I. Counting Atoms
   A. Atomic Masses
   B. Moles
   C. Molar Mass

II. Composition and Formulas for Compounds
   A. Percent Composition
   B. Calculating Empirical Formulas
   C. Calculating Molecular/Ionic Formulas

A day without chemistry is like a day without the Sun!!
I. Counting Atoms
Candy and Other Small Things

Small Items
- rather than count, we get a rough idea about quantity from masses

Chemistry (e.g. 6 Criteria Air Pollutants (EPA))

| CO  | Pb  | NO₂ | O₃  | P.M. | SO₂ |

How do we estimate the amount of SO₂ released?
1. Find **mass** of sulfur that is combusted
2. Relate **mass** of sulfur combusted to a # of atoms combusted
3. Relate # of S atoms combusted to the # of SO₂ molecules produced

Two Types of Mass Units for Atoms
A. Atomic Masses

Need – mass of sulfur atom

Isotopes and Atomic Mass Units

Isotopes – atoms with same atomic # but different mass #'s

Sulfur – three naturally occurring isotopes $^{32}S, ^{33}S, ^{34}S$

Need – average mass of sulfur atom

<table>
<thead>
<tr>
<th>Sulfur isotope</th>
<th>$^{32}S$</th>
<th>$^{33}S$</th>
<th>$^{34}S$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass (g)</td>
<td>5.3152×10^{-23} g</td>
<td>5.4700×10^{-23} g</td>
<td>5.6400×10^{-23} g</td>
</tr>
<tr>
<td>Natural Abundance</td>
<td>95%</td>
<td>4%</td>
<td>1%</td>
</tr>
<tr>
<td>100 atom sample</td>
<td>95</td>
<td>4</td>
<td>1</td>
</tr>
</tbody>
</table>

$$\text{avg mass} = \frac{5.0494\times10^{-21} g + 2.1880\times10^{-22} g + 5.6400\times10^{-23} g}{100\ \text{atoms}}$$

$$= 5.3246\times10^{-23} \text{ g / atom}$$
Example
If I obtain a sample of 89 argon atoms, what is the mass in amu?

B. Moles
Need – a # of atoms that provides an easily measurable mass
Mole – # of atoms that provides an easily measurable mass
The Mole (mol)
Atomic Mass – mass of 1 atom (amu)

Molar Mass – mass of 1 mole of atoms (g)

1 mol of pennies -
### Mole Cont’d

<table>
<thead>
<tr>
<th>Element</th>
<th>Mass of 1 atom (amu)</th>
<th>Mass of $6.022 \times 10^{23}$ atoms (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mg</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Rn</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Conversion factors**
Calculations with the Mole

1. If I have 0.256 g of Cu how many moles of Cu do I have?

How many Cu atoms do I have?

2. If I have $6.25 \times 10^{24}$ atoms of Ar, what is the mass of my sample of Ar?
C. **Molar Mass**
What is the mass of 1 molecule of Methane? \( \text{CH}_4 \)

What is the mass of 1 mole of methane molecules?

What is the molar mass of \( \text{Fe(ClO}_2\text{)}_2 \)?
<table>
<thead>
<tr>
<th>Compound</th>
<th>1 Molecule (# Atoms)</th>
<th>1 Mole (# Moles of atoms)</th>
<th>Mass (Amu) Molar Mass(g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C$_2$H$_6$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ca$_3$(PO$_4$)$_2$</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
4. How many moles of C₂H₆ are in a 4.367 g sample of C₂H₆?

How many molecules of C₂H₆ are in the sample?
5. What is the mass (g) of butanoic acid (C₄H₈O₂) contained in a $2.35 \times 10^{22}$ molecule sample of butanoic acid?
6. How many Cl atoms are in a 2.54 g sample of CCl₄?

7. A sample of H₂S contains 6.65x10²³ H atoms. What is the mass of the sample of H₂S?
II. Experimentally Determining the Composition and Formulas for Compounds

A. Calculating Percent Composition

\[ \text{Percent} = \frac{\text{Part}}{\text{Whole}} \times 100\% \]

Example → Score 94 out of 106 is that an A?

\[ \text{Percent} = \frac{94 \text{ points}}{106 \text{ points}} \times 100\% = 88.67\% \]

Mass Percent in Chemical Compounds
8. The chemical formula for Kryptonite has been proposed to be $Na_3Li_3B_2(Si_2O_7)_2(OH)F$. What is the percent by mass of oxygen in Kryptonite?
B. Calculating Empirical Formulas

Goal → Experimentally Determine Formula for Chemical Compounds

Types of Chemical Formulas
1. Molecular Formula (M.F.)
   - formula of an actual molecule

2. Empirical Formula (E.F.)
   - provides relative numbers of atoms in a molecule
   - subscripts are reduced to the smallest whole number ratios
How do we Calculate Empirical Formulas?

1. Determine the mass of each element present in the compound.

2. Convert mass to number of moles present for each atom.

3. Find the smallest ratio of each atom to all of the other atoms.
### How do we Calculate Empirical Formulas? Cont’d

<table>
<thead>
<tr>
<th>Calculated E.F.</th>
<th>Multip. Factor</th>
<th>New E.F.</th>
<th>Other Sub. Use</th>
</tr>
</thead>
<tbody>
<tr>
<td>C₁H₁.₅</td>
<td>2</td>
<td>C₂H₃</td>
<td>(0.5, 1.5, 2.5, etc)</td>
</tr>
<tr>
<td>C₁H₁.₃₃</td>
<td>3</td>
<td>C₃H₄</td>
<td>(0.33, 1.33, etc)</td>
</tr>
<tr>
<td>C₁H₁.₂₅</td>
<td>4</td>
<td>C₄H₅</td>
<td>(0.25, 1.25, etc)</td>
</tr>
<tr>
<td>C₁H₁.₆₆</td>
<td>3</td>
<td>C₃H₅</td>
<td>(0.66, 1.66, etc)</td>
</tr>
<tr>
<td>C₁H₁.₇₅</td>
<td>4</td>
<td>C₄H₇</td>
<td>(0.75, 1.75, etc)</td>
</tr>
</tbody>
</table>
Example
9. If a 3.000|5 g sample of copper metal is heated in a chlorine gas atmosphere, the mass of copper (I or II?) chloride produced is 6.349 g. Calculate the empirical formula of the copper chloride.
10. Putrescine, an incredibly foul smelling compound that is generated during the decomposition of flesh, contains nitrogen, hydrogen, and carbon. During his quest to obtain the “Ring” and help Sauron conquer Middle Earth, Saruman the White, a magician and master chemist, tried to synthesize putrescine and use it as a weapon against the inhabitants of Middle Earth. In order to obtain the chemical formula of putrescine, Saruman analyzed a sample of putrescine and found it contained 34.13\% N and 58.50\% C. What is the empirical formula for putrescine?

Saruman, you stink!
Obtaining Empirical Formulas from Mass %

1. Determine the mass\% and mass of each element present in the compound.

2. Convert mass to number of moles present for each atom.

3. Find the smallest ratio of each atom to all of the other atoms.
C. Calculating Molecular/Ionic Formulas from Empirical Formulas
How many sticks of gum in a pack of gum?