Chapter 2
Atoms, Molecules, and Ions

When learning a new language:
- Start with the alphabet
- Then, form words
- Finally, form more complex structures such as sentences

Chemistry has an alphabet and a language; in this chapter, the fundamentals of the language of chemistry will be introduced.

Outline
- Atoms and Atomic Theory
- Components of the Atom
- Introduction to the Periodic Table
- Molecules and Ions
- Formulas of Ionic Compounds
- Names of Compounds

The Language of Chemistry
- This chapter introduces the fundamental language of chemistry
  - Atoms, molecules and ions
  - Formulas
  - Names

The Structure of Matter
- Atoms
  - Composed of electrons, protons and neutrons
- Molecules
  - Combinations of atoms
- Ions
  - Charged particles

Atoms and Atomic Theory
- An element is composed of tiny particles called atoms
- All atoms of the same element have the same chemical properties
- In an ordinary chemical reaction
  - There is a change in the way atoms are combined with each other
  - Atoms are not created or destroyed
  - Compounds are formed when two or more atoms of different element combine
Fundamental Laws of Matter

- There are three fundamental laws of matter
  - Law of conservation of mass
    - Matter is conserved in chemical reactions
  - Law of constant composition
    - Pure water has the same composition everywhere
  - Law of multiple proportions
    - Compare Cr₂O₃ to CrO₃
    - The ratio of Cr:O between the two compounds is a small whole number

Components of the Atom

- Atomic theory raised more questions than it answered
  - Could atoms be broken down into smaller particles
  - 100 years after atomic theory was proposed, the answers were provided by experiment

Fundamental Experiments

- J.J. Thomson, Cavendish Laboratories, Cambridge, England
- Ernest Rutherford
  - McGill University, Canada
  - Manchester and Cambridge Universities, England
Electrons

- First evidence for subatomic particles came from the study of the conduction of electricity by gases at low pressures
  - J.J. Thomson, 1897
  - Rays emitted were called cathode rays
  - Rays are composed of negatively charged particles called electrons
  - Electrons carry unit negative charge (-1) and have a very small mass (1/2000 the lightest atomic mass)

The Electron and the Atom

- Every atom has at least one electron
- Atoms are known that have one hundred or more electrons
- There is one electron for each positive charge in an atom
- Electrical neutrality is maintained

Protons and Neutrons – The Nucleus

- Ernest Rutherford, 1911
- Bombardment of gold foil with $\alpha$ particles (helium atoms minus their electrons)
  - Expected to see the particles pass through the foil
  - Found that some of the alpha particles were deflected by the foil
  - Led to the discovery of a region of heavy mass at the center of the atom

Figure 2.2 – J.J. Thomson and Ernest Rutherford

Figure 2.3 – Cathode Ray Apparatus

Figure 2.4 – Rutherford Backscattering
Nuclear Particles

1. Protons
   • Mass nearly equal to the H atom
   • Positive charge
2. Neutrons
   • Mass slightly greater than that of the proton
   • No charge

Mass and the Atom

• More than 99.9% of the atomic mass is concentrated in the nucleus
• The volume of the nucleus is much smaller than the volume of the atom

Table 2.1 – Subatomic Particles

<table>
<thead>
<tr>
<th>Particle</th>
<th>Location</th>
<th>Relative Charge</th>
<th>Relative Mass*</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton</td>
<td>Nucleus</td>
<td>+1</td>
<td>1.00728</td>
</tr>
<tr>
<td>Neutron</td>
<td>Nucleus</td>
<td>0</td>
<td>1.00867</td>
</tr>
<tr>
<td>Electron</td>
<td>Outside nucleus</td>
<td>−1</td>
<td>0.00055</td>
</tr>
</tbody>
</table>

*These are expressed in atomic mass units (Chapter 3).

Terminology

• Atomic number, Z
  • Number of protons in the atom
• Mass number, A
  • Number of protons plus number of neutrons

Isotopes

• Isotopes are two atoms of the same element
  • Same atomic number
• Different mass numbers
  • Number of neutrons is A − Z
  • Number of neutrons differs between isotopes

Nuclear symbolism

\[ ^{A}_{Z}X \]

• A is the mass number
• Z is the atomic number
• X is the chemical symbol
Isotopes of hydrogen

- $^1$H, $^2$H, $^3$H
- Hydrogen, deuterium, tritium
- Different masses

Note that some of the ice is at the bottom of the glass – this is $^2$H$_2$O

Example 2.1

Radioactivity

- Radioactive isotopes are unstable
- These isotopes decay over time
- Emit other particles and are transformed into other elements
- Radioactive decay is not a chemical process!
- Particles emitted
  - High speed electrons: $\beta$ (beta) particles
  - Alpha ($\alpha$) particles: helium nuclei
  - Gamma ($\gamma$) rays: high energy light

Nuclear Stability

- Nuclear stability depends on the neutron/proton ratio
  - For light elements, n/p is approximately 1
  - For heavier elements, n/p is approximately 1.4/1
- The belt of stability

Figure 2.5 – The Nuclear Belt of Stability
Periods and Groups
- Horizontal rows are *periods*
  - First period is H and He
  - Second period is Li-Ne
  - Third Period is Na-Ar
- Vertical columns are *groups*
  - IUPAC convention: use numbers 1-18

Blocks in the Periodic Table
- Main group elements
  - 1, 2, 13-18
- Transition Metals
  - 3-12
- Post-transition metals
  - Elements in groups 13-15 to the right of the transition metals
  - Ga, In, Tl, Sn, Pb, Bi

Families with Common Names
- Alkali Metals, Group 1
- Alkaline Earth Metals, Group 2
- Halogens, Group 17
- Noble Gases, Group 18

Importance of Families
- Elements within a family have similar chemical properties
  - Alkali metals are all soft, reactive metals
  - Noble gases are all relatively unreactive gases; He, Ne and Ar do not form compounds

Arrangement of Elements
- Periods
  - Arranged by increasing atomic number
- Families
  - Arranged by chemical properties

Mendeleev
- Dmitri Mendeleev, 1836-1907
- Arranged elements by chemical properties
  - Left space for elements unknown at the time
  - Predicted detailed properties for elements as yet unknown
    - Sc, Ga, Ge
    - By 1886, all these elements had been discovered, and with properties similar to those he predicted
Metals and Nonmetals
- Diagonal line starting with B separates the metals from the nonmetals
- Elements along this diagonal have some of the properties of metals and some of the properties of nonmetals
- Metalloids
  - B, Si, Ge, As, Sb, Te

A Look at the Sulfur Group
- Sulfur (nonmetal), antimony (metalloid) and silver (metal)

Biological View of the Periodic Table
- “Good guys”
  - Essential to life
  - Carbon, hydrogen, oxygen, sulfur and others
- “Bad guys”
  - Toxic or lethal
  - Some elements are essential but become toxic at higher concentrations
  - Selenium

Molecule
- Two or more atoms may combine to form a molecule
- Atoms involved are often nonmetals
- Covalent bonds are strong forces that hold the atoms together
- Molecular formulas
  - Number of each atom is indicated by a subscript
- Examples
  - Water, H$_2$O
  - Ammonia, NH$_3$

Structural Formulas
- Structural formulas show the bonding patterns within the molecule

Figure 2.8 – Biologically Important and Toxic Elements

Selenium, metalloids, and essential elements in the body.
**Structural Formulas**

- Condensed structural formulas suggest the bonding pattern and highlight specific parts of a molecule, such as the reactive group of atoms.

**Example 2.2**

Gibbs the molecular formulas of (a) ethyl alcohol and (b) ethylamine.

**Strategy**
- Find the molecular formulas, simply add up the atoms of each type and use the same as subscripts in the formulas.

**Solution**
- (a) **CH₃CH₂OH**
- (b) **CH₃CH₂NH₂**

**Reality Check**
- Note that although molecular formulas give the composition of the molecule, they reveal nothing about the way the atoms fit together. In this sense they are less useful than structural formulas.

**Example 2.3**

Increase the following problem about the one described.

**Strategy**
- Find the molecular formulas, simply add up the atoms of each type and use the same as subscripts in the formulas.

**Solution**
- (a) **Na⁺**
- (b) **O₂⁻**

**Reality Check**
- Note that although molecular formulas give the composition of the molecule, they reveal nothing about the way the atoms fit together. In this sense they are less useful than structural formulas.

**Example 2.4**

- When atoms or molecules lose or gain electrons, they form charged particles called **ions**.
  - **Na⁺** + **e⁻** → **Na**
  - **O₂⁻** → **O** + 2e⁻

- Positively charged ions are called **cations**.
- Negatively charged ions are called **anions**.

**Polyatomic Ions**

- Groups of atoms may carry a charge; these are the polyatomic ions.
  - **OH⁻**
  - **NH₄⁺**

**Ball and Stick Models**

- Water molecule
- Ammonia molecule
- Methane molecule
Ionic Compounds

- Compounds can form between anions and cations
- Sodium chloride, NaCl
  - Sodium cations and chloride ions associate into a continuous network

Forces Between Ions

- Ionic compounds are held together by strong forces
- Electrostatic attraction of + and – for each other
- Compounds are usually solids at room temperature
  - High melting points
  - May be water-soluble

Solutions of Ionic Compounds

- When an ionic compound dissolves in water, the ions are released from each other
- Presence of ions in the solution leads to electrical conductivity
  - Strong electrolytes
- When molecular compounds dissolve in water, no ions are formed
- Without ions, solution does not conduct electricity
  - Nonelectrolytes

Example 2.4

Conceptual

The structure of a water solution of KNO₃ containing equal numbers of K⁺ and NO₃⁻ ions might be represented as

Calculating the formula:

- Sodium chloride, CaCl₂
  - Ca²⁺
  - Two Cl⁻ ions are required for charge balance

Figure 2.12 – Electrical Conductivity

- For electrical current to flow and light the bulb, the solution in which the electrodes are immersed must contain ions, which carry electrical charge.

(a) The solution of pure water does not contain ions and thus does not light the bulbs.
(b) The solution of sucrose (table sugar) and pure water also lacks ions, and thus does not light the bulbs.
(c) The solution of sodium chloride (NaCl) and pure water does contain ions, and thus lights the bulbs.

Formulas of Ionic Compounds

- Charge balance
  - Each positive charge must have a negative charge to balance it
  - Calcium chloride, CaCl₂
    - Ca²⁺
    - Two Cl⁻ ions are required for charge balance
### Noble Gas Connections

- Atoms that are close to a noble gas (group 18) form ions that contain the same number of electrons as the neighboring noble gas atom.
- Applies to Groups 1, 2, 16 and 17, plus Al ($\text{Al}^{3+}$) and N ($\text{N}^{3-}$).

<table>
<thead>
<tr>
<th>Group</th>
<th>No. of Electrons in Atom</th>
<th>Charge of Ion Formed</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>2 more than noble gas atom</td>
<td>$-1$</td>
<td>$\text{Na}^+$, $\text{K}^+$</td>
</tr>
<tr>
<td>2</td>
<td>2 more than noble gas atom</td>
<td>$-2$</td>
<td>$\text{Mg}^{2+}$, $\text{Ca}^{2+}$</td>
</tr>
<tr>
<td>16</td>
<td>2 less than noble gas atom</td>
<td>$-2$</td>
<td>$\text{O}^{2-}$, $\text{S}^{2-}$</td>
</tr>
<tr>
<td>17</td>
<td>2 less than noble gas atom</td>
<td>$1$</td>
<td>$\text{F}^-$, $\text{Cl}^-$</td>
</tr>
</tbody>
</table>

### Cations of Transition and Post-Transition Metals

- Iron
  - Commonly forms $\text{Fe}^{2+}$ and $\text{Fe}^{3+}$
- Lead
  - Commonly forms $\text{Pb}^{2+}$ and $\text{Pb}^{4+}$

### Polyatomic Ions

- There are only two common polyatomic cations: $\text{NH}_4^+$ and $\text{Hg}_2^{2+}$
- All other common polyatomic ions are anions.

### Table 2.2 – Polyatomic ions

<table>
<thead>
<tr>
<th>No. of Electrons</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>$\text{NH}_4^+$ (ammonium), $\text{Hg}_2^{2+}$ (mercury(II))</td>
</tr>
<tr>
<td>2</td>
<td>$\text{NH}_4^+$ (ammonium), $\text{Hg}_2^{2+}$ (mercury(II))</td>
</tr>
<tr>
<td>3</td>
<td>$\text{NH}_4^+$ (ammonium), $\text{Hg}_2^{2+}$ (mercury(II))</td>
</tr>
</tbody>
</table>

### Example 2.5

**Predict the formula of the ionic compound**

(a) formed by barium with iodine.
(b) containing a transition metal with a $+1$ charge in period 4 and Group 11 and oxide ions.
(c) containing an alkaline earth in period 5 and halogens.
(d) containing ammonium and phosphate ions.

**Strategy**

First identify the elements, then (using Table 2.2 if necessary) identify the charges of the cation and anion. Finally, balance positive and negative charges.

**SOLUTION**

(a) Barium is in Group 2, so it forms a $+2$ ion. Iodine is in Group 17, forming the $1-$ ion. The formula is then $\text{BaI}_2$.

(b) The cation is $\text{Ca}^{2+}$, the anion $\text{O}^{2-}$. Two cations are required to balance one anion. Hence, the formula is $\text{Ca}_2\text{O}_3$.

(c) The element is $\text{Fe}$ (iron), which is in Group 2 and so forms the $\text{Fe}^{2+}$ ion. Nitrogen as a ion has a $-3$ charge. The formula thus is $\text{Fe}(\text{NO}_3)_3$.

(d) From Table 2.2, the ions are $\text{NH}_4^+$ and $\text{PO}_4^{3-}$. The formula is therefore $\text{NH}_4\text{PO}_4$.

### Names of Compounds - Cations

- Monatomic cations take the name from the metal from which they form.
  - $\text{Na}^+$, sodium ion
  - $\text{K}^+$, potassium ion

- If more than one charge is possible, a Roman numeral is used to denote the charge.
  - $\text{Fe}^{2+}$, iron(II) ion
  - $\text{Fe}^{3+}$, iron(III) ion
Names of Compounds - Anions

- Monatomic anions are named by adding –ide to the stem of the name of the element from which they form
  - Oxygen becomes oxide, $O^2-$
  - Sulfur becomes sulfide, $S^2-$
- Polyatomic ions are given special names (see table 2.3, p. 39)

Oxoanions

- When a nonmetal forms two oxoanions
  - -ate is used for the one with the larger number of oxygens
  - -ite is used for the one with the smaller number of oxygens
- When a nonmetal forms more than two oxoanions, prefixes are used
  - per (largest number of oxygens)
  - hypo (smallest number of oxygens)

Ionic Compounds

- Combine the name of the cation with name of the anion
  - $\text{Cr(NO}_3\text{)}_3$, chromium(III) nitrate
  - $\text{SnCl}_2$, tin(II) chloride

Table 2.3 – Oxoanions of Nitrogen, Sulfur and Chlorine

<table>
<thead>
<tr>
<th>Nitrogen</th>
<th>Sulfur</th>
<th>Chlorine</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{NO}_3^-$ nitrate</td>
<td>$\text{SO}_4^{2-}$ sulfate</td>
<td>$\text{ClO}_4^-$ perchlorate</td>
</tr>
<tr>
<td>$\text{NO}_2^-$ nitrate</td>
<td>$\text{SO}_3^{2-}$ sulfite</td>
<td>$\text{ClO}_3^-$ chlorate</td>
</tr>
</tbody>
</table>

Example 2.6

Example 2.6: Name the following ionic oxoanions.
(a) $\text{CrS}$  (b) $\text{Al(NO}_3\text{)}_3$  (c) $\text{TeCl}_4$

Strategy: To name an ionic oxoanion, you must know the rules for naming individual ions, as discussed previously.

Solution:
(a) chromium sulfide  (b) aluminum nitrate  (c) tellurium tetrachloride

Binary Molecular Compounds

- Unlike ionic compounds, there is no simple way to deduce the formula of a binary molecular compound
- Systematic naming
  1. The first word is the name of the first element in the formula, with a Greek prefix if necessary
  2. The second word consists of
    - The appropriate Greek prefix
    - The stem of the name of the second element
    - The suffix -ide
**Some Examples**

- Binary nonmetallic compounds
  - \( \text{N}_2\text{O}_5 \), dinitrogen pentaoxide
  - \( \text{N}_2\text{O}_4 \), dinitrogen tetraoxide
  - \( \text{NO}_2 \), nitrogen dioxide
  - \( \text{N}_2\text{O}_3 \), dinitrogen trioxide
  - \( \text{NO} \), nitrogen oxide
  - \( \text{N}_2\text{O} \), dinitrogen oxide
- Common names
  - \( \text{H}_2\text{O} \), water
  - \( \text{H}_2\text{O}_2 \), hydrogen peroxide

**Common Molecular Compounds**

To illustrate these rules, consider the names of the several oxides of nitrogen:

- \( \text{N}_2\text{O}_5 \), dinitrogen pentaoxide
- \( \text{N}_2\text{O}_4 \), dinitrogen tetraoxide
- \( \text{NO}_2 \), nitrogen dioxide
- \( \text{N}_2\text{O}_3 \), dinitrogen trioxide
- \( \text{NO} \), nitrogen oxide
- \( \text{N}_2\text{O} \), dinitrogen oxide

When the prefixes \( \text{tetra}, \text{penta}, \text{hexa} \), etc., are followed by the letter \( \text{O} \), the \( \text{O} \) is often dropped. For example, \( \text{NO}_2 \) is often referred to as dinitrogen peroxide.

Many of the best known binary compounds of the elements have acquired common names. These are widely—and in some cases exclusively—used. Examples include:

- \( \text{H}_2\text{O} \), water
- \( \text{H}_2\text{O}_2 \), hydrogen peroxide
- \( \text{H}_2\text{O}_3 \), water peroxide
- \( \text{H}_2\text{O}_4 \), water dioxide
- \( \text{H}_2\text{O}_5 \), water trioxide
- \( \text{H}_2\text{O}_6 \), water tetraoxide
- \( \text{H}_2\text{O}_7 \), water pentaoxide

**Table 2.4 - Greek Prefixes**

<table>
<thead>
<tr>
<th>Number</th>
<th>Prefix</th>
<th>Number</th>
<th>Prefix</th>
<th>Number</th>
<th>Prefix</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>tri</td>
<td>6</td>
<td>hexa</td>
<td>9</td>
<td>nona</td>
</tr>
<tr>
<td>3</td>
<td>tetra</td>
<td>7</td>
<td>hepta</td>
<td>10</td>
<td>decra</td>
</tr>
</tbody>
</table>

*The prefix \( \text{tri} \) is seldom used.

**Example 2.7**

Given the names of the following molecules:

- (a) \( \text{SO}_2 \)
- (b) \( \text{SO}_3 \)
- (c) \( \text{PCl}_3 \)
- (d) \( \text{Cl}_2\text{O}_4 \)

**Strategy**

Start with the prefix denoting the number of atoms (if there is more than one) of the first element followed by the name of that element. Repeat for the second element, ending with the suffix \( \text{-ide} \).

**Solution**

- (a) sulfur dioxide
- (b) sulfur trioxide
- (c) phosphorus trichloride
- (d) chlorine heptaoxide

**Acids**

- Acids ionize to form \( \text{H}^+ \) ions
- Hydrogen and chlorine
  - As a molecule, \( \text{HCl} \) is hydrogen chloride
  - When put in water, \( \text{HCl} \) is hydrochloric acid

**Common Acids**

<table>
<thead>
<tr>
<th>Pure Substance</th>
<th>Water Solution</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{HCl}_g )</td>
<td>Hydrochloric acid</td>
</tr>
<tr>
<td>( \text{HCl}_l )</td>
<td>( \text{H}^+, \text{Cl}^-. )</td>
</tr>
<tr>
<td>( \text{HBr}_g )</td>
<td>Hydrobromic acid</td>
</tr>
<tr>
<td>( \text{HBr}_l )</td>
<td>( \text{H}^+, \text{Br}^- )</td>
</tr>
<tr>
<td>( \text{HI}_g )</td>
<td>Hydroiodic acid</td>
</tr>
<tr>
<td>( \text{HI}_l )</td>
<td>( \text{H}^+, \text{I}^- )</td>
</tr>
</tbody>
</table>
Oxoacids

- Two common oxoacids
  - HNO₃, nitric acid
  - H₂SO₄, sulfuric acid

Oxoacids of Chlorine

Most acids contain oxygen in addition to hydrogen atoms. Such species are referred to as oxoacids. Two oxoacids that you are likely to encounter in the general chemistry laboratory are:

\[
\begin{align*}
HNO_3 & \quad \text{nitric acid} \\
H_2SO_4 & \quad \text{sulfuric acid}
\end{align*}
\]

The names of oxoacids are simply related to those of the corresponding oxoanions. The acid suffix of the anion is replaced by -ic in the acid. In a similar way, the sulfide -ide is replaced by the suffix -ious. The prefixes peri- and hypo- found in the name of the anion are retained in the name of the acid:

\[
\begin{align*}
\text{ClO}_3^- & \quad \text{perchlorate} \\
\text{ClO}_4^- & \quad \text{chlorate} \\
\text{ClO}_2^- & \quad \text{chlorite} \\
\text{ClO}^- & \quad \text{hypochlorite} \\
\end{align*}
\]

Example 2.8

Example 2.8: Give the names of

(a) HCl(g)  (b) HNO₃(aq)  (c) H₂SO₄(aq)  (d) HBr(aq)

Strategy: In (a), note that the species is a gas, not an acid. In (b) and (c), refer back to Table 2.3 for the names of the oxoanion. In (d), note that it is now an acid.

SOLUTION

(a) hydrogen chloride gas  (b) nitric acid
(c) sulfuric acid  (d) hydrobromic acid

Key Concepts

1. Relate a nuclear symbol to the numbers of protons and neutrons in the nucleus.
2. Relate elements and the periodic table.
3. Relate structural, condensed, and molecular formulas.
4. Relate the ionic charge to the number of electrons.
5. Write the formula of ionic compounds from the charges on the ions.

Key Concepts

6. Relate names to formulas for
   - Ionic compounds
   - Binary molecular compounds
   - Oxoanions and oxoacids