Unit 4: Chemical Bonding & Molecules (Chapter 6 in book)

“Perhaps one of you gentlemen would mind telling me just what it is outside the window that you find so attractive...?”

Cartoon courtesy of NearingZero.net
Chemical Bonding

pgs. 161–182
Chemical Bonds

- Attraction between the *nuclei* and *valence electrons* of different atoms that “glues” the atoms together.
  - the difference between materials as diverse as diamonds and pencils is how they're glued together.

- Why?
  - Bonded atoms are more stable than solo atoms

- How?
  - Atoms will share or exchange valence electrons to achieve a full outer shell (usually octet).
3 Main Types of Bonds

**Ionic Bonds** - Transfer of electrons between atoms
- electrical attraction between cations & anions
- *Formed by: metals & non-metals*

**Covalent Bonds** - sharing of electrons between atoms
- “co” = sharing, “valent” = outer electrons
- *Formed by: non-metals & non-metals*

**Metallic Bonds** - Metal atoms that share a “sea of electrons”
- *Formed by: metals & metals*
Predicting Bond Types

• Bonding is not usually purely ionic or covalent, but somewhere in between

• The difference in electronegativity strength of the atoms in a bond can help us estimate what percentage of the bond will be ionic
  (see example on next slide)
Using the Periodic Table to Determine Bond Types

### Electronegativity

<table>
<thead>
<tr>
<th>Element</th>
<th>Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>2.1</td>
</tr>
<tr>
<td>Li</td>
<td>1.0</td>
</tr>
<tr>
<td>Na</td>
<td>0.9</td>
</tr>
<tr>
<td>K</td>
<td>0.8</td>
</tr>
<tr>
<td>Rb</td>
<td>0.8</td>
</tr>
<tr>
<td>Cs</td>
<td>0.8</td>
</tr>
<tr>
<td>Fr</td>
<td>0.7</td>
</tr>
<tr>
<td>Be</td>
<td>1.5</td>
</tr>
<tr>
<td>Mg</td>
<td>1.2</td>
</tr>
<tr>
<td>Ca</td>
<td>1.0</td>
</tr>
<tr>
<td>Sc</td>
<td>1.3</td>
</tr>
<tr>
<td>Ti</td>
<td>1.5</td>
</tr>
<tr>
<td>V</td>
<td>1.6</td>
</tr>
<tr>
<td>Cr</td>
<td>1.6</td>
</tr>
<tr>
<td>Mn</td>
<td>1.5</td>
</tr>
<tr>
<td>Fe</td>
<td>1.8</td>
</tr>
<tr>
<td>Co</td>
<td>1.8</td>
</tr>
<tr>
<td>Ni</td>
<td>1.8</td>
</tr>
<tr>
<td>Cu</td>
<td>1.9</td>
</tr>
<tr>
<td>Zn</td>
<td>1.6</td>
</tr>
<tr>
<td>Ga</td>
<td>1.6</td>
</tr>
<tr>
<td>Ge</td>
<td>1.8</td>
</tr>
<tr>
<td>As</td>
<td>2.0</td>
</tr>
<tr>
<td>Se</td>
<td>2.4</td>
</tr>
<tr>
<td>Br</td>
<td>2.8</td>
</tr>
<tr>
<td>B</td>
<td>2.0</td>
</tr>
<tr>
<td>C</td>
<td>2.5</td>
</tr>
<tr>
<td>N</td>
<td>3.0</td>
</tr>
<tr>
<td>O</td>
<td>3.5</td>
</tr>
<tr>
<td>F</td>
<td>4.0</td>
</tr>
<tr>
<td>Al</td>
<td>1.5</td>
</tr>
<tr>
<td>Si</td>
<td>1.8</td>
</tr>
<tr>
<td>P</td>
<td>2.1</td>
</tr>
<tr>
<td>S</td>
<td>2.5</td>
</tr>
<tr>
<td>Cl</td>
<td>3.0</td>
</tr>
</tbody>
</table>

#### Ionic bond =
- metal (weak) & non-metal (strong)
- huge difference in strength (1.7 or more)

#### Covalent bond =
- 2 non-metals (strong)
- close to same strength

#### Metallic Bond =
- 2 Metals (both weak)
**Summary: Ionic Bonds vs. Covalent Bonds**

<table>
<thead>
<tr>
<th>Ionic</th>
<th>Covalent</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Electrons are</strong></td>
<td><strong>Electrons are</strong></td>
</tr>
<tr>
<td><strong>Transferred</strong> (become charged ions that are attracted)**</td>
<td><strong>shared</strong></td>
</tr>
<tr>
<td><strong>Metal + non-metal</strong> (ex: Li + K)</td>
<td><strong>2 non-metals</strong></td>
</tr>
<tr>
<td></td>
<td><em>(O + O or O + N)</em></td>
</tr>
<tr>
<td><strong>One atom is a lot higher electronegativity than the other (1.7)</strong></td>
<td><strong>Close to equal electronegativities</strong></td>
</tr>
<tr>
<td></td>
<td><em>(less than 1.7)</em></td>
</tr>
</tbody>
</table>
Lewis Dot Structures
Octet Rule

- Most* atom wants to have 8 electrons in their valence shell (outermost shell)

- Chemical compounds form so that each atom can complete their octet by gaining, losing or sharing electrons

- *Exceptions =
  - H & He (they only want 2 electrons in their valence shell)
  - B (forms bonds so it will have only 6 electrons)
  - F, O & Cl (will sometimes be surrounded by more than 8 electrons because they are so electronegative)
Lewis Dot Structure

• Picture showing how many valence electrons an atom has (dots).

• Helps determine how atoms will bond.

Ex: Phosphorus (has 5 valence electrons)
Lewis Dot Structures for Ionic Compounds

• A way to show how atoms achieve the octet with each other.

• Note:
  – the transfer of the electron
  – the charges ions that result

[Na]⁺  [·Cl⁻]⁻

This is how we draw it
Lewis Dot Structures for Covalent Molecules

2 ways to show:

- With electrons being shared in between

- Line showing the sharing of pair of electrons
3 Bonds Types in More Depth
Covalent Bonds

• Result from the sharing of electron pairs between two atoms

• *Molecule* = termed used to describe atoms are held together by covalent bonds
Covalent Bonds

- Occurs between 2 non-metals
- Electrons are shared

- 2 types of covalent bonds: Polar and non-polar (to be discussed later)

  - Ex: Water & most biological molecules (sugars, fats, proteins)

- Can form single, double, or triple bonds
Ionic Bonds

- Forms between: Metal + Non-Metal
- Electrons are transferred
Ionic Bonds (cont.)

Ex: \[ \text{K} \cdot \overset{\cdot}{\overset{\cdot}{\text{Cl}}} \rightarrow \text{K}^+ [\overset{\cdot}{\overset{\cdot}{\text{Cl}}}^-] \]

- Electroneg= .8
- Electroneg= 3.0

- **Cl** is so much stronger that it will "take" K’s electron

- The transfer of electron causes K to be a **cation** (+) and Cl to be an **anion** (-).

- Oppositely charged particles are highly attracted to each other... Ionic bond!
Characteristics of Ionic Compounds

- Shape: crystal lattice of alternating positive and negative ions
- Ex: NaCl and salts
- Ionic bonds are strong so they are:
  - hard
  - have a high melting point
  - high boiling point
Metallic Bonds- “sea of electrons”

- Forms between 2 metals
- Metal atoms valence electrons overlap creating a “sea of electrons”.
  - Electrons do not belong to any one atom, but roam freely throughout the metal atoms
- Ex: Brass (alloy of Cu + Zn)
Metallic Characteristics

- Because of these roaming “sea” of electrons:
  - metals are great conductors of heat/electricity
  - they are ductile (can be made into wire)
  - they are malleable (can be hammered into sheets)
Electrical Conductivity

A. Distilled water does not conduct a current

B. Positive and negative ions fixed in a solid do not conduct a current

C. In solution, positive and negative ions move and conduct a current
Properties & Bonding Type

pgs. 161-182
<table>
<thead>
<tr>
<th>Formation</th>
<th>Covalent</th>
<th>Ionic Bonds</th>
<th>Metallic Bonds</th>
</tr>
</thead>
<tbody>
<tr>
<td>Types of Atoms</td>
<td>Non-metal &amp; non-metal</td>
<td>Non-metal &amp; metal</td>
<td>Metal &amp; metal</td>
</tr>
<tr>
<td>Electron Distribution</td>
<td>Shared</td>
<td>Transferred</td>
<td>Sea of electrons</td>
</tr>
<tr>
<td>Characteristics</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Bond Strength</td>
<td>Strong</td>
<td>Very strong</td>
<td>Varies</td>
</tr>
<tr>
<td>Structure</td>
<td>Neutral group</td>
<td>Crystal lattice</td>
<td>crystalline</td>
</tr>
<tr>
<td>Properties of Compounds</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Type of Compound</td>
<td>Molecular</td>
<td>Ionic</td>
<td>metallic</td>
</tr>
<tr>
<td>Melting Point</td>
<td>Low</td>
<td>Very high</td>
<td>n/a</td>
</tr>
<tr>
<td>Boiling Point</td>
<td>Low</td>
<td>High</td>
<td>Very high</td>
</tr>
<tr>
<td>Malleability</td>
<td>n/a</td>
<td>Not malleable, brittle</td>
<td>Very malleable</td>
</tr>
<tr>
<td>Ductility</td>
<td>n/a</td>
<td>Not ductile</td>
<td>Very ductile</td>
</tr>
<tr>
<td>Conductivity</td>
<td>Not conductive</td>
<td>Conductive</td>
<td>Highly conductive</td>
</tr>
</tbody>
</table>
Bond Energy & Bond Length
**Bond Energy**

- **energy required to break bond**

<table>
<thead>
<tr>
<th>Bond</th>
<th>Bond Length</th>
<th>Bond Energy</th>
</tr>
</thead>
<tbody>
<tr>
<td>Single Bond</td>
<td></td>
<td>Low</td>
</tr>
<tr>
<td>Double Bond</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Triple Bond</td>
<td></td>
<td>High</td>
</tr>
</tbody>
</table>

Only one pair of electrons holding the nuclei together

Two pair of electrons hold the nuclei tighter and closer
## Bond Energy & Bond Lengths

<table>
<thead>
<tr>
<th>Bond</th>
<th>Length (picometers)</th>
<th>Energy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H—Br</td>
<td>141</td>
<td>366</td>
</tr>
<tr>
<td>H—C</td>
<td>109</td>
<td>413</td>
</tr>
<tr>
<td>H—N</td>
<td>101</td>
<td>391</td>
</tr>
<tr>
<td>H—O</td>
<td>96</td>
<td>464</td>
</tr>
<tr>
<td>H—S</td>
<td>93</td>
<td>339</td>
</tr>
<tr>
<td>C—O</td>
<td>143</td>
<td>360</td>
</tr>
<tr>
<td>C=O</td>
<td>129</td>
<td>799</td>
</tr>
<tr>
<td>C—C</td>
<td>154</td>
<td>348</td>
</tr>
<tr>
<td>C=C</td>
<td>134</td>
<td>614</td>
</tr>
<tr>
<td>C C</td>
<td>120</td>
<td>839</td>
</tr>
<tr>
<td>O—O</td>
<td>148</td>
<td>145</td>
</tr>
<tr>
<td>O=O</td>
<td>121</td>
<td>498</td>
</tr>
<tr>
<td>N—N</td>
<td>145</td>
<td>170</td>
</tr>
<tr>
<td>N=N</td>
<td>125</td>
<td>418</td>
</tr>
<tr>
<td>N N N</td>
<td>110</td>
<td>945</td>
</tr>
</tbody>
</table>
Lewis Structures in Covalently Bonded Molecules & HONC Rule

pgs. 183 - 186
Drawing Lewis Dot Structures for Molecules

- **arrange atoms to form a skeleton**
  - Carbon is center atom
  - Hydrogen is never a central atom

- **Pair up all electrons**
  - Unpaired electrons can pair unpaired electron from another atom to form a bond

- **Make sure each atom of the molecule obeys the octet rule & HONC rule**
- **Make sure you have correct # of valence electrons**
Examples of Lewis Dot Structure

\[
\text{CH}_4 \\
\text{H} \quad - \quad \text{C} \quad - \quad \text{H} \\
\text{H} \\

\text{H}_2\text{O} \\
\text{H} \quad - \quad \overset{\cdot}{\text{O}} \quad - \quad \text{H} \\

\text{NH}_3 \\
\text{H} \quad - \quad \overset{\cdot}\overset{\cdot}{\text{N}} \quad - \quad \text{H} \\
\text{H} \\

\text{I}_2 \\
\overset{\cdot}{\text{I}} \quad - \quad \overset{\cdot}{\text{I}}
\]
Multiple Covalent Bonds: Double bonds

Two pairs of shared electrons

$O_2$:

$\text{O} + \text{O} \rightarrow \text{O}=\text{O}$ or $\text{O} \equiv \text{O}$

each oxygen has 8 electrons in the valence shell

$CO_2$:

$\ddot{\text{O}}=\text{C}=\ddot{\text{O}}$
Multiple Covalent Bonds: Triple bonds

Three pairs of shared electrons
Molecular vs. Structural Formulas

- **Molecular formulas** - show how many atoms of each element are in the molecules
  - Ex: $C_6H_{12}O_6 = 6$ carbons, 12 hydrogens & 6 oxygens

- **Structural formulas** - show the 2-dimensional shape of the molecule
  - Ex:
HONC 1-2-3-4 Rule

- Hydrogen, oxygen, nitrogen & carbon are common elements found in biological molecules.
  - Hydrogen needs 1 electron to fill its “octet”
  - Oxygen needs 2 electrons to fill its octet
  - Nitrogen needs 3 electrons to fill its octet
  - Carbon needs 4 electrons to fill its octet
- “1-2-3-4” can be used to predict how these atoms will form bonds with other atoms to build molecules.
Molecular Geometry
Seeing Molecules in 3-D
Molecular Geometry
molecules are really 3-D!

\[ \text{CH}_4 \text{ in 2-D on a sheet of paper} \]

\[ \text{CH}_4 \text{ looks like this in 3-D} \]
Valence Electrons determine Molecular “VSEPR” Shape

• VSEPR = “Valence-Shell Electron Pair Repulsion”

• Electron pairs (bonding or lone pairs) in a molecule repel each other and will try and get as far away from each other as possible... this determines the shape.

\[
\text{NH}_3 \text{ in 2-D}
\]

\[
\text{NH}_3 \text{ VSEPR shape in 3-D}
\]
4 Shapes to Know

Tetrahedral

Pyramidal

Bent

Linear
How Lone Pairs Affect Molecular Shape

“paddles” are lone pairs of electrons.

Remove the paddles and you can see the shapes.

tetrahedral  pyramidal  bent  linear
Steps for Determining Molecular Geometry

1. Draw Lewis dot structure
2. Count number atoms bonded to the central atom
3. Count number of lone-pair electrons on the central atom
4. Look up the Geometry on the chart
Shapes in Large Molecules

Large molecules are composed of the small shapes we’ve studied.

Ex: tetrahedral

ball-and-stick model of citronellol
Why Shape Matters

Ethyl Acetate ($C_4H_8O_2$)  Butyric Acid ($C_4H_8O_2$)

Same formula, but different shapes = very different smells

Rum extract smell  Rancid butter smell
Polarity
Differences In Electronegativities

- Ionic: 100% (3.3)
- Polar-Covalent: 50% (1.7)
- Nonpolar-Covalent: 5% (0.3)
- Nonpolar: 0% (0)
# Practice Problems

<table>
<thead>
<tr>
<th>Bonding Between:</th>
<th>Difference in Electronegativity</th>
<th>Bond Type</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl &amp; Ca</td>
<td>$3.0 - 1.0 = 2.0$</td>
<td>Ionic</td>
</tr>
<tr>
<td>O &amp; H</td>
<td>$3.5 - 2.1 = 1.4$</td>
<td>Polar-covalent</td>
</tr>
<tr>
<td>B &amp; H</td>
<td>$</td>
<td>2.0 - 2.1</td>
</tr>
</tbody>
</table>
2 types of Covalent Bonds:

**Non-Polar**
- Electrons are shared equally
- Usually the same element bonded to itself

**Polar**
- Unequal sharing of electrons between atoms
- More electronegative atoms “hogs” electrons

(Arrow shows F is “pulling” electrons)

“partial positive charge”

“partial negative charge”
Visual Comparison of Bond Types

- Nonpolar covalent bond
- Polar covalent bond
- Ionic bond
Determining Polarity

1. Draw correct VSEPR Shape
2. Determine if molecule is symmetrical.
3. If the molecule is symmetrical = **non-polar**
   - no partial charges are needed!
4. If the molecule is **NOT** symmetrical = **polar**
   - you must show partial charges.
   - always bent or pyramidal shapes

![Linear shape (symmetrical)](image1)

![Bent shape (asymmetrical)](image2)
Ex: $CO_2$

- Carbon dioxide = nonpolar
- has polar bonds, but they cancel each other out.
EX: Water= Polar Molecule

How we know:

1) Cut the molecule on 2 planes
   - see how it’s different above the horizontal line = non symmetrical

1) One atom is “pulling”, look at periodic table to determine which one.

\[ \delta^- \quad \delta^+ \]

Indicates which atom “pulls” the electrons

\[ \delta^- \]

Means oxygen is slightly negative because it “hogs” electrons

\[ H \quad H \]
2 views of Polar water

(a) \[ \delta^+\text{H} - \delta^+\text{O} \]

(b) \[ \delta^+\text{H} \quad \delta^-\text{O} \]

Non Polar Molecules

- **Non Polar molecule** = “no pull”
  - equal sharing of electrons
  - No difference in electronegativity
  - symmetrical in shape

---

Nonpolar covalent bonds

\( \text{Cl - Cl} \)
Examples of Polar & Nonpolar Molecules
## Electronegativity

<table>
<thead>
<tr>
<th>Period</th>
<th>Element</th>
<th>Electronegativity</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>H</td>
<td>2.1</td>
</tr>
<tr>
<td></td>
<td>Li</td>
<td>1.0</td>
</tr>
<tr>
<td></td>
<td>Be</td>
<td>1.6</td>
</tr>
<tr>
<td></td>
<td>Na</td>
<td>0.9</td>
</tr>
<tr>
<td></td>
<td>Mg</td>
<td>1.3</td>
</tr>
<tr>
<td></td>
<td>K</td>
<td>0.8</td>
</tr>
<tr>
<td></td>
<td>Ca</td>
<td>1.3</td>
</tr>
<tr>
<td></td>
<td>Sc</td>
<td>1.4</td>
</tr>
<tr>
<td></td>
<td>Ti</td>
<td>1.5</td>
</tr>
<tr>
<td></td>
<td>V</td>
<td>1.6</td>
</tr>
<tr>
<td></td>
<td>Cr</td>
<td>1.7</td>
</tr>
<tr>
<td></td>
<td>Mn</td>
<td>1.6</td>
</tr>
<tr>
<td></td>
<td>Fe</td>
<td>1.8</td>
</tr>
<tr>
<td></td>
<td>Co</td>
<td>1.9</td>
</tr>
<tr>
<td></td>
<td>Ni</td>
<td>1.9</td>
</tr>
<tr>
<td></td>
<td>Cu</td>
<td>1.9</td>
</tr>
<tr>
<td></td>
<td>Zn</td>
<td>1.7</td>
</tr>
<tr>
<td></td>
<td>Ga</td>
<td>1.6</td>
</tr>
<tr>
<td></td>
<td>Ge</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>As</td>
<td>2.2</td>
</tr>
<tr>
<td></td>
<td>Se</td>
<td>2.5</td>
</tr>
<tr>
<td></td>
<td>Br</td>
<td>3.0</td>
</tr>
<tr>
<td></td>
<td>Kr</td>
<td>3.0</td>
</tr>
<tr>
<td></td>
<td>Rb</td>
<td>0.8</td>
</tr>
<tr>
<td></td>
<td>Sr</td>
<td>1.0</td>
</tr>
<tr>
<td></td>
<td>Y</td>
<td>1.2</td>
</tr>
<tr>
<td></td>
<td>Zr</td>
<td>1.3</td>
</tr>
<tr>
<td></td>
<td>Nb</td>
<td>1.6</td>
</tr>
<tr>
<td></td>
<td>Mo</td>
<td>2.2</td>
</tr>
<tr>
<td></td>
<td>Tc</td>
<td>2.1</td>
</tr>
<tr>
<td></td>
<td>Ru</td>
<td>2.2</td>
</tr>
<tr>
<td></td>
<td>Rh</td>
<td>2.3</td>
</tr>
<tr>
<td></td>
<td>Pd</td>
<td>2.2</td>
</tr>
<tr>
<td></td>
<td>Ag</td>
<td>2.2</td>
</tr>
<tr>
<td></td>
<td>Cd</td>
<td>1.7</td>
</tr>
<tr>
<td></td>
<td>In</td>
<td>1.8</td>
</tr>
<tr>
<td></td>
<td>Sn</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>Sb</td>
<td>2.1</td>
</tr>
<tr>
<td></td>
<td>Te</td>
<td>2.1</td>
</tr>
<tr>
<td></td>
<td>I</td>
<td>2.7</td>
</tr>
<tr>
<td></td>
<td>Xe</td>
<td>2.6</td>
</tr>
<tr>
<td></td>
<td>Cs</td>
<td>0.8</td>
</tr>
<tr>
<td></td>
<td>Ba</td>
<td>0.9</td>
</tr>
<tr>
<td></td>
<td>La</td>
<td>1.1</td>
</tr>
<tr>
<td></td>
<td>Ce</td>
<td>1.3</td>
</tr>
<tr>
<td></td>
<td>Pr</td>
<td>1.5</td>
</tr>
<tr>
<td></td>
<td>Nd</td>
<td>1.5</td>
</tr>
<tr>
<td></td>
<td>Pm</td>
<td>1.9</td>
</tr>
<tr>
<td></td>
<td>Sm</td>
<td>1.7</td>
</tr>
<tr>
<td></td>
<td>Eu</td>
<td>1.9</td>
</tr>
<tr>
<td></td>
<td>Gd</td>
<td>2.2</td>
</tr>
<tr>
<td></td>
<td>Tb</td>
<td>2.2</td>
</tr>
<tr>
<td></td>
<td>Dy</td>
<td>2.2</td>
</tr>
<tr>
<td></td>
<td>Ho</td>
<td>2.4</td>
</tr>
<tr>
<td></td>
<td>Er</td>
<td>2.4</td>
</tr>
<tr>
<td></td>
<td>Tm</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>Yb</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>Lu</td>
<td>2.2</td>
</tr>
<tr>
<td></td>
<td>Th</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>Pa</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>U</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>Np</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>Pu</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>Am</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>Cm</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>Bk</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>Cf</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>Es</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>Fm</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>Md</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>No</td>
<td>2.0</td>
</tr>
<tr>
<td></td>
<td>Lr</td>
<td>2.0</td>
</tr>
</tbody>
</table>
Inter vs. Intra molecular Forces
Why Polarity Matters: Molecular Attractions

• 1 molecule can be attracted to another molecule
  – “inter”molecular force

• You can predict how one molecule might react with another: Ex: HBr + H₂O
Intermolecular Attractions & Smell

• Besides shape, polarity also plays a role in your ability to smell.
  – Polar molecules = smell
  – Non-polar = don’t smell

• Your smell receptors are polar and surrounded by mucous (a watery substance)

Ex: Methane gas is odorless

They add a this stinky chemical to it so that you can smell it:

\[
\text{CH}_4 \quad \text{nonpolar covalent bonds}
\]

\[
\text{CH}_3\text{SH}
\]
**Intermolecular Forces vs. Intramolecular Forces**

**Intramolecular Forces:** (within in a molecule)
- Covalent bond
- Ionic bond
- Metallic bond

**Intermolecular Forces:** (between molecules)
- Hydrogen bonds
- Weaker than covalent, ionic, & metallic bonds
Hydrogen Bonding—(an intermolecular force) in Water

- Water is polar
  - (has a + and - end)
  - It’s “sticky”

- Will stick to any other thing that is:
  - polar (ex: other water molecules)
  - charged ionic substances (NaCl)
Water’s Polarity leads to its ability to dissolve things so well

The slight charges on water attract the NaCl’s ions and cause them to separate from each other.
Unique Properties of Water due to polarity & hydrogen bonding

1) **Surface tension** (hydrogen bonds create surface on water)

2) **Adhesion/ Cohesion** (water is attracted to other water molecules)

3) **Capillary action** water is attracted to other water molecules and will “rise”
Properties of Water due to Hydrogen Bonding & Polarity

- **Cohesion** - water molecules are attracted to one another
  - Causes water to be “Sticky”
  - This is why water forms droplets

- **Adhesion** - water is attracted to other substances
  - Water will “stick” to containers & objects

- **Surface tension** - strong forces between molecules cause the surface of a liquid to contract
More properties...

- **Capillary Action** - the movement of water within the spaces of a porous material due to the forces of adhesion, cohesion, and surface tension.

- **Universal Solvent** things dissolve in water—polarity

Slightly positive hydrogen are attracted to chlorine anions
Slightly negative oxygen are attracted to sodium cations

Hydrogen bonding in Kevlar, a strong polymer used in bullet-proof vests.
Hydrogen Bonding in DNA

- Thymine
- Adenine
- Guanine
- Cytosine

D = Deoxyribose (sugar)
P = Phosphate

Hydrogen Bond
Other Intermolecular Forces
(FYI... not part of this class)

• Van der Waals Forces include:
  - **Dipole-Dipole forces** - results from the tendency of polar molecules to align themselves so that the positive end of one molecule is near the negative end of another molecule.
  - **London (Dispersion) forces** - results from the small, instantaneous dipoles that occur because of the varying positions of the electrons during their motion about nuclei
Organic Chemistry
Organic Chemistry shows the versatility of carbon:

- has 4 valence electrons = 4 bonding spaces available.

- Backbone to many large, complex biological molecules (Carbs, Lipids, Proteins, Nucleic Acids)

- Over 16 million carbon-containing compounds are known.
Monomers combine to make Polymers (small unit) (large)

Common Examples of Polymers:

- Carbohydrates
- Lipids
- Proteins
- Nucleic Acids (CLPN)
Ex: Carbohydrates

**Monomer**

**Monosaccharide**

**Polymer**

**Polysaccharide**

**Examples**

- Starch
- Fiber
- Sucrose
Ex: Lipids

Monomers
Glycerol & Fatty Acid tails

Polymer
Tri-glyceride

Examples
- Saturated Fats
- Unsaturated fats
- Steroids
- Cholesterol
Ex: Proteins

**Monomer**

Amino Acids

**Polymer**

Polypeptide

**Examples**

- enzymes
- pigments
- Meat/dairy
Ex: Nucleic Acids

Monomer | Polymer | Examples
---|---|---
Nucleotide | Polynucleotide | -DNA
| | | -RNA

[Diagram of nucleic acid structure with complementary base pairing and hydrogen bonds]
Distilled Water vs. Tap Water
Water Poisoning/ water Intoxication

**Cause:** excessive consumption of water during a short period of time.

**Why:** leads to a disruption in normal brain function due to the imbalance of **electrolytes** in the body’s fluids.

- can dilute the careful balance of **sodium** compounds in the body fluids

**Who:** individuals in water drinking contests...consume more than 10 liters (10.5 quarts) of water over the course of just a few minutes

- People doing endurance sports which electrolytes are not properly replenished, yet massive amounts of fluid are still consumed
Neural transmission

(a) Polarized membrane
(b) Depolarization (sodium ions flow in)
(c) Sodium ions pumped out of neuron
(d) Flow of depolarization

Dendrites
Direction of impulse
To next neuron
Axon
Nucleus
Direction of impulse
To next neuron
Electric Stimulation Machine—stimulates muscles for you

See video clips on web links