6.1 How Much Sodium?

6.2 Counting Nails by the Pound
A hardware store customer buys 2.60 pounds of nails. How many nails did the customer buy?

A dozen of the nails has a mass of 0.150 pounds.

\[
\text{nails} = \frac{2.60 \text{ lbs} \times 1 \text{ doz. nails}}{0.150 \text{ lbs}} = 208 \text{ nails}
\]

- The customer bought 2.60 lbs of nails and received 208 nails. He counted the nails by weighing them!

6.3 & 6.4 Counting Atoms and Molecules by the Gram
By analogy we can calculate how many atoms or molecules there are in a given mass of an element or compound.

Atoms or Molecules and Moles
- If we can find the mass of a particular number of atoms or molecules, we can use this information to convert the mass of a element or compound sample to the number of atoms or molecules in the sample.
Counting Atoms or Molecules by Moles

The number of atoms or molecules we will use is $6.022 \times 10^{23}$ and we call this a **mole**

- 1 mole = $6.022 \times 10^{23}$ particles
- Like 1 dozen = 12 particles

• The number of particles in 1 mole is called Avogadro's Number = $6.0221421 \times 10^{23}$

We can make two conversion factors:

A) $\frac{1 \text{ mole}}{6.022 \times 10^{23} \text{ atoms}}$
B) $\frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mole}}$

A) For converting atoms $\rightarrow$ mole
B) For converting mole $\rightarrow$ atoms

Practice 1

Conversion sequence: **moles $\rightarrow$ atoms, molecules**

1. How many atoms are in 6.28 moles of aluminum?
2. How many atoms are in 90.43 moles of copper?
3. How many atoms in 7.64 moles of barium?
4. How many molecules in 3.72 moles of sulfur dioxide?
5. 76.4 moles of oxygen difluoride contain how many molecules?

Practice 2

Conversion sequence: **atoms, molecules $\rightarrow$ moles**

1. How many moles of water are represented by $8.33 \times 10^{18}$ molecules of water?
2. How many moles of magnesium is $3.01 \times 10^{22}$ atoms of magnesium?
3. How many moles are $1.20 \times 10^{25}$ atoms of phosphorous?

Moles and Mass

The mass of one mole of atoms or molecules is called the **molar mass**
Moles and Mass (cont.)

The molar mass of an element, in grams, is numerically equal to the element’s atomic mass.

Mole and Mass Relationships

<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>

1 mole Sulfur 32.06 g

1 mole Carbon 12.01 g

Moles and Mass (cont.)

The molar mass of a compound, in grams, is numerically equal to the sum of the atomic masses of the elements in the compounds formula.

The molar mass of water is calculated from the atomic weights of hydrogen and oxygen.

Formula = $H_2O$

Formula Mass = $2(1.01 \text{ amu H}) + 16.00 \text{ amu O} = 18.02 \text{ amu}$

Molar Mass = 18.02 g

Practice 3

Calculate formula mass and Molar Mass

<table>
<thead>
<tr>
<th>FORMULA</th>
<th>FORMULA MASS (amu)</th>
<th>MOLAR MASS (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Br$_2$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>sodium sulfide</td>
<td></td>
<td></td>
</tr>
<tr>
<td>potassium hydroxide</td>
<td></td>
<td></td>
</tr>
<tr>
<td>fluorine</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ni</td>
<td></td>
<td></td>
</tr>
<tr>
<td>BaCl$_2$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Fe(SO$_4$)$_2$</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
**Practice 4**

**Converting Between Grams and Moles**

How many grams for each of the following:

1. 7.24 moles of silver phosphate
2. 2.88 moles of diphosphorous pentoxide
3. 0.0009273 moles of zinc bicarbonate
4. 154.8 moles of silicon tetraiodide
5. 88.624 moles of silver

**Practice 5**

**Converting Between Grams and Number of Atoms or Molecules**

How many moles for each of the following?

1. 28 grams of CO
2. 452 g of argon
3. 9.273 kg of zinc bicarbonate
4. 25.0 g of iron
5. 88.624 mg of silver

**Practice 6**

**Conversion sequence: grams → moles → atoms**

How many atoms or molecules for each of the following?

1. 28 grams of CO
2. 452 g of argon
3. 9.273 kg of zinc bicarbonate
4. 25.0 g of iron
5. 88.624 mg of silver

**Practice 7**

**Conversion sequence: atoms → moles → grams**

How many grams in each of the following?

1. $3.01 \times 10^{23}$ atoms of sodium (Na)
2. $4.5 \times 10^{25}$ atoms of argon
3. $9.27 \times 10^{30}$ molecules of zinc bicarbonate
4. $2.50 \times 10^{19}$ atoms of iron
5. $8.86 \times 10^{15}$ molecules of dinitrogen tetroxide
6.5 Chemical Formulas as Conversion Factors

- 1 spider ≡ 8 legs
- 1 chair ≡ 4 legs
- 1 H₂O molecule ≡ 2 H atoms ≡ 1 O atom

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 mole Na, 1 mole Cl</td>
</tr>
<tr>
<td>1 mol H₂O</td>
<td>2 mol H, 1 mole 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>

Aka… Mole Ratios… always whole number ratios

Practice 8

1. How many moles Cl in 4.7 mol CaCl₂?
2. How many mol of H in 54.1 mol C₁₀H₂₂?
3. How many oxygen atoms in 2.00 mol O₂?
4. How many grams of Cl in 55 g of CF₃Cl?
5. How many grams of Fe in 1.0 x 10³ kg of Fe₂O₃?

Writing Mole Ratios

<table>
<thead>
<tr>
<th>Moles of Compound</th>
<th>Moles of Constituents</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 mol NaCl</td>
<td>1 mole Na, 1 mole Cl</td>
</tr>
<tr>
<td>1 mol H₂O</td>
<td>2 mol H, 1 mole 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>

6.6 Percent Composition

- Percentage of each element in a compound by mass

Determined from
1. The formula of the compound
2. The experimental mass analysis of the compound

\[
\text{Percentage} = \frac{\text{part}}{\text{whole}} \times 100\% 
\]
2. Percent Composition from experiment
A 30.0 g sample of carvone contains 24.0 g of C, 3.2 g O and the rest H?

What is it’s percent composition

Mass Percent as a Conversion Factor

6.8 & 6.9 Empirical and Molecular Formulas

• The simplest, whole-number ratio of atoms in a molecule is called the Empirical Formula
• The Molecular Formula is a multiple of the Empirical Formula

Hydrogen Peroxide
Molecular Formula = H₂O₂
Empirical Formula = HO

Benzone
Molecular Formula = C₆H₆
Empirical Formula = CH

Glucose
Molecular Formula = C₆H₁₂O₆
Empirical Formula = CH₂O

Finding an Empirical Formula

1) convert the percentages to grams (skip if already grams)
2) convert grams to moles (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 whole number (2, 3, 4, 5, etc.) to make all mole ratios whole numbers. (skip if all mole ratios already whole numbers)

Finding an Empirical Formula from Experimental Data
Example:

- A laboratory analysis of aspirin determined the following mass percent composition. Find the empirical formula.

  \[ C = 60.00\% \]
  \[ H = 4.48\% \]
  \[ O = 35.53\% \]

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>
</tr>
</thead>
<tbody>
<tr>
<td>glyceraldehyde</td>
<td>( C_3H_4O_3 )</td>
<td>CH(_2)O</td>
</tr>
<tr>
<td>erythrose</td>
<td>( C_4H_6O_3 )</td>
<td>CH(_2)O</td>
</tr>
<tr>
<td>arabinose</td>
<td>( C_5H_8O_5 )</td>
<td>CH(_2)O</td>
</tr>
<tr>
<td>glucose</td>
<td>( C_6H_12O_6 )</td>
<td>CH(_2)O</td>
</tr>
</tbody>
</table>

Molecular Formulas

- 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

\[
\text{Molar Mass}_{\text{real formula}} = \text{factor used to multiply subscripts} \times \text{Molar Mass}_{\text{empirical formula}}
\]

All these molecules have the same empirical formula. How are the molecules different?

<table>
<thead>
<tr>
<th>Name</th>
<th>Molecular Formula</th>
<th>Molar Mass, g</th>
<th>Empirical Formula</th>
<th>EF Molar Mass, g</th>
<th>FACTOR</th>
</tr>
</thead>
<tbody>
<tr>
<td>glyceraldehyde</td>
<td>( C_3H_4O_3 )</td>
<td>90</td>
<td>CH(_2)O</td>
<td>30</td>
<td>3</td>
</tr>
<tr>
<td>erythrose</td>
<td>( C_4H_6O_3 )</td>
<td>120</td>
<td>CH(_2)O</td>
<td>30</td>
<td>4</td>
</tr>
<tr>
<td>arabinose</td>
<td>( C_5H_8O_5 )</td>
<td>150</td>
<td>CH(_2)O</td>
<td>30</td>
<td>5</td>
</tr>
<tr>
<td>glucose</td>
<td>( C_6H_12O_6 )</td>
<td>180</td>
<td>CH(_2)O</td>
<td>30</td>
<td>6</td>
</tr>
</tbody>
</table>

Determine the Molecular Formula of Cadinene if it has a molar mass of 204 g and an empirical formula of \( C_5H_8 \)

1. Determine the empirical formula
   - May need to calculate it as previous

2. Determine the molar mass of the empirical formula

\[ 5 \times 60.05 \text{ g} + 8 \times 8.064 \text{ g} = 68.11 \text{ g} \]
3. Divide the given molar mass of the compound by the molar mass of the empirical formula
   ✓ Round to the nearest whole number

\[
\frac{204 \text{ g}}{68.11 \text{ g}} = 3
\]

4. Multiply the empirical formula by the factor above to give the molecular formula

\[\text{(C}_5\text{H}_8)_3 = \text{C}_{15}\text{H}_{24}\]