<table>
<thead>
<tr>
<th>TYPE</th>
<th>One unit is called:</th>
<th>Made of:</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Ionic Bonds</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Formula Unit</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Cation (+ Charge)</td>
<td></td>
<td>Metal Atoms</td>
</tr>
<tr>
<td>Anion (-charge)</td>
<td></td>
<td>Non-Metal Atoms</td>
</tr>
<tr>
<td><strong>Metallic Bonds</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Formula Unit</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Metal Atoms</td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>Covalent Bonds</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Molecule</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Non-Metal Atoms</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Chapter 7: Ionic Bonds

Chapter 8: Covalent Bonds
Covalent Bonds

8.1 Molecular Compounds

8.2 The Nature of Covalent Bonds

8.4 Polar Bonds and Molecules
Covalent Bond: atoms held together by sharing electrons

Molecule: neutral group of atoms joined by covalent bonds

Water ($\text{H}_2\text{O}$)  Carbon monoxide ($\text{CO}$)
- Diatomic molecule: molecule with two atoms of one element

Nitrogen: $\text{N}_2$

Hydrogen: $\text{H}_2$
• Molecular compound:
  atoms bonded covalently (sharing valence e-)

• lower melting and boiling points (than ionic compounds)
• gases or liquids at room temperature
• composed of two or more nonmetals

Water (H₂O)  Carbon monoxide (CO)  Nitrogen: N₂
How are molecular compounds different from ionic compounds?

**Ionic compounds:**
- Ions: lose/gain e-
- Higher melting points
- Conduct electricity when melted or dissolved in water

**Molecular compounds:**
- Covalent bonds: share e-
- Lower melting points
- Lower boiling points

---

![Diagram of ionic compound (Table Salt) and molecular compound (Water)](image)
Molecular formula

- Chemical formula of a molecular compound
  - Tells how many atoms of each element in a molecule
- Structural formula: shows the arrangement of atoms
Which molecule has the greatest amount of oxygen atoms?

Water ($H_2O$)  Carbon dioxide ($CO_2$)  Ethanol ($C_2H_6O$)
Octet rule in covalent bonding

- e- shared so each atom has a full valence shell (Noble Gas configuration—usually an octet!)

\[ \text{H}^- + \text{H}^+ \rightarrow \text{H}_2 \]

H:H the two dots represents shared pair of electrons
Single covalent bond: two atoms held together by sharing a pair of e-

H : H becomes H—H

Bond Line
\[
\begin{align*}
\cdot \hat{F} & \quad + \quad \cdot \hat{F} \\
\text{Fluorine atom} & \quad \text{Fluorine atom} & \quad \text{Fluorine molecule} \\
\end{align*}
\]
2H· + :O· → :O:H \text{ or } :O—H

Hydrogen atoms
Oxygen atom

Water molecule

Water molecule
Ammonia $\text{NH}_3$

Unshared Pair: a pair of valence electrons (lone pair) that is not shared between atoms

H• H• H•

Nitrogen atom

Ammonia molecule

N

$1s$

$2s$

$2p$

$\uparrow \downarrow$

$\uparrow \downarrow$

$1s$

$1s$

$1s$

H

H

H

Ammonia molecule
Methane \( \text{CH}_4 \)

- Draw methane structural formula \( \text{CH}_4 \)
Hydrogen Chloride is a molecular compound with a single covalent bond.

Draw the electron dot structure for HCl

\[ \text{H} \cdot + \cdot \text{Cl} : \rightarrow \text{H} : \text{Cl} : \]

Hydrogen chloride molecule

\[ \text{H-Cl} \]
Molecular Model Activity

- Work in lab groups
- Write in lab book as you go.
- Activities do not require a complete lab write-up.
Chapter 8 notes continued
Double and triple bonds

Double covalent bond:
shares two pairs of electrons

Triple covalent bond:
shares three pairs of electrons
<table>
<thead>
<tr>
<th>Name</th>
<th>Chemical formula</th>
<th>Electron dot structure</th>
<th>Properties and uses</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fluorine</td>
<td>F₂</td>
<td>··F—F·</td>
<td>Greenish-yellow reactive toxic gas. Compounds of fluorine, a halogen, are added to drinking water and toothpaste to promote healthy teeth.</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Cl₂</td>
<td>··Cl—Cl·</td>
<td>Greenish-yellow reactive toxic gas. Chlorine is a halogen used in household bleaching agents.</td>
</tr>
<tr>
<td>Bromine</td>
<td>Br₂</td>
<td>··Br—Br·</td>
<td>Dense red-brown liquid with pungent odor. Compounds of bromine, a halogen, are used in the preparation of photographic emulsions.</td>
</tr>
<tr>
<td>Iodine</td>
<td>I₂</td>
<td>··I—I·</td>
<td>Dense gray-black solid that produces purple vapors; a halogen. A solution of iodine in alcohol (tincture of iodine) is used as an antiseptic.</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>H₂</td>
<td>H—H</td>
<td>Colorless, odorless, tasteless gas. Hydrogen is the lightest known element.</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>N₂</td>
<td>··N≡N·</td>
<td>Colorless, odorless, tasteless gas. Air is almost 80% nitrogen by volume.</td>
</tr>
<tr>
<td>Oxygen</td>
<td>O₂</td>
<td>··O=O·</td>
<td>Colorless, odorless, tasteless gas that is vital for life. Air is about 20% oxygen by volume.</td>
</tr>
</tbody>
</table>
Coordinate Covalent bond is a covalent bond in which one atom contributes both bonding electrons.

Oxygen is stable but carbon still is not so oxygen must donate one of its unshared electrons.

Carbon monoxide

\[
\text{C} + \text{O} \rightarrow \text{C} \equiv \text{O}
\]

Carbon dioxide

\[
\text{O} + \text{C} + \text{O} \rightarrow \text{O} \equiv \text{C} \equiv \text{O}
\]
Polyatomic ion

Ammonium ion $\text{NH}_4^+$

Tetrahedral E. P. G.
Tetrahedral Molecular Geometry
Polyatomic ion: combination of covalent and ionic bonds

Hydronium ion

\[
H^+ + \overset{\cdot}{H}O\overset{\cdot}{H} \rightarrow \left[H\overset{\cdot}{O}\overset{\cdot}{H}\right]^+ \quad \text{or} \quad \left[H\overset{\rightarrow}{O}\overset{\leftarrow}{H}\right]^+
\]

Hydrogen ion (proton) \quad \text{Water molecule (H}_2\text{O)} \quad \text{Hydronium ion (H}_3\text{O}^+)
<table>
<thead>
<tr>
<th>Name</th>
<th>Chemical formula</th>
<th>Structure</th>
<th>Properties and uses</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen peroxide</td>
<td>$H_2O_2$</td>
<td><img src="#" alt="Structure" /></td>
<td>Colorless, unstable liquid when pure. It is used as rocket fuel. A 3% solution is used as a bleach and antiseptic.</td>
</tr>
<tr>
<td>Sulfur dioxide</td>
<td>$SO_2$</td>
<td><img src="#" alt="Structure" /></td>
<td>Oxides of sulfur are produced in combustion of petroleum products and coal. They are major air pollutants in industrial areas. Oxides of sulfur can lead to respiratory problems.</td>
</tr>
<tr>
<td>Sulfur trioxide</td>
<td>$SO_3$</td>
<td><img src="#" alt="Structure" /></td>
<td>Oxides of nitrogen are major air pollutants produced by the combustion of fossil fuels in automobile engines. They irritate the eyes, throat, and lungs.</td>
</tr>
<tr>
<td>Nitric oxide</td>
<td>NO</td>
<td><img src="#" alt="Structure" /></td>
<td>Colorless, sweet-smelling gas. It is used as an anesthetic commonly called laughing gas.</td>
</tr>
<tr>
<td>Nitrogen dioxide</td>
<td>$NO_2$</td>
<td><img src="#" alt="Structure" /></td>
<td>Colorless, toxic gas with the smell of almonds.</td>
</tr>
<tr>
<td>Nitrous oxide</td>
<td>$N_2O$</td>
<td><img src="#" alt="Structure" /></td>
<td>Two hydrogen halides, all extremely soluble in water. Hydrogen chloride, a colorless gas with pungent odor, readily dissolves in water to give a solution called hydrochloric acid.</td>
</tr>
</tbody>
</table>
Bond Dissociation Energy

- Energy need to break the bond between two covalently bonded atoms

  - Single carbon-carbon bond about 347 kJ/mol
  - Double carbon-carbon bond about 657 kJ/mol
  - Triple carbon-carbon bond about 908 kJ/mol

- Compounds with C-C and C-H single covalent bonds are unreactive because dissociation energy is high
Resonance Structures:
* 2 ways to draw a dot structure
* use $\leftrightarrow$ and show both structures

Ozone

\[ \begin{array}{c}
\begin{array}{c}
:0\cdot0\cdot0
\end{array}
\end{array} \leftrightarrow
\begin{array}{c}
\begin{array}{c}
0\cdot0\cdot0
\end{array}
\end{array} \]
Exceptions to the Octet Rule

- The octet rule cannot be satisfied in molecules whose total number of valence electrons is an **odd number**.
- There are also molecules in which an atom has **more or less** than a complete octet of valence electron.

Nitrogen dioxide molecule

\[ \text{O} = \text{N} = \text{O} \]

\[ \text{O} - \text{N} = \text{O} \]
Compare/Contrast:

Ionic Bonds

Covalent Bonds
How does electronegativity determine charge distribution in a polar bond?

When the atoms in a bond pull equally (when identical atoms are bonded) the bonding electrons are shared equally and the bond is a **nonpolar covalent bond**.
Polarity: unequal pulling of an e- toward the atom

The chlorine atom attracts the electron cloud more than the hydrogen atom does.
- A **polar covalent bond**, known also as a polar bond, is a covalent bond between atoms in which the electrons are shared unequally.

- The more electronegative atom attracts electrons more strongly and gains a slightly negative charge.

- The less electronegative atom has a slightly positive charge.
In a **polar molecule**, one end of the molecule is slightly negative and the other end is slightly positive.

**Dipole**: a molecule that has two poles (dipolar molecule)
### Periodic Table of the Elements

**Electronegativity**

<table>
<thead>
<tr>
<th>Element</th>
<th>Electronegativity</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>2.20</td>
</tr>
<tr>
<td>Li</td>
<td>0.98</td>
</tr>
<tr>
<td>Na</td>
<td>0.93</td>
</tr>
<tr>
<td>K</td>
<td>0.82</td>
</tr>
<tr>
<td>Rb</td>
<td>0.82</td>
</tr>
<tr>
<td>Cs</td>
<td>0.79</td>
</tr>
<tr>
<td>Fr</td>
<td>0.7</td>
</tr>
</tbody>
</table>

**Lanthanides**

<table>
<thead>
<tr>
<th>Element</th>
<th>Electronegativity</th>
</tr>
</thead>
<tbody>
<tr>
<td>La</td>
<td>1.10</td>
</tr>
<tr>
<td>Ce</td>
<td>1.12</td>
</tr>
<tr>
<td>Pr</td>
<td>1.13</td>
</tr>
<tr>
<td>Nd</td>
<td>1.14</td>
</tr>
<tr>
<td>Pm</td>
<td>1.13</td>
</tr>
<tr>
<td>Sm</td>
<td>1.17</td>
</tr>
<tr>
<td>Eu</td>
<td>1.2</td>
</tr>
<tr>
<td>Gd</td>
<td>1.2</td>
</tr>
<tr>
<td>Tb</td>
<td>1.2</td>
</tr>
<tr>
<td>Dy</td>
<td>1.22</td>
</tr>
<tr>
<td>Ho</td>
<td>1.23</td>
</tr>
<tr>
<td>Er</td>
<td>1.24</td>
</tr>
<tr>
<td>Tm</td>
<td>1.25</td>
</tr>
<tr>
<td>Yb</td>
<td>1.31</td>
</tr>
</tbody>
</table>

**Actinides**

<table>
<thead>
<tr>
<th>Element</th>
<th>Electronegativity</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ac</td>
<td>1.1</td>
</tr>
<tr>
<td>Th</td>
<td>1.3</td>
</tr>
<tr>
<td>Pa</td>
<td>1.5</td>
</tr>
<tr>
<td>U</td>
<td>1.38</td>
</tr>
<tr>
<td>Np</td>
<td>1.36</td>
</tr>
<tr>
<td>Pu</td>
<td>1.28</td>
</tr>
<tr>
<td>Am</td>
<td>1.3</td>
</tr>
<tr>
<td>Cm</td>
<td>1.3</td>
</tr>
<tr>
<td>Bk</td>
<td>1.3</td>
</tr>
<tr>
<td>Cf</td>
<td>1.3</td>
</tr>
<tr>
<td>Es</td>
<td>1.3</td>
</tr>
</tbody>
</table>

***Elements > 104 exist only for very short half-lifes and the data is unknown.***

[Source: chemistry.about.com](http://chemistry.about.com)

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About Chemistry
A. Definition of IMF

InterMolecular Forces: attractive forces between molecules.

Intermolecular attractions are much weaker than chemical bonds within molecules.

◆ a.k.a. van der Waals forces
IMF and Molecular Properties:

These attractions are responsible for determining whether a molecular compound is a gas, a liquid, or a solid at a given temperature.

State of matter: solid, liquid, gas

IE: Melting a solid requires energy to overcome IMF
Types of IMF: Dispersion

Temporary attraction between regions of MOVING polar molecules.

London forces arise from transitory dipoles.

Senese, senese@antoine.frostburg.edu © 1999
Types of IMF: Dipole-Dipole

Attraction between oppositely charged regions of polar molecules

View animation online.

F. Senese, senese@antoine.frostburg.edu © 1999
Types of IMF: Hydrogen Bonding

Hydrogen attracted to:
- Oxygen
- Nitrogen
- Fluorine
Frayer Model: Ionic Bonds

Covalent Bonds
### B. Types of IMF

<table>
<thead>
<tr>
<th></th>
<th>London Dispersion Forces</th>
<th>Dipole-Dipole Forces</th>
<th>Hydrogen Bonding</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Diagram</strong></td>
<td><img src="image1" alt="Diagram of London Dispersion Forces" /></td>
<td><img src="image2" alt="Diagram of Dipole-Dipole Forces" /></td>
<td><img src="image3" alt="Diagram of Hydrogen Bonding" /></td>
</tr>
<tr>
<td><strong>Relative Strength</strong></td>
<td>weakest</td>
<td>medium strength</td>
<td>strongest</td>
</tr>
<tr>
<td><strong>Other</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
C. Determining IMF

\( \text{NCl}_3 \)
- polar = dispersion, dipole-dipole

\( \text{CH}_4 \)
- nonpolar = dispersion

\( \text{HF} \)
- H-F bond = dispersion, dipole-dipole, hydrogen bonding
POLAR BONDS AND MOLECULES

SUMMARIZE:
- IMF
- VAN DER WAALS FORCES
- DISPERSION FORCES
- DIPOLE INTERACTIONS
- HYDROGEN BONDS