Chapter 1
The Atomic Nature of Matter

• 1-1 Chemistry: Science of Change
• 1-2 The Composition of Matter
• 1-3 The Atomic Theory of Matter
• 1-4 Chemical Formulas and Relative Atomic Masses
• 1-5 The Building Blocks of the Atom
• 1-6 Finding Atomic Masses the Modern Way
• 1-7 The Mole Concept: Counting and Weighing Atoms and Molecules
• 1-8 Finding Empirical and Molecular Formulas the Modern Way
• 1-9 Volume and Density
Definitions

- **Analysis** (Take things apart)
- **Synthesis** (Put things together)
- **Physical Properties** (Color, odor, taste, boiling point, etc.)
- **Chemical properties** (with respect to other materials, e.g., uniformity)
- **Substance** (can not be separated by physical means, refers to elements and compounds, never mixtures)
- **Elements** (cannot be decomposed into simpler substances)
- **Compounds** (contain two or more elements bonded together, e.g., NaCl)
- **Molecule** (a few atoms connected together, e.g., CO₂)
- **Mixtures** (can be separated into two or more substances)
- **Homogenous** (uniform throughout, solutions)
- **Heterogeneous** (properties vary from region to region)
- **Phase** (liquid, gas, solid)
Examples

• Table Salt

• Wood

• Mercury

• Air

• Water

• What is an example of a **homogenous** sample that would gradually become **heterogeneous** if left to itself?
Atomic Theory of Matter

- Law of conservation of mass: Mass is neither created nor destroyed in a chemical reaction
- Dalton’s Atomic Theory of Matter (1808):
  1. All matter consists of solid and indivisible atoms
  2. All atoms of a given chemical element are identical in mass and in all other properties
  3. Different elements have different kinds of atoms; these atoms differ in mass from element to element
  4. Atoms are indestructible and retain their identity in all chemical reactions
  5. The formation of a compound from its elements occurs through the combination of atoms of unlike elements in small whole-number ratio.
Chemical Formulas and Relative Atomic Masses

• **Chemical Formulas** display symbols for the elements and the relative number of atoms

• **Molecules** are groupings of two or more atoms bound closely together by strong forces that maintain them in a persistent combination
Building Blocks of the Atom

- Electrons, Protons and Neutrons
  - **Electrons** discovered in 1897 by Thomson
  - Rutherford proposed that the atomic nucleus was composed of neutral particles called **Neutrons** and positively charged particles called **protons**
  - Neutron number = N
  - Atomic number = \( Z = \) number of Protons
  - Atomic mass number = \( A \)
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<table>
<thead>
<tr>
<th></th>
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</thead>
<tbody>
<tr>
<td>6</td>
<td>7</td>
<td>14</td>
<td>15</td>
</tr>
<tr>
<td>12.011</td>
<td>14.007</td>
<td>28.086</td>
<td>30.794</td>
</tr>
<tr>
<td>C</td>
<td>N</td>
<td>Si</td>
<td>P</td>
</tr>
<tr>
<td>Carbon</td>
<td>Nitrogen</td>
<td>Silicon</td>
<td>Phosphorus</td>
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Atomic Mass

\[ A = Z + N \]

Atomic Mass = \# Protons + \# Neutrons

For Carbon, \( 12 = 6 + \text{Neutrons} \)

Neutrons = 6

Every Carbon atom has 6 electrons, 6 protons and 6 neutrons
Mass Spectrometry and Isotopes

- **Mass Spectrometer** accelerates ions (or molecular ions) in an electric field and then separates those ions by relative mass in a magnetic field.

![Mass Spectrometer Separation of Chlorine](image)

<table>
<thead>
<tr>
<th>Relative Mass</th>
<th>Relative Amount</th>
</tr>
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<tbody>
<tr>
<td>35</td>
<td>80</td>
</tr>
<tr>
<td>37</td>
<td>20</td>
</tr>
</tbody>
</table>

17  
35.453  
Cl  
Chlorine
Atoms

- **Avogadro’s Number** is the number of $^{12}\text{C}$ atoms in exactly 12 grams of carbon
  \[ N_0 = 6.0221420 \times 10^{23} \]

- The mass, in grams, of Avogadro's number of atoms of an element is numerically equal to the relative atomic mass of that element
Molecules

• **Relative Molecular Mass** of a molecule equals the sum of the relative atomic masses of all of the atoms making up the molecule
Moles

• A mole measures the chemical amount of a substance
• Mole is an abbreviation of gram molecular weight
• One mole of a substance equals the amount that contains Avogadro's number of atoms, molecules.
• One mole = Molar mass (M) of that element or molecule
Exercise 1-6

- Molecules of isoamyl acetate have the formula $C_7H_{14}O_2$. Calculate (a) how many moles and (b) how many molecules are present in 0.250g of isoamyl acetate.

- **Strategy:**
  1. Calculate molar mass of $C_7H_{14}O_2$
  2. Calculate the number of moles in 0.250 grams
  3. Using Avogadro’s number to calculate the number of molecules in $X$ moles of $C_7H_{14}O_2$
Exercise 1-6

- Molecules of isoamyl acetate have the formula $\text{C}_7\text{H}_{14}\text{O}_2$. Calculate (a) how many moles and (b) how many molecules are present in 0.250g of isoamyl acetate.

Solution:

1. Calculate molar mass of $\text{C}_7\text{H}_{14}\text{O}_2$

2. Calculate the number of moles in 0.250 grams

3. Using Avogadro’s number calculate the number of molecules in X moles of $\text{C}_7\text{H}_{14}\text{O}_2$
Percentage Composition from Empirical or Molecular Formula
Exercise 1-8

• Tetrodotoxin, a potent poison found in the ovaries and liver of the globefish, has the empirical formula \( \text{C}_{11}\text{H}_{17}\text{N}_{3}\text{O}_{8} \). Calculate the mass percentages of the four element in this compound.

**Strategy:**

1. Calculate molar mass of \( \text{C}_{11}\text{H}_{17}\text{N}_{3}\text{O}_{8} \), by finding the mass contributed by each element

2. Divide the mass for each element by the total mass of the compound.
Exercise 1-8

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Solution:

1. Calculate molar mass of $C_{11}H_{17}N_3O$, by finding the mass contributed by each element

2. Divide the mass for each element by the total mass of the compound.
Exercise 1-10

- Moderate Heating of 97.44 mg of a compound containing nickel, carbon and oxygen and no other elements drives off all of the carbon and oxygen in the form of carbon monoxide (CO) and leaves 33.50 mg of metallic nickel behind. Determine the empirical formula of the compound.

Strategy:

1. Write the reaction
2. Use the conservation of mass to find the amount of CO
3. Find the number of moles of CO and Nickel
4. Find the ratios of the moles for each substance by dividing each by the smallest one, i.e., normalize to the smallest.
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Volume and Density
Exercise 1-13

• The density of liquid mercury at 20 deg C is 13.594 g cm\(^{-3}\). A chemical reaction requires 0.560 mol of mercury. What volume (in cubic centimeters) of mercury should be measured out at 20°C?

**Strategy:**
1. Use density and mass to find volume. Rearrange

\[ d = \frac{m}{V} \quad V = \frac{m}{d} \]

2. Density is given, can find mass from the number of moles of mercury which is given

Exercise 1-13

- The density of liquid mercury at 20 deg C is 13.594 g cm\(^{-3}\). A chemical reaction requires 0.560 mol of mercury. What volume (in cubic centimeters) of mercury should be measured out at 20°C?

**Solution:**

1. Use density and mass to find volume.

\[ V = \frac{m}{d} \]

2. Density is given, can find mass from the number of moles of mercury which is given

• **WebCT problems?** Contact Dr. George McKelvy in Chem Annex Room 41 or via email george.mckelvy@chemistry.gatech.edu

• **PRS unit registration.** You must register your unit before midnight on **August 20** so that I update the PRS class file with your name and number. If you change your unit at any time send me email

• **Examples and/or Exercises** in the text are for your learning and practices. They are not to be turned in.

• If you did not purchase “Chemistry” combination lock, you can do so at GT Bookstore in Technology Square

• **Wear safety goggles** in the lab at all times! Their your eyes, *protect them*. Don’t rely on someone else to remind you.
Chapter 1
The Atomic Nature of Matter

Examples / Exercises (for your practice only, not to be turned in)
– All (1-1 thru 1-13)

HW Problems (to be performed on WebCT)
– 10, 16, 24, 28, 32, 36, 48, 54, 72