Chapter 2 Atoms, Molecules, and Ions

The Atomic Theory

Conservation of Mass and the Law of Definite Proportions

*The law of conservation of mass:* the total mass of material present after a chemical reaction is the same as the total mass before the reaction.

Dalton’s Atomic Theory and the Law of Multiple Proportions

1. Each element is composed of extremely small particles called atoms.

2. All atoms of a given element are identical; the atoms of different elements are different and have different properties (including masses).

3. Atoms of an element are not changed into different types of atoms by chemical reactions; atoms are neither created nor destroyed in chemical reactions.

4. Compounds are formed when atoms of more than one element combine; a given compound always has the same relative number and kind of atoms.

Atoms are the basic building blocks of matter. They are the smallest particles of an element that retain the chemical identity of the element.

*Law of multiple proportions:* if two elements A and B combine to form more than one compound, the masses of B that can combine with a given mass of A are in a ratio of small whole numbers.
The Structure of Atoms: Electrons

Atoms are composed of subatomic particles.

**Radiation** - the emission and transmission of energy through space in the form of waves.

There are two types of electrical charge, positive (+) (+ve) and negative (-) (-ve).

**Law of Electrostatic Attraction**: like charges repel one another, unlike charges attract.

**Cathode rays and electrons**

Thomson 1.76x10^8 coulombs per gram

Millikan oil-drop experiment: charge of an electron 1.6x10^-19 C

\[
\text{Mass} = \frac{1.60 \times 10^{-19} \text{C}}{1.76 \times 10^8 \text{C/g}} = 9.10 \times 10^{-28} \text{g}
\]

**Radioactivity** - spontaneous emission of particles and/or radiation.

<table>
<thead>
<tr>
<th>Name</th>
<th>Symbols</th>
<th>Charge</th>
<th>mass (g/particle)</th>
</tr>
</thead>
<tbody>
<tr>
<td>alpha particles</td>
<td>$\frac{4}{2}^4\text{He}$, $\frac{4}{2}\alpha$</td>
<td>2+</td>
<td>$6.65 \times 10^{-24}$</td>
</tr>
<tr>
<td>beta particle</td>
<td>$0^-\text{e}$, $0^-\beta$</td>
<td>1-</td>
<td>$9.11 \times 10^{-28}$</td>
</tr>
<tr>
<td>gamma ray</td>
<td>$0^-\gamma$, $\gamma$</td>
<td>0</td>
<td>0</td>
</tr>
</tbody>
</table>
The Structure of Atoms: Protons and Neutrons

When discussing the mass of atoms we will use the atomic mass unit \((\text{amu})\). The amu is \(1.66054 \times 10^{-24}\) g.

<table>
<thead>
<tr>
<th>Particle</th>
<th>Charge</th>
<th>Mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton</td>
<td>Positive (1+)</td>
<td>1.0073</td>
</tr>
<tr>
<td>Neutron</td>
<td>None (neutral)</td>
<td>1.0087</td>
</tr>
<tr>
<td>Electron</td>
<td>Negative (1-)</td>
<td>(5.486 \times 10^{-4})</td>
</tr>
</tbody>
</table>

Most of the mass is in the nucleus (protons and neutrons).

The size of atoms is small! Atomic diameters are on the order of \(1 \times 10^{-10}\) m to \(5 \times 10^{-10}\) m (100 – 500 pm). A convenient unit (not SI) to express atomic diameters is the angstrom (Å).

\(1\ \text{Å} = 10^{-10}\ m\). Atoms are 1 – 5 Å in diameter.

Atomic Number

Isotopes, Atomic Numbers and Mass Numbers

All atoms of an element have the same number of protons in the nucleus.

The number of protons determines the type of atom.

In an atom, the number of electrons equals the number of protons.

Atoms of the same element that differ in the number of neutrons, and mass, are called isotopes.
Atomic number (Z) (subscript): number or protons in nucleus. Often omitted since the atomic symbol indicates this.

Mass number (A) (superscript): total number of protons and neutrons.

\[ ^{12}\text{C} \quad ^{12}\text{C} \quad "carbon\ 12" \]

\[ ^{14}\text{C} \quad "carbon\ 14" \quad 6\ \text{protons\ (carbon)} \quad 8\ \text{neutrons} \]

An atom of a specific isotope is called a nuclide.

Difference in elements is due to the difference in the number of subatomic particles.

**Elements and the Periodic Table**

periods

- groups

- metals

- nonmetals

- metalloids or semimetals
Periodic Table of the Elements
Main groups

Group 1A - Alkali metals

Group 2A - Alkaline earth metals

Group 7A - Halogens

Group 8A - Noble gases

Transition metal groups

Inner transition metal groups

Molecules, Ions and Chemical Bonds

A molecule is an assembly of two or more atoms tightly bound together.

Molecules and Chemical Formulas.

Chemical formulas give the composition substances.

The subscripts in the formula tell us the number of that type of atom present in the molecule.

Example: \( O_2 \)  two oxygen atoms

\( O_3 \)  three oxygen atoms

\( H_2O \)  two hydrogen atoms one oxygen atom
Molecules containing two atoms are called **diatomic**. Ex: CO, HF, NO, HCl

Elements that occur as diatomic molecules are: N\textsubscript{2}, O\textsubscript{2}, H\textsubscript{2}, F\textsubscript{2}, Cl\textsubscript{2}, Br\textsubscript{2}, and I\textsubscript{2}.

When we speak of these elements we are referring to the diatomic form listed above.

**Molecular compounds** are compounds that are composed of molecules.

Composition of compounds is given by their **chemical formulas**.

*Most molecular substances that we will encounter contain only nonmetals.*

**Molecular and Empirical Formulas**

**Molecular formulas** indicate the actual number and types of atoms in a molecule.

**Empirical formulas** give only the relative number of atoms of each type. The subscripts are always the smallest whole number ratio.

Example:

- H\textsubscript{2}O\textsubscript{2} \quad \text{HO}
- C\textsubscript{2}H\textsubscript{4} \quad \text{CH\textsubscript{2}}
- C\textsubscript{6}H\textsubscript{12} \quad \text{CH\textsubscript{2}}
- H\textsubscript{2}O \quad \text{H\textsubscript{2}O}

**Picturing Molecules**

The structural formula of a substance shows which atoms are attached to which within the molecule.

- **Water**
- **Hydrogen Peroxide**
- **Methane**
Addition or removal of electrons from a neutral atom results in the formation of a charged particle called an **ion**.

An ion with positive charge is called a **cation**.

A negatively charged ion is called an **anion**.

The net charge is represented by a superscript.

Superscripts +, 2+, and 3+ mean a net charge resulting from the loss of one, two, or three electrons.

Superscripts -, 2-, and 3- mean a net charge resulting from the gain of one, two, or three electrons.

In general, metal atoms lose electrons; nonmetals tend to gain electrons.
Naming Chemical Compounds

Four types of Inorganic Compounds: ionic, molecular, acid and bases, and hydrates.

Predicting Ionic Charges

Many atoms gain or lose electrons so as to end up with as many electrons as the closest noble gas.

Group 1A atoms form 1+ ions 
Group 2A atoms form 2+ ions 
Group 7A atoms form 1- ions 
Group 6A atoms form 2– ions

Ionic Compounds

Ionic compounds contain positively charged ions and negatively charged ions. (salts)

Generally, cations are metal ions and anions are nonmetal ions.

Ionic compounds are generally combinations of metals and nonmetals.

Molecular compounds are generally composed of nonmetals only.

Only empirical formulas can be written for most ionic compounds. These will be given such that the total positive charge equals the total negative charge.

NaCl

BaCl₂

Mg₃N₂

Naming Inorganic Compounds (see tables of page 60)

Names and Formulas of Ionic Compounds

1. Positive Ions (Cations)

a. Cations formed from metal atoms have the same name as the metal.
b. If a metal can form cations of differing charges, the positive charge is given by a Roman numeral in parentheses following the name of the metal:

\[
\begin{align*}
\text{Fe}^{2+} & \quad \text{Iron(II)} \quad \text{(also ferrous)} \\
\text{Fe}^{3+} & \quad \text{Iron(III)} \quad \text{(Stock system) (also ferric)}
\end{align*}
\]

c. Cations formed from nonmetals atoms have names that end in -ium: Ex: \( \text{NH}_4^+ \), ammonium

2. Negative Ions (Anions)

a. Monatomic (one-atom) anions have names formed by dropping the ending of the name of the element and adding the ending -ide.

\[
\begin{align*}
\text{F}^- & \quad \text{Cl}^- & \quad \text{Br}^- & \quad \text{I}^- & \quad \text{H}^- \\
\text{O}^{2-} & \quad \text{S}^{2-} & \quad \text{N}^{3-} & \quad \text{P}^{3-} & \quad \text{C}^{4-}
\end{align*}
\]

b. Polyatomic (many-atom) anions containing oxygen have names ending in -ate or -ite. The one with more oxygens will end in -ate. The one with less will end in -ite.

\[
\begin{align*}
\text{NO}_2^- & \quad \text{NO}_3^- \\
\text{SO}_3^{2-} & \quad \text{SO}_4^{2-} \\
\text{PO}_4^{3-}
\end{align*}
\]

c. Anions derived by adding \( \text{H}^+ \) to an oxyanion are named by adding as a prefix the word hydrogen or dihydrogen, as appropriate.

\[
\begin{align*}
\text{HCO}_3^- \\
\text{HSO}_4^- \\
\text{H}_2\text{PO}_4^{3-}
\end{align*}
\]
3. Ionic Compounds

Names of ionic compounds are the cation name followed by the anion name:

- BaBr$_2$  
  barium bromide

- Al(NO$_3$)$_3$  
  Aluminum nitrate

- Cu(ClO$_4$)$_2$  
  copper (II) perchlorate (or cupric perchlorate)

Formula from names

- Aluminum oxide  
  Copper(II) chloride

- Silver oxide  
  Zinc bromide

Names and Formulas of Binary Molecular Compounds

1. The name of the element farthest to the left in the periodic table is usually written first.

2. If both elements are in the same group in the periodic table, the lower one is named first.

3. The name of the second element is given an -ide ending.

4. Greek prefixes are used to indicate the number of atoms of each element. The prefix mono- is never used with the first element. When the prefix ends in a or o and the name of the second element begins with a vowel (such as oxide), the a or o is often dropped.

- NF$_3$  
  P$_2$O$_5$

- N$_2$O$_4$  
  S$_2$F$_{10}$
Acids and Bases

Names and Formulas of Acids

Acids - a substance that yields hydrogen ions (H+) when dissolved in water.

1. Acids Based on Anions Whose Name End in -ide. Anions whose names end in -ide have associated acids that have the hydro- prefix and an -ic ending as in the following examples.

<table>
<thead>
<tr>
<th>Anion</th>
<th>Corresponding Acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl(^-) (chloride)</td>
<td>HCl (hydrochloric acid)</td>
</tr>
<tr>
<td>S(^2-) (sulfide)</td>
<td>H(_2)S (hydrosulfuric acid)</td>
</tr>
</tbody>
</table>

2. Acids Based on Anions Whose Names End in -ate or -ite. Anions whose names end in -ate have associated acids with an -ic ending, whereas anions whose names end in -ite have acids with an -ous ending.

<table>
<thead>
<tr>
<th>oxoanion</th>
<th>Oxoacid</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hypochlorite, ClO(^-)</td>
<td>Hypochlorous acid, HClO</td>
</tr>
<tr>
<td>Chlorite, ClO(_2)(^-)</td>
<td>Chlorous acid, HClO(_2)</td>
</tr>
<tr>
<td>Chlorate, ClO(_3)(^-)</td>
<td>Chloric acid, HClO(_3)</td>
</tr>
<tr>
<td>Perchlorate, ClO(_4)(^-)</td>
<td>Perchloric acid, HClO(_4)</td>
</tr>
</tbody>
</table>

This pattern also applies to Br and I.

If all of the hydrogens are not removed, you must indicate the number of hydrogens present.

Ex: \(\text{H}_3\text{PO}_4\) phosphoric acid

\(\text{NaH}_2\text{PO}_4\) sodium dihydrogen phosphate

\(\text{Na}_2\text{HPO}_4\) sodium hydrogen phosphate

\(\text{Na}_3\text{PO}_4\) sodium phosphate
**Bases** - substance that yield hydroxide ions (OH\(^-\)) when dissolved in water.

Sodium hydroxide, NaOH

Potassium hydroxide, KOH

Barium hydroxide, Ba(OH)\(_2\)

Ammonia, NH\(_3\).

\[ \text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{OH}^- (\text{aq}) \]

**Hydrates** - contain specific number of water molecules.

BaCl\(_2\)-2H\(_2\)O

MgSO\(_4\)-7H\(_2\)O

**Organic compounds**

The simplest are hydrocarbons, consisting of hydrogen and carbon.

One class of hydrocarbons is the alkanes.

<table>
<thead>
<tr>
<th>Number of carbons</th>
<th>Name</th>
<th>Molecular formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Methane</td>
<td>CH(_4)</td>
</tr>
<tr>
<td>2</td>
<td>Ethane</td>
<td>CH(_3)CH(_3) or C(_2)H(_6)</td>
</tr>
<tr>
<td>3</td>
<td>Propane</td>
<td>CH(_3)CH(_2)CH(_3) or C(_3)H(_8)</td>
</tr>
<tr>
<td>4</td>
<td>Butane</td>
<td>CH(_3)(CH(_2))(_2)CH(_3) or C(_4)H(_10)</td>
</tr>
<tr>
<td>5</td>
<td>Pentane</td>
<td>CH(_3)(CH(_2))(_3)CH(_3) or C(_5)H(_12)</td>
</tr>
<tr>
<td>6</td>
<td>Hexane</td>
<td>CH(_3)(CH(_2))(_4)CH(_3) or C(_6)H(_14)</td>
</tr>
<tr>
<td>7</td>
<td>Heptane</td>
<td>CH(_3)(CH(_2))(_5)CH(_3) or C(_7)H(_16)</td>
</tr>
<tr>
<td>8</td>
<td>Octane</td>
<td>CH(_3)(CH(_2))(_6)CH(_3) or C(_8)H(_18)</td>
</tr>
<tr>
<td>9</td>
<td>Nonane</td>
<td>CH(_3)(CH(_2))(_7)CH(_3) or C(_9)H(_20)</td>
</tr>
<tr>
<td>10</td>
<td>Decane</td>
<td>CH(_3)(CH(_2))(_8)CH(_3) or C(_10)H(_22)</td>
</tr>
</tbody>
</table>
Functional Groups

Replacing an H atom in methane with different functional groups produce new compounds.

\[
\begin{align*}
\text{methane} & \quad \begin{array}{c}
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\end{array} \\
\text{methanol} & \quad \begin{array}{c}
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\end{array} \\
\text{methylamine} & \quad \begin{array}{c}
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\end{array} \\
\text{acetic acid} & \quad \begin{array}{c}
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\end{array}
\end{align*}
\]