Chapter 2: Atoms, Molecules, and Ions

Learning outcomes:
- Learn the basic postulates of Dalton’s atomic theory.
- Describe the key experiments that led to the discovery of electrons and the nuclear model of the atom.
- Describe the structure of the atom in terms of protons, neutrons, and electrons and express the relative electrical charges and masses of these subatomic particles.
- Use chemical symbols together with atomic number and mass number to express the subatomic composition of isotopes.
- Calculate the atomic weight of an element from the masses of individual atoms and a knowledge of natural abundances.
- Describe how elements are organized in the periodic table by atomic number and by similarities in chemical behavior, giving rise to periods and groups.
- Identify the locations of metals and nonmetals in the periodic table.
- Distinguish between molecular substances and ionic substances in terms of their composition.
- Distinguish between empirical formulas and molecular formulas.
- Describe how molecular formulas and structural formulas are used to represent the compositions of molecules.
- Explain how ions are formed by the gain or loss of electrons and use the periodic table to predict the charges of common ions.
- Write the empirical formulas of ionic compounds, given the charges of their component ions.
- Write the name of an ionic compound given its chemical formula or write the chemical formula given its name.
- Name or write chemical formulas for binary inorganic compounds and for acids.
- Identify organic compounds and name simple alkanes and alcohols.

Atomic Theory of Matter

- Democritus (Greek philosopher) believed that there was a smallest particle—“atomos” (uncuttable)—that made up all of nature.
- Experiments in the eighteenth and nineteenth centuries led to an organized atomic theory by John Dalton in the early 1800s, which explained several laws known at that time:
  - The law of constant composition
  - The law of conservation of mass
  - The law of multiple proportions

Law of Constant Composition
- Also known as the law of definite proportions.
- The elemental composition of a pure substance never varies.
- In a given compound, the relative numbers and kinds of atoms are constant.
- Basis of Dalton’s Postulate #4
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 the ratio of small whole numbers.

Law of Conservation of Mass (Matter)

- The total mass of substances present at the end of a chemical process is the same as the mass of substances present before the process took place.
- Basis of Dalton’s Postulate #3
- Can’t create matter in a chemical reaction!

Dalton’s Atomic Theory

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

2. All atoms of a given element are identical, but the atoms of one element are different from the atoms of all other elements.

3. Atoms of one element cannot be changed into atoms of a different element 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.
Discovery of Atomic Structure

Mid 1800’s – scientists studied electrical discharge through partially evacuated tubes.

Path of cathode rays deflected by presence of a magnet.

Proposed in 1897 that Cathode Rays were actually particles (negatively charged) that we now know are electrons.

Thomson measured the charge/mass ratio of the electron to be $1.76 \times 10^8$ coulombs/gram (C/g).
• Once the charge/mass ratio of the electron was known, determination of either the charge or the mass of an electron would yield the other.

• Robert Millikan (1909) determined the charge on the electron, equal to $1.60 \times 10^{-19}$ C.

• Mass of an $e^-$ could be calculated as $9.11 \times 10^{-28}$ g.

Radioactivity:

The spontaneous emission of radiation by an atom.

First observed by Henri Becquerel.  
(1852-1908)

Also studied by Marie (1867-1934) and Pierre Curie (1857-1906). Discovered Po and Ra. Suggested that atoms of certain substances can disintegrate.
Radioactivity

Three types of radiation were discovered by Ernest Rutherford:

\( \alpha \) particles positively charged \((+2)\), large mass
\( \beta \) particles negatively charged \((-1)\), small mass
\( \gamma \) rays no charge, no mass

The Atom, circa 1900

- “Plum pudding” model, put forward by Thomson.
- Positive sphere of matter with negative electrons imbedded in it.

Sir Joseph John Thomson (1856-1940)
Discovery of the Nucleus

Ernest Rutherford shot $\alpha$ particles at a thin sheet of gold foil and observed the pattern of scatter of the particles.

The Nuclear Atom

• Since some particles were deflected at large angles, Thomson’s model could not be correct.

• Rutherford’s nuclear model of the atom: all of the positive charge and most of the mass is concentrated at the center – the nucleus. Electrons occupy the rest of the space (volume) of the atom.
Subatomic Particles

Neutrons were discovered by James Chadwick in 1932.

Neutrons were discovered by James Chadwick in 1932.

TABLE 2.1 Comparison of the Proton, Neutron, and Electron

<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>

Atomic and Mass Numbers

Mass number (number of protons plus neutrons)

Atomic number (number of protons or electrons)

Elements are represented by a one or two letter symbol, for which the first letter is always capitalized. C is the symbol for carbon.

Atomic number: equal to the number of protons in the nucleus. All atoms of the same element have the same number of protons. Denoted by “Z”.

Mass number: equal to the sum of the number of protons and neutrons for an atom.
The mass of an atom in atomic mass units (amu) is the total number of protons and neutrons in the atom.

A carbon (C) atom with 6 protons and 6 neutrons is assigned a mass of exactly 12 amu.

Atomic mass unit (amu) is one-twelfth of the mass of an atom of carbon with 6 protons and 6 neutrons.

\[
1 \text{ amu} = 1.66054 \times 10^{-24} \text{ g}
\]

\[
1 \text{ g} = 6.02214 \times 10^{23} \text{ amu}
\]

Isotopes

- Isotopes are atoms of the same element with different masses.
- Isotopes have different numbers of neutrons, thus different mass numbers.

\[
\begin{array}{c}
\text{11}\ _6^6\text{C} \\
\text{12}\ _6^6\text{C} \\
\text{13}\ _6^6\text{C} \\
\text{14}\ _6^6\text{C}
\end{array}
\]

TABLE 2.2  Some Isotopes of Carbon\(^a\)

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Number of Protons</th>
<th>Number of Electrons</th>
<th>Number of Neutrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>(^{11}\text{C})</td>
<td>6</td>
<td>6</td>
<td>5</td>
</tr>
<tr>
<td>(^{12}\text{C})</td>
<td>6</td>
<td>6</td>
<td>6</td>
</tr>
<tr>
<td>(^{13}\text{C})</td>
<td>6</td>
<td>6</td>
<td>7</td>
</tr>
<tr>
<td>(^{14}\text{C})</td>
<td>6</td>
<td>6</td>
<td>8</td>
</tr>
</tbody>
</table>

\(^a\) Almost 99% of the carbon found in nature is \(^{12}\text{C}\).
Atomic Mass

Atomic and molecular masses (actually the mass to charge ratio) can be measured with great accuracy with a mass spectrometer.

Atomic Masses of the Elements

- **Isotopic mass** is the mass in amu (u), of a particular isotope of an element.
- Different isotopes of an element all react essentially the same, so a *weighted average* of isotopic masses can be used in calculations.
- The **atomic weight** is the weighted average mass, of the naturally occurring element. It is calculated from the isotopes of an element weighted by their relative abundances.

\[
\text{Atomic weight} = \sum \left[ (\text{isotope mass}) \times (\text{fractional isotope abundance}) \right]
\]

Average atomic mass is known as the atomic weight.
Boron has two naturally occurring isotopes, $^{10}\text{B}$ and $^{11}\text{B}$ with isotopic mass 10.01293 and 11.0093 amu, respectively. The average atomic mass of boron is 10.811 amu. Determine the fractional abundance of $^{10}\text{B}$.
A repeating pattern of chemical and physical properties is observed.

- **Law of chemical periodicity**: the properties of the elements are periodic functions of atomic number.

- Elements in the same *group* have similar chemical and physical properties.

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### Table 2.3 Names of Some Groups in the Periodic Table

<table>
<thead>
<tr>
<th>Group</th>
<th>Name</th>
<th>Elements</th>
</tr>
</thead>
<tbody>
<tr>
<td>1A</td>
<td>Alkali metals</td>
<td>Li, Na, K, Rb, Cs, Fr</td>
</tr>
<tr>
<td>2A</td>
<td>Alkaline earth metals</td>
<td>Be, Mg, Ca, Sr, Ba, Ra</td>
</tr>
<tr>
<td>6A</td>
<td>Chalcogens</td>
<td>O, S, Se, Te, Po</td>
</tr>
<tr>
<td>7A</td>
<td>Halogens</td>
<td>F, Cl, Br, I, At</td>
</tr>
<tr>
<td>8A</td>
<td>Noble gases</td>
<td>He, Ne, Ar, Kr, Xe, Rn</td>
</tr>
</tbody>
</table>
The Periodic Table: Metals

Metals are on the left side of the periodic table.
- Shiny luster, ductile, malleable
- Conducting heat and electricity
- Solids (except mercury)

Nonmetals are on the right side of the periodic table (with the exception of H).
- Have a wide variety of properties (solids, liquids and gases) and do not conduct electricity well (except C as graphite).

The Periodic Table: Metalloids

Metalloids or semimetals: some physical characteristics of metals and chemical characteristics of nonmetal.

Metalloids border the stair-step line (with the exception of Al and Po).
Chemical Formulas

- The subscript to the right of the symbol of an element tells the number of atoms of that element in one molecule of the compound.
- Molecular compounds often contain only nonmetals.
- The attraction between molecules are often relatively weak, explaining why gases and liquids are common among molecular substances.
- Carbon is typically listed first in the formula, unless C is part of a polyatomic ion.

Diatomic Molecules

Seven elements occur naturally as molecules containing two atoms:
- Hydrogen
- Nitrogen
- Oxygen
- Fluorine
- Chlorine
- Bromine
- Iodine
Types of Formulas

- **Molecular formulas** give the exact number of atoms of each element in a compound.
- **Empirical formulas** give the lowest whole-number ratio of atoms of each element in a compound.
- **Structural formulas** show the order in which atoms are bonded.
- **Perspective drawings** also show the three-dimensional shape of atoms in a compound.

**Ions**

When atoms lose or gain electrons, they become ions.

- **Cations are positive** and are formed by elements on the left side of the periodic table.
- **Anions are negative** and are formed by elements on the right side of the periodic table.

The element’s symbol is followed by a superscript number and a sign that shows the charge on the ion in electron charge units.

If the ionic charge is one unit, the number is often omitted, e.g. Na\(^+\) is the symbol for a sodium ion.
Ionic Bonds

Ionic compounds (such as NaCl) are generally formed between metals and nonmetals.

Writing Formulas

\[ \text{Mg}^{2+} \quad \text{N}^{3-} \quad \rightarrow \quad \text{Mg}_3\text{N}_2 \]

- Because compounds are electrically neutral, one can determine the formula of a binary compound this way:
  - The charge on the cation becomes the subscript on the anion.
  - The charge on the anion becomes the subscript on the cation.
  - If these subscripts are not in the lowest whole-number ratio, divide them by the greatest common factor.
Naming Monoatomic Ions

**Cations**
- metal + “ion” or “cation”
  - $\text{Al}^{3+}$ aluminum ion or aluminum cation
- For metals that have more than one oxidation state, the charge is indicated by Roman numerals.
  - $\text{Fe}^{2+}$ iron (II) ion, $\text{Fe}^{3+}$ iron (III) ion
  - $\text{Bi}^{3+}$ bismuth (III) ion, $\text{Bi}^{5+}$ bismuth (V) ion

**Anions**
- -ide ending
  - $\text{O}^{2-}$ oxide ion, $\text{I}^{-}$ iodide ion, $\text{H}^{+}$ hydride ion

**TABLE 2.4 Common Cations**

<table>
<thead>
<tr>
<th>Charge</th>
<th>Formula</th>
<th>Name</th>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>+1</td>
<td>$\text{H}^{+}$ hydrogen ion</td>
<td>$\text{NH}_4^{+}$ ammonium ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>$\text{Li}^{+}$ lithium ion</td>
<td>$\text{Cu}^{+}$ copper(I) or cuprous ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>$\text{Na}^{+}$ sodium ion</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>$\text{K}^{+}$ potassium ion</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>$\text{Cs}^{+}$ cesium ion</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>$\text{Ag}^{+}$ silver ion</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>+2</td>
<td>$\text{Mg}^{2+}$ magnesium ion</td>
<td>$\text{CO}^{2+}$ cobalt(II) or cobaltous ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>$\text{Ca}^{2+}$ calcium ion</td>
<td>$\text{Cu}^{2+}$ copper(II) or cupric ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>$\text{Sr}^{2+}$ strontium ion</td>
<td>$\text{Fe}^{3+}$ iron(III) or ferric ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>$\text{Ba}^{2+}$ barium ion</td>
<td>$\text{Mn}^{2+}$ manganese(II) or manganous ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>$\text{Zn}^{2+}$ zinc ion</td>
<td>$\text{Hg}^{2+}$ mercury(I) or mercurous ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>$\text{Cd}^{2+}$ cadmium ion</td>
<td>$\text{Hg}^{+}$ mercury(II) or mercuric ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>$\text{Ni}^{2+}$ nickel(II) or nickelous ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>$\text{Pb}^{2+}$ lead(II) or plumbois ion</td>
<td>$\text{Sn}^{2+}$ tin(II) or stannous ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td>+3</td>
<td>$\text{Al}^{3+}$ aluminum ion</td>
<td>$\text{Cr}^{3+}$ chromium(III) ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>$\text{Fe}^{3+}$ iron(III) ion</td>
<td>$\text{Cu}^{3+}$ cupric ion</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

*The ions we use most often in this course are in boldface. Learn them first.*

$\text{Fe}^{2+}$ iron(II) ion, $\text{Cu}^{3+}$ copper(I) ion, $\text{Fe}^{3+}$ ferrous ion, $\text{Cu}^{3+}$ cuprous ion, $\text{Fe}^{3+}$ ferric ion, $\text{Cu}^{3+}$ cupric ion
## TABLE 2.5 Common Anions\(^a\)

<table>
<thead>
<tr>
<th>Charge</th>
<th>Formula</th>
<th>Name</th>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>1−</td>
<td>H(^+)</td>
<td>hydride ion</td>
<td>CH(_3)COO(^-) (or C(_2)H(_3)O(_2)(^-))</td>
<td>acetate ion</td>
</tr>
<tr>
<td></td>
<td>F(^-)</td>
<td>fluoride ion</td>
<td>ClO(_3)(^-)</td>
<td>chlorate ion</td>
</tr>
<tr>
<td></td>
<td>Cl(^-)</td>
<td>chloride ion</td>
<td>ClO(_4)(^-)</td>
<td>perchlorate ion</td>
</tr>
<tr>
<td></td>
<td>Br(^-)</td>
<td>bromide ion</td>
<td>NO(_3)(^-)</td>
<td>nitrate ion</td>
</tr>
<tr>
<td></td>
<td>I(^-)</td>
<td>iodide ion</td>
<td>MnO(_4)(^-)</td>
<td>permanganate ion</td>
</tr>
<tr>
<td></td>
<td>CN(^-)</td>
<td>cyanide ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>OH(^-)</td>
<td>hydroxide ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2−</td>
<td>O(^{2-})</td>
<td>oxide ion</td>
<td>CO(_3)(^2-)</td>
<td>carbonate ion</td>
</tr>
<tr>
<td></td>
<td>O(_2)(^{2-})</td>
<td>peroxide ion</td>
<td>CrO(_4)(^2-)</td>
<td>chromate ion</td>
</tr>
<tr>
<td></td>
<td>S(^{2-})</td>
<td>sulfide ion</td>
<td>Cr(_2)O(_7)(^2-)</td>
<td>dichromate ion</td>
</tr>
<tr>
<td></td>
<td>SO(_4)(^2-)</td>
<td>sulfate ion</td>
<td></td>
<td>sulfate ion</td>
</tr>
<tr>
<td>3−</td>
<td>N(_3)(^{3-})</td>
<td>nitride ion</td>
<td>PO(_4)(^3-)</td>
<td>phosphate ion</td>
</tr>
</tbody>
</table>

\(\text{\(a\)}}\) The ions we use most often are in boldface. Learn them first.
Nomenclature: Ionic Compounds

- Write the name of the cation.
- If the anion is an element, change its ending to -ide; if the anion is a polyatomic ion, simply write the name of the polyatomic ion.
- If the cation can have more than one possible charge (e.g., iron), write the charge as a Roman numeral in parentheses.

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

<table>
<thead>
<tr>
<th>Chemical formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>CaCl₂</td>
<td>calcium chloride</td>
</tr>
<tr>
<td>Al(NO₃)₃</td>
<td>aluminum nitrate</td>
</tr>
<tr>
<td>Cu(ClO₄)₂</td>
<td>copper(II) perchlorate (or cupric perchlorate)</td>
</tr>
</tbody>
</table>

First, write the name of the cation. **magnesium**

Second, if the anion is an element, change its ending to -ide. **nitrogen** → **nitride**

**magnesium nitride**

Chromium(III) oxide, used as a green paint pigment, is composed of Cr³⁺ and O²⁻ ions.

What is the formula of chromium(III) oxide?

Strontium oxide, is composed of Sr²⁺ and O²⁻ ions.

What is the formula for this compound?
Potassium chromate, is a brightly colored compound of chromium. Chromium comes from the Greek word *chroma*, meaning color. The chromate anion is $\text{CrO}_4^{2-}$.

What is the ionic formula for this compound?

Iron(II) phosphate is found in the hydrated form in the mineral Vivianite.

What is the formula for this compound?

**Patterns in Oxyanion Nomenclature**

- When there are **two** oxyanions involving the same element:
  - The one with fewer oxygens ends in **-ite**
    - $\text{NO}_2^-$: nitrite; $\text{SO}_3^{2-}$: sulfite
  - The one with more oxygens ends in **-ate**
    - $\text{NO}_3^-$: nitrate; $\text{SO}_4^{2-}$: sulfate

Most common form of the oxyanion.
Patterns in Oxyanion Nomenclature

• When there are **more than two** oxyanions involving the same element:

  • The one with the second **fewest** oxygens ends in **-ite**
    - ClO\(_2^-\) : chlorite
  • The one with the second **most** oxygens ends in **-ate**
    - ClO\(_3^-\) : chlorate

Patterns in Oxyanion Nomenclature

• The one with the fewest oxygens has the prefix **hypo-** and ends in **-ite**
  - ClO\(^-\) : hypochlorite
• The one with the most oxygens has the prefix **per-** and ends in **-ate**
  - ClO\(_4^-\) : perchlorate
Acid Nomenclature

• If the anion in the acid ends in \textit{ide}, change the ending to \textit{ic acid} and add the prefix \textit{hydro}-:
  - HCl: hydrochloric acid
  - HBr: hydrobromic acid
  - HI: hydroiodic acid

<table>
<thead>
<tr>
<th>Anion</th>
<th>Corresponding Acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl\textsuperscript{-} (chloride)</td>
<td>HCl (hydrochloric acid)</td>
</tr>
<tr>
<td>S\textsuperscript{2-} (sulfide)</td>
<td>H\textsubscript{2}S (hydrosulfuric acid)</td>
</tr>
</tbody>
</table>

Acid Nomenclature

• If the anion in the acid ends in \textit{ate}, change the ending to \textit{ic acid}:
  - HClO\textsubscript{3}: chloric acid
  - HClO\textsubscript{4}: perchloric acid

Name the corresponding acid to the SO\textsubscript{4}\textsuperscript{2-} anion.
**Acid Nomenclature**

- If the anion in the acid ends in *-ite*, change the ending to *-ous acid*:
  - HClO: hypochlorous acid
  - HClO₂: chlorous acid

Name the corresponding acid to the SO₃²⁻ anion.

<table>
<thead>
<tr>
<th>Anion</th>
<th>Corresponding Acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>ClO₄⁻</td>
<td>HClO₄</td>
</tr>
<tr>
<td>ClO₃⁻</td>
<td>HClO₃</td>
</tr>
<tr>
<td>ClO₂⁻</td>
<td>HClO₂</td>
</tr>
<tr>
<td>ClO⁻</td>
<td>HClO</td>
</tr>
</tbody>
</table>

(perchlorate) (perchloric acid)
(chlorate) (chloric acid)
(chlorite) (chlorous acid)
(hypochlorite) (hypochlorous acid)
Nomenclature of Binary Compounds (molecular) Formed Between Nonmetals.

- The less electronegative atom (farthest to left) is usually listed first.
- A Greek prefix is used to denote the number of atoms of each element in the compound (mono- is not used on the first element listed, however.)
- The ending on the more electronegative element is changed to -ide.
  - CCl₄: carbon tetrachloride
  - CO₂: carbon dioxide
  - CO: carbon monoxide
  - N₂O₄: dinitrogen tetroxide
- In nonmetal compounds containing carbon, the C atom is listed first in the chemical formula, followed by H.

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Meaning</th>
</tr>
</thead>
<tbody>
<tr>
<td>mono-</td>
<td>1</td>
</tr>
<tr>
<td>di-</td>
<td>2</td>
</tr>
<tr>
<td>tri-</td>
<td>3</td>
</tr>
<tr>
<td>tetra-</td>
<td>4</td>
</tr>
<tr>
<td>penta-</td>
<td>5</td>
</tr>
<tr>
<td>hexa-</td>
<td>6</td>
</tr>
<tr>
<td>hepta-</td>
<td>7</td>
</tr>
<tr>
<td>octa-</td>
<td>8</td>
</tr>
<tr>
<td>nona-</td>
<td>9</td>
</tr>
<tr>
<td>deca-</td>
<td>10</td>
</tr>
</tbody>
</table>

**Table 2.6 Prefixes Used in Naming Binary Compounds Formed between Nonmetals**

If the prefix ends with a or o and the name of the element begins with a vowel, the two successive vowels are often elided into one:

N₂O₅: dinitrogen pentoxide

<table>
<thead>
<tr>
<th>Compound</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl₂O</td>
<td>dichlorine monoxide</td>
</tr>
<tr>
<td>N₂O₄</td>
<td>dinitrogen tetroxide</td>
</tr>
<tr>
<td>NF₃</td>
<td>nitrogen trifluoride</td>
</tr>
<tr>
<td>P₄S₁₀</td>
<td>tetraphosphorus decasulfide</td>
</tr>
</tbody>
</table>
Summary of Nomenclature of Binary Compounds

Cations
1. Cations formed from metal atoms have the same name as the metal.

2. If a metal can form different cations, the positive charge is indicated by a Roman numeral in parenthesis following the name of the metal.
   - Sometimes the ending –ous and –ic will be used to distinguish lower and higher charged ions, respectively.

3. Cations formed from nonmetal atoms have names that end in –ium.

Anions
1. Names of monatomic anions formed by replacing the ending of the element with –ide.

2. Polyatomic ions (oxyanions) containing oxygen have names that end in –ate or –ite.

3. Anions derived by adding H+ to an oxyanion are named by adding as a prefix the word hydrogen or dihydrogen, as appropriate.

Common Names to Know

<table>
<thead>
<tr>
<th>Compound</th>
<th>Common Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>NH₃</td>
<td>ammonia</td>
</tr>
<tr>
<td>H₂O</td>
<td>water</td>
</tr>
<tr>
<td>CH₄</td>
<td>methane</td>
</tr>
<tr>
<td>C₂H₆</td>
<td>ethane</td>
</tr>
<tr>
<td>C₃H₈</td>
<td>propane</td>
</tr>
<tr>
<td>CH₃OH</td>
<td>methanol</td>
</tr>
<tr>
<td>C₂H₅OH</td>
<td>ethanol</td>
</tr>
<tr>
<td>C₃H₇OH</td>
<td>propanol</td>
</tr>
</tbody>
</table>
Nomenclature of Organic Compounds

- **Organic chemistry** is the study of carbon and has its own system of nomenclature.
- The simplest hydrocarbons (compounds containing only carbon and hydrogen) are **alkanes**.
- The first part of the names just listed correspond to the number of carbons (*meth-* = 1, *eth-* = 2, *prop-* = 3, etc.).
- It is followed by **-ane**.

When a hydrogen in an alkane is replaced with a **functional group**, like –OH in the compounds above, the name is derived from the name of the alkane.
- The ending denotes the type of compound.
  - An **alcohol** ends in **-ol**
Isomers form when two or more molecules have the same chemical formula, but different structures.

1-Propanol and 2-propanol have the oxygen atom connected to different carbon atoms, but both have the formula C₃H₈O.