Unit 5 Notes
Covalent Bonding, Covalent Compounds, and Intermolecular Forces

Chemical Bond—a mutual electrical attraction between the nuclei and valence electrons of different atoms that binds the atoms together. Chemical Bonds create more stable arrangements of matter.

Covalent Bonding—bonding that results from the sharing of electron pairs.

2 Types of Covalent Bonding
- Nonpolar Covalent Bonding (also called a “pure covalent bond”)
- Polar Covalent Bonding

Nonpolar Covalent Bond—a covalent bond in which the bonding electrons are shared equally, resulting in a balanced distribution of electrical charge (value 0-0.4).

Polar Covalent Bond—covalent bond in which the unbonded atoms have an unequal attraction for the shared electrons (value 0.4-1.7).

Calculating Bond Type—subtract electronegativity values

Molecular Compounds
- A molecule is a group of atoms held together by covalent bonds (usually composed only of nonmetals).
- The composition of a compound is given by its chemical formula. A chemical formula indicates the numbers of atoms in a chemical compound using symbols and subscripts.

Diatomic Molecule—molecules containing only 2 atoms.

7 Diatomic Elements—H₂, O₂, F₂, Cl₂, Br₂, I₂, and N₂

Characteristics of the Covalent Bond

Bond energy is the energy required to break a chemical bond and form neutral atoms. Atoms want to combine to reach stable (noble gas) electron configurations.

The Octet Rule

Atoms can fill their outermost s and p sublevels by sharing electrons through covalent bonding. Such bond formation follows the octet rule: Chemical compounds tend to form so that each atom, by gaining, losing, or sharing electrons, has an octet of electrons in its highest energy level.

Exceptions to the Octet Rule
- Exceptions to the octet rule include those for atoms that cannot fit eight electrons, and for those that can fit more than eight electrons.
- Common exceptions to the octet rule:
  - Phosphorous (can often have more than 8 electrons)
  - Boron (only wants 6 electrons)
  - Hydrogen (only wants 2 electrons)
Electron-Dot Notation

- **Electron-dot notation** is an electron-configuration notation in which only the valence electrons of an atom are shown, indicated by dots placed around the element’s symbol.

Lewis Structures

- Electron-dot notation can also be used to represent molecules.

![H:H](image source: http://wps.prenhall.com/wps/media/objects/4678/4790506/images/aabjuzxa.jpg)

The pair of dots between the two symbols represents a shared electron pair (a covalent bond).

- Bond pair—a pair of electrons involved in bonding.

- In addition, each fluorine atom is surrounded by three pairs of electrons that are not shared in bonds.

![F:F](Image Source: http://wps.prenhall.com/wps/media/objects/4678/4790506/images/aabjuzxa.jpg)

- An *unshared pair*, also called a *lone pair or nonbonding pair*, is a pair of electrons that is not involved in bonding.

- The pair of dots representing a shared pair of electrons in a covalent bond is often replaced by a long dash.

**Example:**

![H:H](Image Source: https://upload.wikimedia.org/wikipedia/commons/thumb/c/c3/Wasserstoff.svg/220px-Wasserstoff.svg.png)

![F:F](Image Source: https://classes.lt.unt.edu/Fall_2010/CECS_5030_026/mrp0113/imageswebproject/fluorine%20lewis%20dot%20symbols.jpg)
A **structural formula** indicates the kind, number, and arrangement, and bonds but not the unshared pairs of the atoms in a molecule.

![F—F](https://classes.lt.unt.edu/Fall_2010/CECS_5030_026/mrp0113/imageswebproject/fluorine%20lewis%20dot%20symbols.jpg)

A **single bond** is a covalent bond in which one pair of electrons is shared between atoms.


A **double bond** is a covalent bond in which two pairs of electrons are shared between atoms.

![H=C=C=H](http://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological/section_16/40088d8077ec636f7a9b00d3c5be9715.jpg)

A **triple bond** is a covalent bond in which three pairs of electrons are shared between atoms.

![H—C≡C—H](http://chemwiki.ucdavis.edu/@api/deki/files/1518/ethyne.png?size=bestfit&width=208&height=141&revision=1)

*ethyne*
Rules for Drawing Lewis Structures

1. Determine the total number of valence electrons in the atoms.
2. Arrange atoms to form a skeleton structure. If C is present, it is the central atom. Otherwise the least electronegative atom is central (except for hydrogen which is never central). Connect the atoms by bonds.
3. Add unshared pairs of electrons so that each hydrogen atom has 2 electrons and each other nonmetal has 8 electrons (there are some exceptions).
4. Count electrons to be sure the number used equals the number available.

Comparing Ionic and Covalent Compounds

<table>
<thead>
<tr>
<th>Ionic</th>
<th>Covalent (Molecular)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bonds formed by attraction between oppositely charged ions.</td>
<td>Bonds formed when electrons are shared.</td>
</tr>
<tr>
<td>High melting point—compounds usually solids at room temperature</td>
<td>Low melting points—compounds may be solids, liquids, or gases at room temperature.</td>
</tr>
<tr>
<td>Ionic Compounds—good conductors when dissolved in water</td>
<td>Covalent Compounds—poor conductors when dissolved in water</td>
</tr>
<tr>
<td>Formed between metal and nonmetal ions</td>
<td>Formed between nonmetals</td>
</tr>
</tbody>
</table>

Polyatomic Ions

Polyatomic ion—a group of atoms with a charge.

Drawing Lewis Structures for Polyatomic Ions

- If ion is positive subtract electrons
- If ion is negative add electrons
- Add brackets and charge of the ion outside the brackets

\[
\text{ammonium} - \text{NH}_4^+ \\
\begin{array}{c}
\text{N} - 5 \\
\text{H} - 4 \\
\hline
9 \\
-1 \\
8
\end{array}
\]

Resonance—having more than one possible Lewis structure. Occurs most commonly when the double bond of a structure can be in more than one location.
**Expanded Octet**—occurs in atoms that have an empty d sublevel (3rd row of the periodic table and down). These atoms will sometimes accommodate more than 8 electrons.

**Molecular Geometry and VSEPR Theory**

**Molecular geometry**—the three-dimensional arrangement of a molecule.

**VSEPR**—valence-shell electron-pair repulsion

**VSEPR theory**—theory used to predict the shapes of molecules based on the number of atoms bonded to the central atom and the number of lone pair electrons.
**Na**
m ing —

  Binary —

  Molecular —

  (Covalent) —

  Compounds

Unlike ionic compounds, molecular compounds are composed of individual covalently bonded units, or molecules (usually 2 nonmetals).

The system of naming molecular compounds is based on the use of prefixes.
The first element in the compound has a prefix **only** if there is more than one atom.

The second element **always** has a prefix.

<table>
<thead>
<tr>
<th>Molecular shape</th>
<th>Atoms bonded to central atom</th>
<th>Lone pairs of electrons</th>
<th>Formula example</th>
<th>Lewis structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>Linear</td>
<td>2</td>
<td>0</td>
<td>BeF₂</td>
<td>:(\text{\text{Be}})-(\text{\text{F}});</td>
</tr>
<tr>
<td>Bent</td>
<td>2</td>
<td>1</td>
<td>SnCl₂</td>
<td>(\text{\text{Sn}})-</td>
</tr>
<tr>
<td>Trigonal-planar</td>
<td>3</td>
<td>0</td>
<td>BF₃</td>
<td>(\text{\text{B}})-</td>
</tr>
<tr>
<td>Tetrahedral</td>
<td>4</td>
<td>0</td>
<td>CH₄</td>
<td>H—C—H</td>
</tr>
</tbody>
</table>

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<thead>
<tr>
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<th>Lewis structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>Trigonal-pyramidal</td>
<td>3</td>
<td>1</td>
<td>NH₃</td>
<td>(\text{\text{N}})-(\text{\text{H}});</td>
</tr>
<tr>
<td>Bent</td>
<td>2</td>
<td>2</td>
<td>H₂O</td>
<td>(\text{\text{O}})-(\text{\text{H}});</td>
</tr>
<tr>
<td>Trigonal-bipyramidal</td>
<td>5</td>
<td>0</td>
<td>PCl₅</td>
<td>(\text{\text{P}})-(\text{\text{Cl}});</td>
</tr>
<tr>
<td>Octahedral</td>
<td>6</td>
<td>0</td>
<td>SF₆</td>
<td>(\text{\text{F}})-(\text{\text{F}});</td>
</tr>
</tbody>
</table>

**Naming Binary Molecular (Covalent) Compounds**

Unlike ionic compounds, molecular compounds are composed of individual covalently bonded units, or molecules (usually 2 nonmetals).

The system of naming molecular compounds is based on the use of prefixes.
The first element in the compound has a prefix **only** if there is more than one atom.

The second element **always** has a prefix.
Prefixes for Naming Covalent Compounds

<table>
<thead>
<tr>
<th>#</th>
<th>Prefix</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Mono-</td>
</tr>
<tr>
<td>2</td>
<td>Di-</td>
</tr>
<tr>
<td>3</td>
<td>Tri-</td>
</tr>
<tr>
<td>4</td>
<td>Tetra-</td>
</tr>
<tr>
<td>5</td>
<td>Penta-</td>
</tr>
<tr>
<td>6</td>
<td>Hexa-</td>
</tr>
<tr>
<td>7</td>
<td>Hepta-</td>
</tr>
<tr>
<td>8</td>
<td>Octa-</td>
</tr>
<tr>
<td>9</td>
<td>Nona-</td>
</tr>
<tr>
<td>10</td>
<td>Deca-</td>
</tr>
</tbody>
</table>

Oxidation Numbers

To indicate the distribution of electrons among the atoms in a molecular compound or a polyatomic ion, oxidation numbers are assigned to the atoms.

Rules for Assigning Oxidation Numbers

1. An atom in its elemental form has an oxidation number of zero.
2. Fluorine has an oxidation number of -1 in all of its compounds.
3. Oxygen has an oxidation number of -2 unless it is combined with F, then it is +2.
4. Hydrogen can have an oxidation number of +1 or -1.
5. Oxidation numbers must add to zero in a compound.

- Many elements can have more than one oxidation number.
- These numbers can be used in the same way as charges.

Intermolecular Forces

*Intermolecular forces*—forces of attraction between molecules

- Higher boiling point—stronger molecular forces

Types of Intermolecular Forces

1. Dipole—intermolecular force created by equal but opposite charges separated by a short distance.

   - A dipole is represented by an arrow pointing toward the negative pole and a crossed tail at the positive pole.

Example:

![Dipole Image](http://3.bp.blogspot.com/-4XwAEJYsY04/UEXGEnF51WI/AAAAAAAAABt8/bhDwaVbMvx0/s1600/HCl+polarity.png)
Dipole-Dipole Forces—forces of attraction between polar molecules.


3. **Hydrogen bonding**—intermolecular force in which a hydrogen atom is bonded to a highly electronegative atom and attracted to an unshared pair of electrons in a nearby molecule.

![Hydrogen Bonding](http://www.physicsofmatter.com/NotTheBook/Talks/Ice/Image7.gif)

4. **London dispersion forces (also known as Van der Waals forces)**—intermolecular attractions resulting from the constant motion of electrons.

![London Dispersion Forces](https://i.ytimg.com/vi/P2CF2JBZAY8/maxresdefault.jpg)

**Cohesion vs. Adhesion**

Cohesion is the property of molecules of the same type sticking together, being mutually attractive. This property is caused by the shape and structure of its molecules creating electrical attraction.
Adhesion is the tendency of different types of particles to cling to one another.

Both cohesion and adhesion are the result of intermolecular attractions.