text{ g Ag}_2\text{CrO}_4} \times \frac{1 \text{ mol CrO}_4^{2-}}{1 \text{ mol Ag}_2\text{CrO}_4} \times \frac{6.02 \times 10^{23} \text{ ions CrO}_4^{2-}}{1 \text{ mol CrO}_4^{2-}} = 4.68 \times 10^{22} \text{ atoms CrO}_4^{2-}
\]
c. What is the mass in grams of one formula unit of silver chromate (Ag₂CrO₄).

\[
\frac{331.74 \text{ g Ag₂CrO₄}}{1 \text{ mol Ag₂CrO₄}} \times \frac{1/\text{mol Ag₂CrO₄}}{6.02 \times 10^{23} \text{ f.u. Ag₂CrO₄}} = \frac{5.51 \times 10^{-22} \text{ g Ag₂CrO₄}}{\text{f.u. Ag₂CrO₄}}
\]

- 2 mol Ag = 107.87 g
- 1 mol Cr = 52.00 g
- 4 mol O = 4(16.00 g)
- 1 mol Ag₂CrO₄ = 331.74 g
Percent Composition of Compounds
Percent composition is the percent by mass of each element the compound contains.

Obtained by dividing the mass of each element in one mole of the compound by the molar mass of the compound and multiplying by 100%
Example 45

Calculate the percent composition by mass of H, P and O for one mole of phosphoric acid (H₃PO₄)

Molar mass = 3(1.008g) + 30.97g + 4(16.00)  
= 97.99
Molar mass = 3(1.008g) + 30.97g + 4(16.00) = 97.99

$$\% \text{H} = \frac{3(1.008g)}{97.99g} \times 100\% = 3.086\%$$

$$\% \text{P} = \frac{30.97g}{97.99g} \times 100\% = 31.61\%$$

$$\% \text{O} = \frac{4(16.00)}{97.99g} \times 100\% = 65.31\%$$
Example

A sample of a compound containing carbon and oxygen had a mass of 88g. Of this sample 24g was carbon, 64g was oxygen. What is the percent composition of this compound.

\[
\text{% carbon} = \frac{24\text{g}}{88\text{g}} \times 100\% = 27\%
\]

\[
\text{% oxygen} = \frac{64\text{g}}{88\text{g}} \times 100\% = 73\%
\]
Determining Formula
Levels of Structure

Elemental Composition
Empirical Formula
Molecular Formula
Constitution
Configuration
Conformation
# Elemental Composition

**Examples:**

<table>
<thead>
<tr>
<th>Formaldehyde</th>
<th>Glucose</th>
</tr>
</thead>
<tbody>
<tr>
<td>C: 40.00%</td>
<td>C: 40.00%</td>
</tr>
<tr>
<td>H: 6.73%</td>
<td>H: 6.73%</td>
</tr>
<tr>
<td>O: 53.27%</td>
<td>O: 53.27%</td>
</tr>
</tbody>
</table>
Levels of Structure

Elemental Composition ✓

Empirical Formula

Molecular Formula

Constitution

Configuration

Conformation
Empirical Formula

The empirical formula tells us which elements are present and the simplest whole-number ratio of their atoms.
Empirical Formula

Examples: Formaldehyde and Glucose

Elemental Composition

C: 40.00%  40.00 g
H: 6.73%   6.73 g
O: 53.27%  53.27 g

assume a 100g sample

calculate atom ratios by dividing by atomic weight
Calculating Empirical Formula

\[
\begin{align*}
\text{C: } & \quad 40.00 \text{ g} \times \frac{1 \text{ mol}}{12.01 \text{ g}} = 3.33 \text{ mol} \\
\text{H: } & \quad 6.73 \text{ g} \times \frac{1 \text{ mol}}{1.00 \text{ g}} = 6.73 \text{ mol} \\
\text{O: } & \quad 53.27 \text{ g} \times \frac{1 \text{ mol}}{16.0 \text{ g}} = 3.33 \text{ mol}
\end{align*}
\]
Empirical Formula

Examples: Formaldehyde and Glucose

Elemental Composition

C: 40.00% 40.00 g 3.33 mol
H: 6.73%  6.73 g 6.73 mol
O: 53.27% 53.27 g 3.33 mol

assume a 100g sample

calculate atom ratios by dividing by atomic weight

determine the smallest whole number ratio by dividing by the smallest molar value
Calculating Empirical Formula

\[ \text{C: } 40.00 \text{ g} \times \frac{1 \text{ mol}}{12.01 \text{ g}} = \frac{3.33 \text{ mol}}{3.33 \text{ mol}} = 1.00 \]

\[ \text{H: } 6.73 \text{ g} \times \frac{1 \text{ mol}}{1.00 \text{ g}} = \frac{6.73 \text{ mol}}{3.33 \text{ mol}} = 2.02 \]

\[ \text{O: } 53.27 \text{ g} \times \frac{1 \text{ mol}}{16.0 \text{ g}} = \frac{3.33 \text{ mol}}{3.33 \text{ mol}} = 1.00 \]
Empirical Formula

Examples: Formaldehyde and Glucose

Elemental Composition

- **C**: 40.00% 40.00 g 3.33 mol 1
- **H**: 6.73% 6.73 g 6.73 mol 2
- **O**: 53.27% 53.27 g 3.33 mol 1

Empirical Formula: CH₂O
A 1.723 g sample of aluminum oxide (which consists of aluminum and oxygen only) contains 0.912 g of Al. Determine the empirical formula of the compound.

1.723 g sample - 0.912 g Al = 0.811 g O

\[
\begin{align*}
0.912 \text{ g Al} & \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = 0.0338 \text{ mol} \\
0.811 \text{ g O} & \times \frac{1 \text{ mol O}}{16.0 \text{ g O}} = 0.0507 \text{ mol}
\end{align*}
\]

\[
\begin{align*}
0.0338 \text{ mol Al} & = 1.0 \times 2 = 2 \\
0.0338 \text{ mol O} & = 1.5 \times 2 = 3
\end{align*}
\]

Therefore, the empirical formula of the compound is \(\text{Al}_2\text{O}_3\).
Levels of Structure

- Elemental Composition
- Empirical Formula
- Molecular Formula
- Constitution
- Configuration
- Conformation
Molecular Formula

determined from empirical formula and experimentally determined molecular mass

<table>
<thead>
<tr>
<th>Compound</th>
<th>Empirical Formula</th>
<th>Molar mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>formaldehyde</td>
<td>CH$_2$O</td>
<td>30</td>
</tr>
<tr>
<td>glucose</td>
<td>CH$_2$O</td>
<td>180</td>
</tr>
</tbody>
</table>
Calculation of empirical mass

1 mol C = 12.01 g
2 mol H = 2 x 1.016 g
1 mol O = 16.00 g

\[
\text{1 mol C} = 12.01 \text{ g} \\
\text{2 mol H} = 2 \times 1.016 \text{ g} \\
\text{1 mol O} = 16.00 \text{ g}
\]

\[
30.026 \text{ g}
\]

Molecular mass

\[
\frac{180 \text{ g}}{30 \text{ g}} = 6
\]

Empirical mass

\[
\frac{30 \text{ g}}{30 \text{ g}} = 1
\]

\[
\text{glucose} \\
\text{formaldehyde}
\]
Molecular Formula

determined from empirical formula and experimentally determined molecular mass

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<td>CH₃O</td>
</tr>
<tr>
<td>glucose</td>
<td>CH₂O</td>
<td>180</td>
<td>C₆H₁₂O₆</td>
</tr>
</tbody>
</table>
Elemental Composition

Example:

Lysine

C: 49.20%
H: 9.66%
N: 19.20%
O: 21.94%
Levels of Structure

Elemental Composition ✓

Empirical Formula

Molecular Formula

Constitution

Configuration

Conformation
Elemental Formula

Example:

Lysine

- C: 49.20%  49.20 g
- H: 9.66%  9.66 g
- N: 19.20%  19.20 g
- O: 21.94%  21.94 g

assume a 100-g sample

calculate atom ratios by dividing by atomic weight

determine smallest whole-number ratio
by dividing by smallest number
49.20 g C \times \frac{1 \text{ mol}}{12 \text{ g}} = 4.10 \frac{\text{mol}}{1.37 \text{ mol}} = 3 \\
9.66 \text{ g H} \times \frac{1 \text{ mol}}{1.0 \text{ g}} = 9.58 \frac{\text{mol}}{1.37 \text{ mol}} = 7 \\
19.20 \text{ g N} \times \frac{1 \text{ mol}}{14.0 \text{ g}} = 1.37 \frac{\text{mol}}{1.37 \text{ mol}} = 1 \\
21.94 \text{ g O} \times \frac{1 \text{ mol}}{16.0 \text{ g}} = 1.37 \frac{\text{mol}}{1.37 \text{ mol}} = 1
Elemental Formula (cont’d)

Example:

Lysine

\[
\begin{align*}
C & : 4.10 \text{ mol C atoms} & 3 \\
H & : 9.58 \text{ mol H atoms} & 7 \\
N & : 1.37 \text{ mol N atoms} & 1 \\
O & : 1.37 \text{ mol O atoms} & 1 \\
\end{align*}
\]

determine smallest whole-number ratio by dividing by smallest number (1.37 mol)

\[C_3H_7ON\]
Levels of Structure

Elemental Composition ✓

Empirical Formula ✓

Molecular Formula

Constitution

Configuration

Conformation
determined from empirical formula and molar mass

<table>
<thead>
<tr>
<th>Compound</th>
<th>Empirical Formula</th>
<th>Molar mass</th>
<th>Molecular formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>lysine</td>
<td>C₃H₇ON</td>
<td>~150</td>
<td>C₆H₁₄O₂N₂</td>
</tr>
</tbody>
</table>

\[
\frac{150g}{73g} = 2
\]
Write the empirical formulas for the following molecules: (a) acetylene (C₂H₂), (b) dinitrogen tetroxide (N₂O₄), (c) glucose (C₆H₁₂O₆), diiodine pentoxide (I₂O₅).

This problem is not realistic. Molecular formulas are derived from empirical formulas, not vice versa. Empirical formulas come from experiment.
Example

Calculate the number of moles of chloroform (CHCl$_3$) in 198 g of chloroform.

Molecular mass of chloroform:
1 mol C = 12.01 g
1 mol H = 1.008 g
3 mol Cl = 3(35.46 g) = 106.38 g

1 mol CHCl$_3$ = 119.4 g

\[
\frac{198 \text{ g CHCl}_3}{119.4 \text{ g CHCl}_3} \times 1 \text{ mol CHCl}_3 = 1.66 \text{ mol CHCl}_3
\]
Example

Caffeine contains 49.48% carbon, 5.15% hydrogen, 28.87% nitrogen, and 16.49% oxygen by mass and has a molar mass of 194.2g. Determine the molecular formula formula of caffeine
First determine the mass of each element in one mole of caffeine

\[
\begin{array}{ccc}
49.48\text{g C} & \times & 194.2\text{g caffeine} \\
100\text{g caffeine} & & 1 \text{ mol} \\
\hline
96.09\text{g C} \\
1 \text{ mol caffeine} \\
\end{array}
\]

\[
\begin{array}{ccc}
5.15\text{g H} & \times & 194.2\text{g caffeine} \\
100\text{g caffeine} & & 1 \text{ mol} \\
\hline
10.0\text{g H} \\
1 \text{ mol caffeine} \\
\end{array}
\]

\[
\begin{array}{ccc}
28.87\text{g N} & \times & 194.2\text{g caffeine} \\
100\text{g caffeine} & & 1 \text{ mol} \\
\hline
56.07\text{g N} \\
1 \text{ mol caffeine} \\
\end{array}
\]

\[
\begin{array}{ccc}
16.49\text{g O} & \times & 194.2\text{g caffeine} \\
100\text{g caffeine} & & 1 \text{ mol} \\
\hline
32.02\text{g O} \\
1 \text{ mol caffeine} \\
\end{array}
\]
Example

then convert to moles

\[
\begin{align*}
96.09\text{g C} & \quad \times \quad \frac{1 \text{ mol C}}{12.011\text{g C}} = 8.00 \text{ mol C} \\
10.0\text{g H} & \quad \times \quad \frac{1 \text{ mol H}}{1.008\text{g H}} = 9.92 \text{ mol H} \\
56.07\text{g N} & \quad \times \quad \frac{1 \text{ mol N}}{14.01\text{g N}} = 4.00 \text{ mol N} \\
32.02\text{g O} & \quad \times \quad \frac{1 \text{ mol O}}{16.00\text{g O}} = 2.00 \text{ mol O}
\end{align*}
\]

\(\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2\)