Chapter 7 Goals

Major Goals of Chapter 7:
1. Counting - by numbers, by weighing and by volume.
2. Summing atomic masses for elements in correctly written chemical formulas.
3. Recognizing the difference between a formula mass (amu) & atomic mass (g/mol).
4. Converting grams to moles & Converting moles to grams.
5. Calculating percentage composition need in determining an empirical formula.

Before viewing, read the Chapter 7 Review:
7.1 Atomic Mass and Formula Mass
7.2 The Mole
7.3 Molar Mass
7.4 Calculations Using Molar Mass
7.5 Percent Composition and Empirical Formulas
7.6 Molecular Formulas
Chapter 7 summary

Physical Methods for counting amounts

- by numbers
  - 1 pr = 2
  - 1 doz = 12
  - 1 case = 12 or 24
  - 1 hand = 5 fingers
  - 1 foot = 12 inches

- by weighing
  - 1 gummy bear = 1 gram
  - 1 penny = 1 gram
  - 1 proton = 1 amu
  - 1 neutron = 1 amu
  - 1 hydrogen-1 atom = 1 amu
  - 1 carbon-12 atom = 12 amu

- Avogadro’s number
  - 1 mole = $6.02 \times 10^{23}$

- molar masses defined

Chemical Methods for counting amounts

- by volume
  - 1 mole any gas = 22.4 L @ 0°C, 1 atm

- by balanced reaction

\[ 2 \text{ H}_2\text{O} \rightarrow 2 \text{ H}_2 + 1 \text{ O}_2 \]

- multiply through by A number
  - 2 x A number
  - 2 x A number
  - 1 x A number

- A number equals Avogadro’s #
  - $2 \times 6.02 \times 10^{23}$
  - $2 \times 6.02 \times 10^{23}$
  - $1 \times 6.02 \times 10^{23}$

- Avogadro’s # equals 1 mole
  - 2 x 1 mole
  - 2 x 1 mole
  - 1 x 1 mole

- count by weighing

- 2 x 18 grams H\textsubscript{2}O per 1 mole H\textsubscript{2}O
- 2 x 2 grams H\textsubscript{2} per 1 mole H\textsubscript{2}
- 1 x 32 grams O\textsubscript{2} per 1 mole O\textsubscript{2}

the electrolysis of 36 grams of water will produce

4 grams of hydrogen and 32 grams oxygen
In chemistry there are two general methods to count particles:

A. Physical Methods
   - for counting amounts
   - by numbers
   - by weighing
   - by volume

Chapter 7 deals with the physical measuring of amounts
- mole, and the mole mass ratio.

B. Chemical Methods
   - for counting amounts
   - by a balanced reaction

Chapter 8 deals with the chemical measuring of amounts
- balancing a chemical reaction
Section 7.2 - The Mole (amounts)

Counting by numbers:
• one pair of anything is the same as two of anything.
• one dozen of anything is the same as twelve of anything.
• one case of wine is the same as twelve bottles of wine.
• one case of soda is the same as twenty four cans of soda.
• one hand is the same as five fingers.
• one foot is the same as twelve inches.

The mole is a counting device like those given above:
• one mole of anything is the same as $6.023 \times 10^{23}$ anything
• one mole of anything equals Avogadro’s number of anything
Section 7.2 - The Mole (amounts)

Counting by weighing using a mass to amount ratio:
• one gummy bear weighs the same as one gram
• one penny weighs the same as one gram
• one proton weighs the same as one amu (atomic mass unit)
• one neutron weighs the same as one amu (atomic mass unit)
• one hydrogen-1 atom weighs the same as one amu (exactly)
• one carbon-12 atom weighs the same as twelve amu (exactly)

An atomic mass unit has a mass of $1.67 \times 10^{-24}$ grams (rounded)

\[
12 \text{ pennies} \times \frac{1 \text{ gram}}{1 \text{ penny}} = 12 \text{ grams}
\]

\[
12 \text{ amu} \times \frac{1.67 \times 10^{-24} \text{ grams}}{1 \text{ amu}} = 2.00 \times 10^{-25} \text{ grams}
\]
COUNTING

Physical Methods for counting amounts

by volume

1 mole any gas = 22.4 L @ 0°C, 1 atm

Counting by volume using a volume to amount ratio:

• one mole of any gas is the same as 22.4 liters @ 0°C and 1 atm.

• 1 mole of any gas @ 0°C and 1 atm contains 6.023x10^{23} particles.

• EQUAL volumes of any gas contain EQUAL amounts of particles.

Section 7.2 - The Mole (amounts)
Section 7.1 - Atomic Mass and Formula Mass (mass to amount ratio)

Counting by weighing using a mass to amount ratio
A. A mass to amount (grams per moles) ratio is called a molar mass

Avogadro's number
1 mole = 6.02x10^{23}
anything anything

1 mole carbon-12 = 12 grams carbon-12 exactly
(6.02x10^{23}) (6.02x10^{23})

For HCl, 1 H = 1.007 amu
1Cl = 35.45 amu
1 mole HCl = 36.457 grams

\[ \frac{12.0 \text{ mole C}}{1 \text{ mole C}} \times \frac{12.01 \text{ grams C}}{1 \text{ mole C}} = 144 \text{ grams C} \]

\[ \frac{12.0 \text{ moles HCl}}{1 \text{ mole HCl}} \times \frac{36.457 \text{ grams HCl}}{1 \text{ mole HCl}} = 437 \text{ grams HCl} \]
Section 7.1- Atomic Mass and Formula Mass (mass to amount ratio)

Summary:
Be sure to know that “moles of substance” multiplied by a molar mass ratio equals grams.

\[
12.0 \text{ mole C} \times \frac{12.01 \text{ grams C}}{1 \text{ mole C}} = 144 \text{ grams C}
\]

\[
12.0 \text{ moles HCl} \times \frac{36.457 \text{ grams HCl}}{1 \text{ mole HCl}} = 437 \text{ grams HCl}
\]
I. Counting Devices

A. Familiar counting devices for counting physical objects larger than atoms:

1 dozen of _______________ = _______________ units of anything

A basket has 18 eggs. How many dozen eggs are contained in the basket?

\[ 18 \text{ eggs} \times \frac{1 \text{ dozen eggs}}{12 \text{ eggs}} = 1.5 \text{ dozen eggs} \]

B. In chemistry, our counting device is the mole:

\[ 1 \text{ mole of } \text{anything} = 6.023 \times 10^{23} \text{ units of anything} \]

\[ 1 \text{ mole of Gummy bears} = 6.023 \times 10^{23} \text{ Gummy bears} \]

\[ 1 \text{ mole of atoms} = 6.023 \times 10^{23} \text{ atoms} \]

\[ 1 \text{ mole of pennies} = 6.023 \times 10^{23} \text{ pennies} \]

Called Avogadro’s Number
Section 7.1 - Atomic Mass and Formula Mass   Supplemental packet page 120

C. In chemistry, we can't physically count atoms by inspection, so we like to count the total number of atoms, and molecules by weighing.

• Counting atoms - The mole has been given a precise definition as the number of atoms contained in 12.000000 grams "exactly" of pure carbon-12.

\[
1 \text{ mole of } ^{12}_6 \text{C} = 12.000000 \text{ grams of } ^{12}_6 \text{C} = 6.023 \times 10^{23} \text{ atoms of } ^{12}_6 \text{C}
\]

• **Molar Mass** is the mass in grams of one mole of any substance, and is also numerically equal to atomic mass unit [amu] expressed in grams.

\[
\begin{align*}
1 \text{ mole of carbon} & = 12.01 \text{ grams of carbon} \\
3 \text{ moles of iron} & = 55.85 \text{ grams of iron} \\
9 \text{ moles of } \text{Fe(NO}_3\text{)}_3 & = 241.85 \text{ grams of iron}
\end{align*}
\]

D. Molar Mass is numerically equal to atomic mass unit (amu) expressed in grams.
Section 7.4 - Calculations Using Molar Mass

I. How can Avogadro's number be used in calculating the:

A. number of atoms for a given number of moles of an pure element?

How many grams is the same as 18 moles of lead, Pb?

\[
18 \text{ moles} \times \frac{6.023 \times 10^{23} \text{ atoms}}{1 \text{ mole atoms}} = 1.0841 \times 10^{25} \text{ atoms Pb}
\]

\[
1.1 \times 10^{25} \text{ (correct Sig Figs)}
\]

B. number of molecules for a given number of moles of a pure substance?

How many molecules is the same as 18 moles of moles?

\[
18 \text{ moles} \times \frac{6.023 \times 10^{23} \text{ molecules}}{1 \text{ mole molecules}} = 1.1 \times 10^{25} \text{ molecules}
\]

C. number of atoms for a given number of grams of a pure substance?

How many atoms is the same as 18 grams of water?

\[
18 \text{ grams} \times \frac{1 \text{ mole H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{6.023 \times 10^{23} \text{ molecules}}{1 \text{ mole of H}_2\text{O}} \times \frac{3 \text{ atoms}}{1 \text{ molecule of H}_2\text{O}} = 4.8 \times 10^{24} \text{ atoms}
\]

D. number of molecules for a given number of grams of a pure substance?

\[
18.0 \text{ grams} \times \frac{1 \text{ mole H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{6.023 \times 10^{23} \text{ molecules}}{1 \text{ mole H}_2\text{O}} = 6.02 \times 10^{23} \text{ molecules H}_2\text{O}
\]
### Counting Atoms

**How many iron atoms are present in 3.00 moles of iron metal?**

1 mol Fe = 55.85 g Fe = 6.02 x 10³ atoms Fe

\[
x \text{ atoms Fe} = 3.00 \text{ mol Fe} \times \frac{6.02 \times 10^{23} \text{ atoms Fe}}{1 \text{ mol Fe}} = 1.81 \times 10^{24} \text{ atoms Fe}
\]

Work out the following problems (show math set-ups)

**How many sulfur atoms are present in 0.174 moles of S nonmetal?**

\[
x \text{ atoms S} = 0.174 \text{ mol S} \times \frac{6.02 \times 10^{23} \text{ atoms S}}{1 \text{ mol S}} = 1.05 \times 10^{23} \text{ mol S}
\]

**How many moles of K are present in 5.92 x10²⁴ atoms of K metal?**

\[
x \text{ mol K} = 5.92 \times 10^{24} \text{ atoms K} \times \frac{1 \text{ mol K}}{6.02 \times 10^{23} \text{ atoms K}} = 9.83 \text{ atoms K}
\]
Let's go over this example together

\[ x \text{ atoms C} = 27.4 \text{ g C} \]

**How many atoms of C are present in 27.4 grams of carbon nonmetal?**

\[ 12.01 \text{ g C} = 6.02 \times 10^{23} \text{ atoms C} \]

Combined we have a grams to particles equivalent statement

<table>
<thead>
<tr>
<th>27.4 g C</th>
<th>x</th>
<th>1 mol C</th>
<th>x</th>
<th>6.02 x 10^{23} \text{ atoms C}</th>
<th>=</th>
<th>1.37 x 10^{24} \text{ atoms C}</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td>12.01 g C</td>
<td></td>
<td>1 mol C</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

(grams) \times (\text{mol per grams}) \times \text{Avogadro’s number}

\[ \text{moles} \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mole}} \]
How many atoms are present in a formula unit of sodium sulfate Na$_2$SO$_4$?

Just as a mole of atoms is based on the atomic mass or atomic weight, a mole of a compound is based upon the formula mass or formula weight.

<table>
<thead>
<tr>
<th>sodium sulfate, Na$_2$SO$_4$</th>
</tr>
</thead>
<tbody>
<tr>
<td>First:                      How many atoms are there present per formula unit of Na$_2$SO$_4$? 7 atoms</td>
</tr>
<tr>
<td>Second:                     What is the mass in amu of one molecule of sodium sulfate? 142.06 amu</td>
</tr>
<tr>
<td>Third:                      What is the mass—in grams—of one mole of sodium sulfate? 142.06 g</td>
</tr>
<tr>
<td>Fourth:                     How many moles of Na$_2$SO$_4$ are in 16.0 g Na$_2$SO$_4$? 1.13 x 10$^{-1}$ mol</td>
</tr>
</tbody>
</table>

\[
\text{Na}_2\text{SO}_4 \quad 2 \text{ Na} \quad 2 \text{ Na} \times 22.99 = 45.98 \\
\quad 1 \text{ S} \quad 1 \text{ S} \times 32.07 = 32.07 \\
\quad 4 \text{ O} \quad 4 \text{ O} \times 16.00 = 64.00 \\
\quad \frac{142.06 \text{ g Na}_2\text{SO}_4}{7 \text{ atoms}}
\]

\[
x \text{ mol Na}_2\text{SO}_4 = 16.0 \text{ g Na}_2\text{SO}_4 \times \frac{1 \text{ mol Na}_2\text{SO}_4}{142.06 \text{ g Na}_2\text{SO}_4} = 1.13 \times 10^{-1} \text{ mol Na}_2\text{SO}_4
\]
**Molar Mass Calculations; one mole amount of a substance in grams**

<table>
<thead>
<tr>
<th>CH₄</th>
<th>CuSO₄ • 5H₂O</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 C x 12.0 = 12.0</td>
<td>1 Cu x 63.6 = 63.6</td>
</tr>
<tr>
<td>4 H x 1.0 = 4.0</td>
<td>1 S x 32.0 = 32.0</td>
</tr>
<tr>
<td></td>
<td>4 O x 16.0 = 64.0</td>
</tr>
<tr>
<td></td>
<td>5 H₂O x 18.0 = 90.0</td>
</tr>
<tr>
<td>ANS: 16.0</td>
<td>ANS: 249.6</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>C₃H₅Br₂</th>
<th>aluminum nitrate Al(NO₃)₃</th>
</tr>
</thead>
<tbody>
<tr>
<td>3 C x 12.0 = 36.0</td>
<td>1 Al x 27.0 = 27.0</td>
</tr>
<tr>
<td>5 H x 1.0 = 5.0</td>
<td>3 N x 14.0 = 42.0</td>
</tr>
<tr>
<td>2 Br x 78.9 = 157.8</td>
<td>9 O x 16.0 = 144.0</td>
</tr>
<tr>
<td></td>
<td>Note you must be able to derive correct formulas from names ANS: 213.0</td>
</tr>
<tr>
<td>ANS: 200.9</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>C₃H₇OH</th>
<th>calcium dihydrogen phosphate Ca(H₂PO₄)₂</th>
</tr>
</thead>
<tbody>
<tr>
<td>3 C x 12.0 = 36.0</td>
<td>1 Ca x 40.1 = 40.1</td>
</tr>
<tr>
<td>8 H x 1.0 = 8.0</td>
<td>4 H x 1.0 = 4.0</td>
</tr>
<tr>
<td>1 O x 16.0 = 16.0</td>
<td>2 P x 31.0 = 62.0</td>
</tr>
<tr>
<td></td>
<td>8 O x 16.0 = 128.0</td>
</tr>
<tr>
<td>ANS: 60.0</td>
<td>ANS: 234.1</td>
</tr>
</tbody>
</table>
How many moles are there in 41.7 g of NaNO₃?

1. In every calculation problem **ALWAYS** calculate molar mass; **MAKE a Table and Do it**.

<table>
<thead>
<tr>
<th>1 mol NaNO₃ = 85.0 g NaNO₃</th>
<th>1 mol NaNO₃ = 85.0 g NaNO₃</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 Na x 23.0 = 23.0</td>
<td>1 N x 14.0 = 14.0</td>
</tr>
<tr>
<td>1 O x 16.0 = 16.0</td>
<td>3 O x 16.0 = 48.0</td>
</tr>
<tr>
<td>85.0 g NaNO₃ molar mass</td>
<td>85.0 g NaNO₃ molar mass</td>
</tr>
</tbody>
</table>

2. Then, the grams to moles conversion set-up

   \[
   \frac{x \text{ moles}}{85.0 \text{ g NaNO}_3} = \frac{41.7 \text{ g NaNO}_3}{85.0 \text{ g NaNO}_3} = 0.491 \text{ mol NaNO}_3
   \]

Quickly convert to moles by dividing grams by molar mass.

**Memorize this!!!!**
Calculate the number of moles in:

<table>
<thead>
<tr>
<th>Substance</th>
<th>Mass</th>
<th>Molar Mass (g/mol)</th>
<th>Moles Calculation</th>
<th>Moles</th>
</tr>
</thead>
<tbody>
<tr>
<td>Calcium Sulfate (CaSO₄)</td>
<td>12.6 g</td>
<td>136.2</td>
<td>1 Ca x 40.1 = 40.1</td>
<td>9.25 x 10⁻² mol</td>
</tr>
<tr>
<td>Ammonium Carbonate (NH₄₂CO₃)</td>
<td>6.18 x 10³ g</td>
<td>96.01</td>
<td>2 NH₄⁺ x 14.0 = 28.0</td>
<td>64.4 mol</td>
</tr>
</tbody>
</table>

Most students fail in the second half of the semester because they cannot correctly calculate a molar mass 
MAKE at TABLE when calculating molar mass.
How many moles of calcium sulfate atoms are present in 12.6 grams of calcium sulfate ionic salt?

How many moles of ammonium carbonate are present in 6.18 x 10^3 grams of ammonium carbonate ionic salt?

**grams to moles!!!!!!!**
Converting Mole Amounts to Grams

Calculate the number of grams in (show math set-ups):

<table>
<thead>
<tr>
<th>4.22 moles of KCl</th>
<th>0.0196 moles barium nitrate</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>KCl</strong></td>
<td><strong>Ba(NO₃)₂</strong></td>
</tr>
<tr>
<td>1 K x 39.1 = 39.1</td>
<td>1 Ba x 137.3 = 137.3</td>
</tr>
<tr>
<td>1 Cl x 35.5 = 35.5</td>
<td>2 N x 14.0 = 28.0</td>
</tr>
<tr>
<td><strong>74.6 g/mol</strong></td>
<td>6 O x 16.0 = 96.0</td>
</tr>
<tr>
<td>4.22 mol KCl x 74.6 g KCl =</td>
<td>0.0196 mol Ba(NO₃)₂ x 261.3 g Ba(NO₃)₂</td>
</tr>
<tr>
<td>1 mol KCl</td>
<td>1 mol Ba(NO₃)₂</td>
</tr>
</tbody>
</table>

ANS: 3.15 x 10² grams KCl

ANS: 5.12 grams Ba(NO₃)₂

Converting Mole Amounts to Grams, just take moles multiplied by molar mass

Memorize this!!!!
Section 7.5 - Percent Composition and Empirical Formulas

\[
\text{Part} \div \text{Whole} \times 100\% = \text{percentage composition}
\]

What is the percentage of potassium in potassium chloride?
1. Write the correct formula for the substance
2. Calculate the molar mass of the substance (the whole)
3. Divide the individual parts by the whole times 100%

\[
\text{KCl} \quad \begin{align*}
1 \text{ K} & \times 39.1 = 39.1 \\
1 \text{ Cl} & \times 35.5 = 35.5
\end{align*}
\]
\[
\frac{39.1 + 74.6}{74.6} = 52.4\%
\]

What is the percentage of barium in barium nitrate?

\[
\text{Ba(NO}_3\text{)}_2 \quad \begin{align*}
1 \text{ Ba} & \times 137.3 = 137.3 \\
2 \text{ N} & \times 14.0 = 28.0 \\
6 \text{ O} & \times 16.0 = 96.0
\end{align*}
\]
\[
\frac{137.3 + 28.0 + 96.0}{261.3} = 52.5\% \quad 10.7\% \quad 36.7\%
\]
1. Explain the difference between the empirical formula and the molecular formula of a compound.

   An empirical formula is the smallest whole number ratio for a molecular formula. For example, $C_6H_{12}O_6$ would have an empirical formula $C_1H_2O_1$.

2. The molecular formula of the gas acetylene is $C_2H_2$. What is the empirical formula?

   $C_2H_2$ is divisible by “n ratio factor” of two; thus $C_1H_1$ is the empirical formula.

\[
\text{“n ratio factor”} = \frac{\text{molar mass}}{\text{empirical mass}}
\]

\[
2 = \frac{C_2H_2}{C_1H_1}
\]
3. The empirical formula for a compound used as a green paint pigment is $\text{C}_2\text{H}_3\text{As}_3\text{Cu}_2\text{O}_8$.

The molar mass is 1013.71 grams. What is the molecular formula?

Solve for the “n ratio factor”

"n ratio factor" = \( \frac{\text{molar mass}}{\text{empirical mass}} \)

by first calculating the empirical mass

\[
\begin{align*}
2\text{C} \times 12.0 &= 24.0 \\
3\text{H} \times 1.0 &= 3.0 \\
3\text{As} \times 74.9 &= 224.7 \\
2\text{Cu} \times 62.9 &= 125.6 \\
8\text{O} \times 16.0 &= 128.0 \\
\hline
\text{empirical mass} &= 505.3 \text{ g/mol}
\end{align*}
\]

then, by dividing molar mass by empirical mass

"n ratio factor" = \( \frac{1013.71 \text{ g/mol}}{505.3 \text{ g/mol}} \)

Continued on the next slide
Section 7.5 - Percent Composition and Empirical Formulas

Continued from the previous slide

Finally, multiply empirical by $n$

$[\text{C}_2\text{H}_3\text{As}_3\text{Cu}_2\text{O}_8] \times 2$

empirical formula

$\text{C}_4\text{H}_6\text{As}_6\text{Cu}_4\text{O}_{16}$

molecular formula exists as two empirical formulas

molecular formula 1013.71 g/mol
Section 7.5 - Percent Composition and Empirical Formulas

4. A sugar which is broken down by the body to produce energy has the following percent composition (C = 39.99 %, H = 6.713 % and O = 53.29 %) and a molar mass of 210.18 g. What is the empirical formula? What is the molecular formula?

Consider the following, if you had a 100 gram sample of this substance, how many grams of it would be carbon, hydrogen and oxygen?

39.99% C x 100 gram sample = 39.99 grams carbon in sample
6.713% H x 100 gram sample = 6.713 grams hydrogen in sample
53.29% O x 100 gram sample = 53.29 grams oxygen in sample

On the next slide, we’ll convert grams of sample to moles.

From knowing the number of moles, a mole ratio of atoms in the chemical formula will be calculated.
Section 7.5 - Percent Composition and Empirical Formulas

MAKE A TABLE!!!!!! And convert grams to moles which will be the mole ratio

4. A sugar which is broken down by the body to produce energy has the following percent composition (C = 39.99 %, H = 6.713 % and O = 53.29 %) and a molar mass of 210.18 g. What is the empirical formula? What is the molecular formula?

<table>
<thead>
<tr>
<th>analysis</th>
<th>C = 39.99 %</th>
<th>H = 6.713 %</th>
<th>O = 53.29 %</th>
</tr>
</thead>
<tbody>
<tr>
<td>grams</td>
<td>39.99 grams</td>
<td>6.713 grams</td>
<td>53.29 grams</td>
</tr>
<tr>
<td>assume 100 g sample</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>MM (molar mass)</td>
<td>12.0 g/mol</td>
<td>1.0 g/mol</td>
<td>16.0 g/mol</td>
</tr>
<tr>
<td>mole</td>
<td>3.33 mol</td>
<td>6.71 mol</td>
<td>3.33 mol</td>
</tr>
<tr>
<td>ratio</td>
<td>3.33 mol</td>
<td>6.71 mol</td>
<td>3.33 mol</td>
</tr>
<tr>
<td>divide by the smallest</td>
<td>3.33 mol</td>
<td>3.33 mol</td>
<td>3.33 mol</td>
</tr>
<tr>
<td>whole number ratio</td>
<td>C₁</td>
<td>H₂</td>
<td>O₁</td>
</tr>
</tbody>
</table>

1C to 2H to 1O is a 1 : 2 : 1 mole ratio of carbon to hydrogen to oxygen in the empirical formula for a substance of 39.99g C, 6.713gH, 53.29 C by mass
Section 7.5 - Percent Composition and Empirical Formulas

To check your work, consider calculating a percentage composition.

The empirical formula for the sugar used in the analysis is $C_1H_2O_1$. Calculate a percentage composition of each element in the formula.

1. Calculate the empirical mass
   
   $1C \times 12.0 = 12.0 \div 30.0 = 0.400 \times 100\% = 40.0\%
   
   $2H \times 1.0 = 2.0 \div 30.0 = 0.67 \times 100\% = 6.7\%
   
   $1O \times 16.0 = 16.0 \div 30.0 = 0.400 \times 100\% = 53.3\%

   \text{empirical mass} \quad \text{30.0 g/mol}

2. Divide the parts by the whole and times by 100 %

   \[ C_1 \quad H_2 \quad O_1 \]

   This type of analysis always produces the lowest whole number ratio, thus we have now just calculated the empirical formula.
5. Using the empirical formula and molar mass that was given determine the molecular formula for the substance given the molar mass is 210.18 g/mol.

The empirical formula for a compound used in the past as C$_1$H$_2$O$_1$.

The molar mass is 210.18 g/mol. What is the molecular formula?

First, calculate the “n ratio factor”

\[
\text{“n ratio factor”} = \frac{\text{molar mass}}{\text{empirical mass}}
\]

Second, calculate the empirical mass

\[
\begin{align*}
1\text{C} \times 12.0 &= 12.0 \\
2\text{H} \times 1.0 &= 2.0 \\
1\text{O} \times 16.0 &= 16.0 \\
\hline
\text{empirical mass} &= 30.0 \text{ g/mol}
\end{align*}
\]

Third, divide molar mass by empirical mass

\[
\text{“n ratio factor”} = \frac{210.18 \text{ g/mol}}{30.0 \text{ g/mol}} = 6
\]

Continued on the next slide
The molecular formula is $C_6H_{12}O_6$.

Continued from the previous slide:

fourth, multiply empirical formula by the “n ratio factor”

\[
\text{“n ratio factor”} = \frac{210.18 \text{ g/mol}}{30.0 \text{ g/mol}}
\]

\[
[C_1H_2O_1] \times 6
\]

empirical formula

\[
\text{“n ratio factor”} = 6
\]

$C_6H_{12}O_6$

molecular formula 210.18 g/mol

Section 7.6 - Molecular Formulas