Chapter 3
Chemical Compounds

An Introduction to Chemistry
by Mark Bishop
Elements, Compounds, and Mixtures

- **Element**: A substance that cannot be chemically converted into simpler substances; a substance in which all of the atoms have the same number of protons and therefore the same chemical characteristics.

- **Compound**: A substance that contains two or more elements, the atoms of these elements always combining in the same whole-number ratio.

- **Mixture**: A sample of matter that contains two or more pure substances (elements and compounds) and has variable composition.
Classification of Matter

Matter
Does it have a constant composition? Can it be described with a chemical formula?

Yes
Pure Substance
Can it be described with a single symbol?

Yes
Element
hydrogen, H₂

No
Compound
water, H₂O

No
Mixture
coffee with cream and sugar
Elements and Compounds

**ELEMENTS**

Hydrogen is composed of molecules with 2 hydrogen atoms.

- **H₂ molecule**

Neon is composed of independent atoms.

- **Neon atom**

Silver exists as an assembly of silver atoms.

- **Silver atom**

**COMPOUNDS**

Water is composed of molecules that contain one oxygen atom and two hydrogen atoms.

- **Water molecule, H₂O**

Sodium chloride exists as an assembly of sodium and chloride ions, always in a one-to-one ratio.

- **Sodium ion**
- **Chloride ion**
Exhaust – a Mixture

- Unburned gasoline
- Water, $H_2O$
- Nitrogen oxides, NO and NO$_2$
- Carbon dioxide, CO$_2$
Covalent Bond Formation

Two hydrogen atoms interact to form one hydrogen molecule.
A link between atoms due to the sharing of two electrons. This bond forms between atoms of two nonmetallic elements.

- If the electrons are shared equally, there is an even distribution of the negative charge for the electrons in the bond, so there is no partial charges on the atoms. The bond is called a **nonpolar covalent bond**.

- If one atom in the bond attracts electrons more than the other atom, the electron negative charge shifts to that atom giving it a partial negative charge. The other atom loses negative charge giving it a partial positive charge. The bond is called a **polar covalent bond**.
Polar Covalent Bond

Electrons shift toward the chlorine atom, forming partial plus and minus charges.

δ+  δ−

Hydrogen attracts electrons less.  Chlorine attracts electrons more.
• The attraction between cation and anion.
• Atoms of nonmetallic elements often attract electrons so much more strongly than atoms of metallic elements that one or more electrons are transferred from the metallic atom (forming a positively charged particle or cation), to the nonmetallic atom (forming a negatively charged particle or anion).
• For example, an uncharged chlorine atom can pull one electron from an uncharged sodium atom, yielding $\text{Cl}^-$ and $\text{Na}^+$. 
Ionic Bond Formation

- **Chlorine gas, Cl₂**
- **Sodium metal, Na**
- **Sodium ion, Na⁺**
- **Chlorine atom, Cl⁻**
- **Chlorine ion, Cl⁻**

Each Na atom loses one electron and gets smaller.
Each Cl atom gains one electron and gets larger.

Ionic bond, an attraction between a cation and an anion.
Sodium Chloride, NaCl, Structure

Each chloride anion is surrounded by 6 cations.

Each sodium cation is surrounded by 6 anions.

Ball-and-stick model

Space-filling model

Salt (sodium chloride)
**Bond Types**

**Nonpolar Covalent Bond**
Equal sharing of electrons

Both atoms attract electrons equally (or nearly so).
No significant charges form.

**Polar Covalent Bond**
Unequal sharing of electrons

Partial positive charge
This atom attracts electrons more strongly.
Partial negative charge.

**Ionic Bond**
Strong attraction between positive and negative charges.

This atom loses one or more electrons and gains a positive charge.

This atom attracts electrons so much more strongly than the other atom that it gains one or more electrons and gains a negative charge.
Types of Compounds

- All nonmetallic atoms usually leads to all covalent bonds, which from molecules. These compounds are called *molecular compounds*.
- Metal-nonmetal combinations usually lead to ionic bonds and *ionic compounds*. 
Classification of Compounds

**Molecular compound**
Hydrogen chloride, HCl, gas

- HCl molecule
- Nonmetal
- Nonmetal
- Covalent bond

**Ionic compound**
Sodium chloride, NaCl, solid

- Nonmetallic anions
- Metallic cations
Summary

- **Nonmetal-nonmetal** combinations (e.g. HCl)
  - Covalent bonds
  - Molecules
  - Molecular Compound

- **Metal-nonmetal** combinations (e.g. NaCl)
  - Probably ionic bonds
  - Alternating cations and anions in crystal structure
  - Ionic compound
The valence electrons for each atom are the most important electrons in the formation of chemical bonds.

The number of valence electrons for the atoms of each element is equal to the element’s A-group number on the periodic table.

Covalent bonds often form to pair unpaired electrons and give the atoms of the elements other than hydrogen and boron eight valence electrons (an octet of valence electrons).
Valence Electrons and A-Group Numbers

One valence electron

| 1 | H |

Number of valence electrons equals the A-group number

<table>
<thead>
<tr>
<th>A</th>
<th>3A</th>
<th>4A</th>
<th>5A</th>
<th>6A</th>
<th>7A</th>
<th>8A</th>
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<tbody>
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<td>6</td>
<td>7</td>
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<td>9</td>
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<td>As</td>
<td>Se</td>
<td>Br</td>
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<td></td>
<td>Kr</td>
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<td>34</td>
<td>35</td>
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<td></td>
<td></td>
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<td>53</td>
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<tr>
<td>4H</td>
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<tr>
<td>5H</td>
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</tr>
</tbody>
</table>

8A He

18 Ar

36 Kr

54 Xe
Electron-Dot Symbols and Lewis Structures

- **Electron-dot symbols** show valence electrons.

  ![Cl Electron Dot Symbol]

- Nonbonding pairs of valence electrons are called **lone pairs**.
• *Lewis structures* represent molecules using element symbols, lines for bonds, and dots for lone pairs.
**Most Common Bonding Patterns for Nonmetals**

<table>
<thead>
<tr>
<th>Element</th>
<th># Bonds</th>
<th># lone pairs</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>1</td>
<td>0</td>
</tr>
<tr>
<td>C</td>
<td>4</td>
<td>0</td>
</tr>
<tr>
<td>N, P</td>
<td>3</td>
<td>1</td>
</tr>
<tr>
<td>O, S, Se</td>
<td>2</td>
<td>2</td>
</tr>
<tr>
<td>F, Cl, Br, I</td>
<td>1</td>
<td>3</td>
</tr>
</tbody>
</table>
• Chapter 12 describes procedure that allows you to draw Lewis structures for many different molecules.
• Many Lewis structures can be drawn by attempting to give each atom in a molecule its most common bonding pattern.
Lewis Structure for Methane, CH$_4$

- Carbon atoms usually have 4 bonds and no lone pairs.
- Hydrogen atoms have 1 bond and no lone pairs.
Tetrahedral Geometry

The shaded shape is a regular tetrahedron.
Methane, $\text{CH}_4$
Lewis Structure for Ammonia, NH₃

- Nitrogen atoms usually have 3 bonds and 1 lone pair.
- Hydrogen atoms have 1 bond and no lone pairs.
Ammonia,
$\text{NH}_3$

H–N–H

Space-filling model

Ball-and-stick model

Geometric sketch

$107^\circ$
Lewis Structure for Water, $\text{H}_2\text{O}$

- Oxygen atoms usually have 2 bonds and 2 lone pairs.
- Hydrogen atoms have 1 bond and no lone pairs.
Water, $\text{H}_2\text{O}$

Lewis structure | Space-filling model | Ball-and-stick model | Geometric Sketch

$\text{H} - \text{O} - \text{H}$
Water Attractions

Attraction between partial positive charge and partial negative charge
Liquid Water

Attractions exist between hydrogen and oxygen atoms of different water molecules.

Molecules break old attractions and make new ones as they tumble throughout the container.
Binary covalent compound \( A_a B_b \)

Nonmetallic elements

Subscripts (may be absent)
Common Names

- $\text{H}_2\text{O}$, water
- $\text{NH}_3$, ammonia
- $\text{CH}_4$, methane
- $\text{C}_2\text{H}_6$, ethane
- $\text{C}_3\text{H}_8$, propane
Naming Binary Covalent Compounds

• If the subscript for the first element is greater than one, indicate the subscript with a prefix.
  – We do not write mono- on the first name.
  – Leave the "a" off the end of the prefixes that end in "a" and the “o” off of mono- if they are placed in front of an element that begins with a vowel (oxygen or iodine).

• Follow the prefix with the name of the first element in the formula.
Naming Binary Covalent Compounds

- Write a prefix to indicate the subscript for the second element.
- Write the root of the name of the second symbol in the formula.
- Add -ide to the end of the name.
<table>
<thead>
<tr>
<th>Prefixes</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>mon(o)</td>
<td>hex(a)</td>
</tr>
<tr>
<td>di</td>
<td>hept(a)</td>
</tr>
<tr>
<td>tri</td>
<td>oct(a)</td>
</tr>
<tr>
<td>tetr(a)</td>
<td>non(a)</td>
</tr>
<tr>
<td>pent(a)</td>
<td>dec(a)</td>
</tr>
</tbody>
</table>
Roots of Nonmetals

H hydr-  F fluor-
C carb-  Cl chlor-
N nitr-  Br brom-
P phosph- I iod-
O ox- Se selen-
Forms of Binary Covalent Names

• prefix(name of nonmetal) prefix(root of name of nonmetal)ide
  (for example, dinitrogen pentoxide)
• or (name of nonmetal) prefix(root of name of nonmetal)ide
  (for example, carbon dioxide)
• or (name of nonmetal) (root of nonmetal)ide
  (for example, hydrogen fluoride)
Writing Binary Covalent Formulas

• Write the symbols for the elements in the order mentioned in the name.
• Write subscripts indicated by the prefixes. If the first part of the name has no prefix, assume it is mono-.
The Making of an Anion

When a hydrogen atom gains one electron, or when an atom in group 15 gains three electrons, or when an atom in group 16 gains two electrons, or when an atom in group 17 gains one electron, it has the same number of electrons as an atom of the nearest noble gas.

Atomic number equals number of electrons.
The Making of a Cation

When an atom in group 1 loses one electron,

or when an atom in group 2 loses two electrons,

or when an atom in group 3 loses three electrons,

or when an aluminum atom loses three electrons,

it has the same number of electrons as an atom of the nearest noble gas.

Atomic number equals number of electrons.
# Monatomic Ions

<table>
<thead>
<tr>
<th>Period</th>
<th>1A</th>
<th>2A</th>
<th>3A</th>
<th>4A</th>
<th>5A</th>
<th>6A</th>
<th>7A</th>
<th>8A</th>
</tr>
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<tbody>
<tr>
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<td>Li+</td>
<td>Be²⁺</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td>Na⁺</td>
<td>Mg²⁺</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>K⁺</td>
<td>Ca²⁺</td>
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<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>4</td>
<td>Rb⁺</td>
<td>Sr²⁺</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>5</td>
<td>Cs⁺</td>
<td>Ba²⁺</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>6</td>
<td>Fr⁺</td>
<td>Ra²⁺</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

- **Li⁺**: Single electron removed from the outermost shell, resulting in a +1 charge.
- **Be²⁺**: Two electrons removed, resulting in a +2 charge.
- **Na⁺**: Single electron removed, resulting in a +1 charge.
- **Mg²⁺**: Two electrons removed, resulting in a +2 charge.
- **K⁺**: Single electron removed, resulting in a +1 charge.
- **Ca²⁺**: Two electrons removed, resulting in a +2 charge.
- **Rb⁺**: Single electron removed, resulting in a +1 charge.
- **Sr²⁺**: Two electrons removed, resulting in a +2 charge.
- **Cs⁺**: Single electron removed, resulting in a +1 charge.
- **Ba²⁺**: Two electrons removed, resulting in a +2 charge.
- **Fr⁺**: Single electron removed, resulting in a +1 charge.
- **Ra²⁺**: Two electrons removed, resulting in a +2 charge.

- **Sc³⁺**: Three electrons removed, resulting in a +3 charge.
- **Fe²⁺**: Two electrons removed, resulting in a +2 charge.
- **Fe³⁺**: Three electrons removed, resulting in a +3 charge.
- **Cu²⁺**: Two electrons removed, resulting in a +2 charge.
- **Zn²⁺**: Two electrons removed, resulting in a +2 charge.
- **Ag⁺**: Single electron removed, resulting in a +1 charge.
- **Cd²⁺**: Two electrons removed, resulting in a +2 charge.

- **H⁻**: Single electron added to the outermost shell, resulting in a -1 charge.
- **N³⁻**: Three electrons added, resulting in a -3 charge.
- **O²⁻**: Two electrons added, resulting in a -2 charge.
- **F⁻**: Single electron added, resulting in a -1 charge.
- **Cl⁻**: Single electron added, resulting in a -1 charge.
- **Br⁻**: Single electron added, resulting in a -1 charge.
- **I⁻**: Single electron added, resulting in a -1 charge.
Monatomic Ion Names

• Monatomic Cations
  – (name of metal)
    • Groups 1, 2, and 3 metals
    • $\text{Al}^{3+}$, $\text{Zn}^{2+}$, $\text{Cd}^{2+}$, $\text{Ag}^{+}$
  – (name of metal)(Roman numeral)
    • All metallic cations not mentioned above

• Monatomic Anions
  – (root of nonmetal name)ide
Monatomic Anions

Hydride $H^-$
Nitride $N^{3-}$
Phosphide $P^{3-}$
Oxide $O^{2-}$
Sulfide $S^{2-}$
Selenide $Se^{2-}$

Fluoride $F^-$
Chloride $Cl^-$
Bromide $Br^-$
Iodide $I^-$
Sodium Chloride, NaCl, Structure

Each chloride anion is surrounded by 6 cations.
Each sodium cation is surrounded by 6 anions.

Ball-and-stick model

Space-filling model

Salt (sodium chloride)
CsCl and NH₄Cl structure

Cesium chloride

Each Cs⁺ cation is surrounded by 8 Cl⁻ anions.

Each Cl⁻ anion is surrounded by 8 Cs⁺ cations.

Ammonium chloride

Each NH₄⁺ cation is surrounded by 8 Cl⁻ anions.

Each Cl⁻ anion is surrounded by 8 NH₄⁺ cations.
<table>
<thead>
<tr>
<th>Ion</th>
<th>Name</th>
<th>Ion</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>NH$_4^+$</td>
<td>ammonium</td>
<td>NO$_3^-$</td>
<td>nitrate</td>
</tr>
<tr>
<td>OH$^-$</td>
<td>hydroxide</td>
<td>SO$_4^{2-}$</td>
<td>sulfate</td>
</tr>
<tr>
<td>CO$_3^{2-}$</td>
<td>carbonate</td>
<td>C$_2$H$_3$O$_2^-$</td>
<td>acetate</td>
</tr>
<tr>
<td>PO$_4^{3-}$</td>
<td>phosphate</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Polyatomic Ions with Hydrogen

- $\text{HCO}_3^-$ hydrogen carbonate
- $\text{HSO}_4^-$ hydrogen sulfate
- $\text{HS}^-$ hydrogen sulfide
- $\text{HPO}_4^{2-}$ hydrogen phosphate
- $\text{H}_2\text{PO}_4^-$ dihydrogen phosphate
Recognizing Ionic Compounds

- Metal-nonmetal...binary ionic compound
  - Metallic element
  - Nonmetallic element
  - Binary ionic compound \( M_aA_b \)
  - Subscripts (may be absent)

- Metal-polyatomic ion
- Ammonium-nonmetal or ammonium polyatomic ion
Converting Ionic Formulas to Names

- Name
  - (name of cation) (name of anion)
<table>
<thead>
<tr>
<th>Metals with one possible charge (Al, Zn, Cd, and Groups 1, 2, 3)</th>
<th>name of metal</th>
</tr>
</thead>
<tbody>
<tr>
<td>Metals with more than one possible charge (the rest)</td>
<td>name (Roman numeral)</td>
</tr>
<tr>
<td>Polyatomic cations (e.g. ammonium)</td>
<td>name of polyatomic ion</td>
</tr>
</tbody>
</table>
### Anion Names

<table>
<thead>
<tr>
<th>Type of Anion</th>
<th>Naming Convention</th>
</tr>
</thead>
<tbody>
<tr>
<td>Monatomic anion</td>
<td>(root of nonmetal name)ide</td>
</tr>
<tr>
<td>Polyatomic anion</td>
<td>Name of polyatomic ion</td>
</tr>
</tbody>
</table>
Converting Ionic Names to Formulas

• Determine the formula, including charge, for the cation and anion.
• Determine the ratio of the ions that yields zero overall charge.