**BONDING**

**Chemical Bond**
Attraction that holds atoms together

- Types include **IONIC**, **METALLIC**, or **COVALENT**
- Differences in electronegativity determine the bond type

**Ionic Bond**
TRANSFER of electrons between atoms

- Each atom achieves a noble gas configuration (full valence shell)
- Usually between a METAL and a NONMETAL
- Formula Unit: lowest whole-number ratio of ions in an ionic compound (ex: NaCl or MgCl₂)

**Ionic Bonding**

- **EXAMPLE:** Show the electron dot structures for sodium and chlorine using arrows to indicate the transfer of e⁻.

  - **Na⁺** (Wants to lose ONE electron!)
  - **Cl⁻** (Wants to gain ONE electron!)

**To write the formula unit for the compound:**
write the element symbol for the cation followed by the anion symbol... to indicate the # of each use subscripts (none needed if just one)!!
Ionic Bonding

- **EXAMPLE:**
  Show the electron dot structures for sodium and chlorine using arrows to indicate the transfer of e-.

\[ \text{Na}^+ \quad \text{Cl}^- \]

\[ \text{NaCl} \]

- **EXAMPLE:**
  Show the electron dot structures for calcium and phosphorus using arrows to indicate the transfer of e-.

\[ \text{Ca}^{2+} \quad \text{P}^3- \]

**Wants to lose TWO electrons!**

**Wants to gain THREE electrons!**

**PHOSPHORUS STILL NEEDS ONE MORE ELECTRON!!!**

**Ca**

**CALCIUM STILL NEEDS TO LOSE ONE MORE ELECTRON!!!**

\[ \text{Ca}^{2+} \quad \text{P}^3- \]

**Ca**
**Ionic Bonding**

- **EXAMPLE:** Show the electron dot structures for calcium and phosphorus using arrows to indicate the transfer of e-.

\[
\begin{align*}
\text{Ca}^{2+} & \quad \text{P}^{3-} \\
\text{Ca}^{2+} & \quad \text{P}^{3-} \\
\text{Ca}^{2+} & \quad \text{Ca}_3\text{P}_2
\end{align*}
\]

**Coulomb’s Law**

*Determines the energy / strength of an ionic bond*

\[
F = \frac{kq_1q_2}{r^2}
\]

- \( r \) = distance between ions in nm (SIZE MATTERS)
- \( Q_1 \) and \( Q_2 \) = numerical ion charges (with signs)
- \( F \) = electrical force between the atoms

**Lattice Energy**

*Change in energy that takes place when separated GASEOUS ions are packed together to form an ionic solid... (tells how much energy needed to separate ions)*

- Ionic compounds form solid crystals organized in a CRYSTAL LATTICE of alternating + and - ions
- Relates to Coulomb’s Law:
  - GREATER charges and SMALLER size mean HIGHER lattice energy (STRONGER BOND)

*Coulomb’s Law* determines the energy / strength of an ionic bond:

- Force of attraction between two oppositely charged particles is directly proportional to the MAGNITUDE of the charges and inversely proportional to the DISTANCE between those charges...

  - GREATER charges = GREATER attraction
  - CLOSER (smaller ions) = GREATER attraction
  - (Attraction \( \rightarrow \) Opposites / Repulsion \( \rightarrow \) Same)

  *When calculated for an ionic compound, energy is NEGATIVE, indicating an attractive force that has a LOWER energy than the separated ions!!

*Lattice Energy* energy needed to separate ions (how strong is the ionic bond)
Lattice Energy

Energy needed to separate ions (how strong is the ionic bond)

<table>
<thead>
<tr>
<th>Cation</th>
<th>F⁻</th>
<th>O²⁻</th>
<th>Na⁺</th>
<th>Cl⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li⁺</td>
<td>3036</td>
<td>855</td>
<td>169</td>
<td>2925</td>
</tr>
<tr>
<td>Na⁺</td>
<td>923</td>
<td>787</td>
<td>704</td>
<td>2905</td>
</tr>
<tr>
<td>K⁺</td>
<td>823</td>
<td>515</td>
<td>469</td>
<td>2364</td>
</tr>
<tr>
<td>Mg²⁺</td>
<td>3987</td>
<td>3203</td>
<td>214</td>
<td>2000</td>
</tr>
<tr>
<td>Ca²⁺</td>
<td>3633</td>
<td>2526</td>
<td>2176</td>
<td>2074</td>
</tr>
<tr>
<td>Al³⁺</td>
<td>3213</td>
<td>3492</td>
<td>3764</td>
<td>3218</td>
</tr>
</tbody>
</table>

**Lattice Energy**

**Energy needed to separate ions (how strong is the ionic bond)**

**Example: Formation of lithium fluoride**

<table>
<thead>
<tr>
<th>Process</th>
<th>Description</th>
<th>Energy Change (kJ)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li(s) → Li(g)</td>
<td>Sublimation energy</td>
<td>161</td>
</tr>
<tr>
<td>Li(g) → Li⁺(g) + e⁻</td>
<td>Ionization energy</td>
<td>500</td>
</tr>
<tr>
<td>½Cl₂(g) → Cl(g)</td>
<td>Bond energy (½ mole)</td>
<td>77</td>
</tr>
<tr>
<td>Cl(g) + e⁻ → Cl⁻(g)</td>
<td>Electron affinity</td>
<td>-328</td>
</tr>
<tr>
<td>Li⁺(g) + Cl⁻(g) → LiCl(s)</td>
<td>Lattice energy</td>
<td>-7547</td>
</tr>
<tr>
<td>Li(s) + ½Cl₂(g) → LiCl(s)</td>
<td>ΔH</td>
<td>-417</td>
</tr>
</tbody>
</table>

*Notice the “-” energy indicating the ions are attracted and the lattice energy is high because both atoms have a small atomic size!*

**Lattice Energy**

Calculate the ΔHₚ for Sodium Chloride:

Na(s) + ½ Cl₂(g) → NaCl(s)

-786 kJ/mol

**Lattice Energy**

Calculate the ΔHₚ for Sodium Chloride:

Na(s) + ½ Cl₂(g) → NaCl(s)

-786 kJ/mol

**Lattice Energy**

Calculate the ΔHₚ for Sodium Chloride:

Na(s) + ½ Cl₂(g) → NaCl(s)

-786 kJ/mol

**Properties of Ionic Compounds**

1) **STRONGLY** bonded: strong attraction between the ions

2) Solid **CRYSTAL** lattice structure

**Ionic Compounds**

- Formed from **METAL** and **NONMETAL** ions bonded together
- Known as “**SALTS**”
- Formed from regular repeating arrangement of formula units

**Properties of Ionic Compounds**

1) **STRONGLY** bonded: strong attraction between the ions

2) Solid **CRYSTAL** lattice structure

3) **HIGH** melting point and boiling point... due to the strong attractions between ions and the stable structure

4) **SOLUBLE** in water / **DISSOCIATE** (break apart into ions)
Properties of Ionic Compounds

5) Conduct **ELECTRICITY** when melted or dissolved in water (because they form ions)... too orderly when solid

Properties of Ionic Compounds

6) Extremely **BRITTLE**... if you hit them hard enough they will shatter because they don’t want to bend and there will be a strong repulsive force

Lattice Energy and Properties

• Which has a **STRONGER** bond? NaCl or KCl

• Which is more **SOLUBLE** in water? NaCl or AlCl₃

Metallic Bonds

Attraction of the free-floating valence electrons for the positively charged metal ions

• Forces of attraction hold metals together
• Not ionic
• Have similar properties to ionic compounds
• Metals weakly hold on to their valence e-
• Positive ions (cations) are floating in a “SEA OF ELECTRONS”

“Sea of Electrons”

ELECTRONS ARE FREE TO MOVE!!

Properties of Metals

1) Malleable
2) Ductile
3) Conductive

All due to the fact that their valence electrons are **MOBILE!**
**Properties of Metals**

Electrons allow the atoms of metals to **SLIDE** like grease!

**Alloys**

*Mixtures of two or more elements with at least one being a metal*

- Formed by melting a mix of ingredients and then cooling
- Examples: brass (Cu and Zn), bronze (Cu and Sn), and steel (Fe, C, etc.)

**Why Make Alloys?**

- Properties are **SUPERIOR** to individual elements
- Sterling Silver (92.5% Ag, 7.5% Cu): harder, more durable than pure Ag / soft enough to work
- Steels: corrosion resistant, ductile, hard, tough, and cost effective

*MORE TO COME IN SECTION 3!*

**Covalent Bond**

*SHARING of electrons between atoms*

- Involves two **NONMETALS**
- Known as covalent or molecular compounds
- **MOLECULE**: group of atoms joined by a covalent bond
- **DIATOMIC MOLECULES**: elements that cannot exist as single atoms
  - Ex: H₂, N₂, O₂, F₂, Cl₂, Br₂, and I₂

**Bond Length**

*Distance where energy is at a MINIMUM*

- When two atoms approach, two “bad” things occur... electron/electron repulsion and proton/proton repulsion
- One “good” thing occurs... proton/electron attraction
- When **ATTRACTIVE** forces offset **REPULSIVE** forces, energy decreases and a bond is formed... Always looking for LOWEST energy!
**Bond Length**

*Distance where energy is at a MINIMUM*

*Note how to find the bond length and energy from the graph!!*

**Properties of Covalent Compounds**

1) Low melting and boiling points
2) Don’t usually conduct electricity
3) Not usually soluble in water (some dissolve but don’t dissociate... SUGAR!)
4) Don’t form crystals (most liquids or gases at room temperature)

**Dissociation vs. Dissolving**

**Dissociation**

\[ \text{Na}^+ \quad \text{Cl}^- \]
\[ \text{Na}^+ \quad \text{Cl}^- \]
\[ \text{Na}^+ \quad \text{Cl}^- \]
\[ \text{Na}^+ \quad \text{Cl}^- \]

Salt = **IONIC**

**Dissolving**

\[ \text{C}_12\text{H}_{22}\text{O}_{11} \]
\[ \text{C}_12\text{H}_{22}\text{O}_{11} \]
\[ \text{C}_12\text{H}_{22}\text{O}_{11} \]
\[ \text{C}_12\text{H}_{22}\text{O}_{11} \]

Sugar = **COVALENT**

**Properties of Covalent Compounds**

![Salt vs. Sugar](image)

**Attractions and Properties**

- Wide range of physical properties among covalent compounds due to varying intermolecular attractions
- **NETWORK SOLIDS**: very stable structure consisting of all atoms being covalently bonded to each other
  - High melting points (1000°C or higher)
  - To melt, ALL the covalent bonds need to be broken
  - EX: Diamonds, SiO₂, SiC

**Types of Covalent Bonding**

- There are three types of covalent bonds:
  1) **SINGLE**: sharing of only one pair of e⁻ (2 total)
  2) **DOUBLE**: sharing of two pairs of e⁻ (4 total)
  3) **TRIPLE**: sharing of three pairs of e⁻ (6 total)

Remember the octet rule... **EIGHT** electrons are needed!!
As the number of bonds between two atoms **INCREASE**, the bond length **DECREASES**... the bonds become **SHORTER** and **STRONGER**!!

**WHY?**: More electrons between the nuclei which **DECREASES** nuclear repulsions and **INCREASES** nucleus to electron attractions!!

**Molecular Dot Structures**

- **Visual representation of how the atoms are bonded together in a molecule**
- **Shows valence e- as dots... can see type of covalent bond as well (sharing of **TWO** dots = a **SINGLE** bond)**

**Molecular Dot Structures**

- **How to draw a structure:**
  1) Determine the **TOTAL** number of valence e- available for the entire molecule from the amount each element has... This is the # of e- that MUST be in the final structure
  2) Position any C or N atoms in the center... With atoms other than C or N, put the least electronegative atom in the center (usually the atom that there is **LESS** of in the formula)... H never in center unless it’s the only element
  3) Whatever atoms are remaining in the formula go around the “center” (evenly distributed when possible)

- **How to draw a structure:**
  4) Any atoms next to the center need to **SHARE** at least a pair of electrons between the atoms (covalent bonding)... Insert pairs of e- between bonded atoms
  5) Give each atom a complete octet by adding **UNSHARED** (lone) pairs until each atom is “happy” with its number
  6) Total the e- in the molecule and see if total from step #1 is reached... If so, then done
  7) **IF OVER** the #, add double or triple bonds and remove pairs... **IF UNDER** the #, add lone pairs (extra) to the center atom

**HONC 1234**

- **Certain elements** **USUALLY** have a specific number of bonds (shared pairs) in a molecule!!
  - **Hydrogen** – **1 bond**
  - **Oxygen** – **2 bonds**
  - **Nitrogen** – **3 bonds**
  - **Carbon** – **4 bonds**

**HONC IF YOU LOVE CHEMISTRY!!**

**Molecular Dot Structures**

- **EXAMPLE:**
  Write the dot structure for: NH₃.
Molecular Dot Structures

• EXAMPLE:
  Write the dot structure for: H₂O.

Molecular Dot Structures

• EXAMPLE:
  Write the dot structure for: CO₂.

Molecular Dot Structures

• EXAMPLES:
  Write the dot structure for:

Exceptions to the Octet Rule

If the number of electrons doesn’t work out...

1) H - full valence shell is 2 e-
2) B - generally satisfied with 6 valence e-
3) Be - generally satisfied with 4 valence e-
4) N - can be satisfied with 7 valence e-
5) As, S, I, Se and P can expand their octet to have 10, 12, or 14 valence e-

Exceptions to the Octet Rule

Why are some elements able to EXPAND their octet and take more than EIGHT?

• Can occur when an element has an EMPTY d-orbital... Electrons get put in there!
• Notice this only occurs in elements in the 3rd period or higher... WHY?
• Ex: Phosphorus

Structural Formulas

• Uses a dash (-) to represent a covalent bond
• Unshared pairs still shown as dots
**Structural Formulas**

- Uses a dash (-) to represent a covalent bond
- Unshared pairs still shown as dots

**Resonance Structures**

*When more than one valid dot structure is possible for the same substance*

- Found in many **DOUBLE** bonded molecules
- Ex: CO$_3^{2-}$ or C$_6$H$_6$
- Each MUST be shown in brackets with charge in upper right... include double arrow as well!

**Formal Charge**

*Tells which atom carries the charge in nonequivalent Lewis structures... assumes perfect sharing of electrons*

\[
FC = \frac{\text{# Valence } e^- \text{ in Free Atom}}{ \text{Group #} } - \frac{\text{Total # of Nonbonding } e^-}{2} - \frac{\text{Total # of Bonding } e^-}{2}
\]

- Sum of all FC **MUST EQUAL** the ion's charge (if no charge, then should add up to zero)

*If MORE than one possible structure could be drawn...*

1. Want structure with the **LOWEST** sum when the absolute values of all the FCs are added
2. If still more than one, structure with the "-" FC on the **MOST ELECTRONEGATIVE** element is the best
Formal Charge

Tells which atom carries the charge in nonequivalent Lewis structures... assumes perfect sharing of electrons

**EXAMPLE:**
Which is the correct structure for sulfuric acid?

\[
\begin{align*}
\text{H} & \quad \text{O} & \quad \text{S} & \quad \text{O} & \quad \text{H} \\
\vdots & \quad \vdots & \quad \vdots & \quad \vdots & \quad \vdots
\end{align*}
\]

\[
\begin{align*}
\text{H} & \quad \text{O} & \quad \text{S} & \quad \text{O} & \quad \text{H} \\
\vdots & \quad \vdots & \quad \vdots & \quad \vdots & \quad \vdots
\end{align*}
\]

Formal Charge

Tells which atom carries the charge in nonequivalent Lewis structures... assumes perfect sharing of electrons

**EXAMPLE:**
Which is the correct structure for SCN\(^-\)?

\[
\begin{align*}
\text{S} & \quad \text{C} & \quad \text{N} \\
\vdots & \quad \vdots & \quad \vdots
\end{align*}
\]


Bond Polarity

- Covalent bonds involve the sharing of electrons but not all atoms share equally!
- Electrons are in a "TUG OF WAR" between atoms

\[
\begin{align*}
\text{H} & \quad \text{F} \\
\end{align*}
\]

Nonpolar Covalent Bonds

**EQUAL** sharing of electrons

- Compounds **NOT** soluble in water ("like dissolves like")
- LOWER boiling and melting point than polar compounds

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\end{align*}
\]

Polar Covalent Bonds

**Electrons are NOT** shared equally (one atom pulls the e- more)

- Compounds are **SOLUBLE** in water ("like dissolves like")
- HIGHER boiling and melting point than nonpolar compounds

\[
\begin{align*}
\text{D} & \quad \text{H} \\
\end{align*}
\]

\[
\begin{align*}
\text{O}_2 \\
\end{align*}
\]
**Polarity**

- Difference in **Electronegativity** (ability of an atom to pull e− towards itself) between the atoms determines what type of bond forms...

<table>
<thead>
<tr>
<th>Electronegativity Difference</th>
<th>Type of Bond</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.0 - 0.4</td>
<td>Nonpolar covalent</td>
<td>S-Se</td>
</tr>
<tr>
<td>0.5 - 2.0</td>
<td>Polar covalent</td>
<td>Cl-P</td>
</tr>
<tr>
<td>&gt; 2.0</td>
<td>Ionic</td>
<td>Na-F</td>
</tr>
</tbody>
</table>

**Example:**

- HCl
  - **H** = electronegativity of 2.1
  - **Cl** = electronegativity of 3.0
  - **Polar bond** → difference is 0.9
  - Cl “hogs” shared e− giving it a slight negative charge
  - H has shared e− less time giving it a slight positive charge

**Partial Charges**

- Slight regions of charge that form due to e− “hogging”
- Much less than a 1+ or 1− in ionic bonding
- **Delta** denotes partial charges (δ+, δ−)
- Written as:
  
  ![Partial Charges Diagram](image)

**Polar Molecules**

- For a molecule to be **Polar** it must have:
  - At least one polar bond
  - Asymmetrical shape (lone pairs) or asymmetrical atoms

**Example:**

- **H₂O**
  - 2 Polar Bonds (O–H)
  - Highly electronegative O pulls the e− away from the H
  - Asymmetrical shape

**Example:**

- **CO₂**
  - 2 Polar Bonds (C–O)
  - Highly electronegative O pulls the e− away from the C
  - Symmetrical shape / cancels dipole

**Just having polar bonds does NOT make a molecule polar!!**
Properties

"LIKE DISSOLVES LIKE"

Bond Enthalpy

Energy required to break the bond between two covalently bonded atoms (aka Bond Energy)

• Bonds BREAKING = Endothermic
• Bonds FORMING = Exothermic

• Change in enthalpy (energy) for a reaction can be found using the bond energies:

\[ \Delta H_{\text{rxn}} = \Sigma (\text{energies of bonds BROKEN}) - \Sigma (\text{energies of bonds FORMED}) \]

Average Bond Energies (kJ/mol)

<table>
<thead>
<tr>
<th>Bond</th>
<th>Energy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H-H</td>
<td>432</td>
</tr>
<tr>
<td>H-F</td>
<td>565</td>
</tr>
<tr>
<td>H-Cl</td>
<td>427</td>
</tr>
<tr>
<td>H-Br</td>
<td>363</td>
</tr>
<tr>
<td>H-I</td>
<td>295</td>
</tr>
<tr>
<td>C-H</td>
<td>413</td>
</tr>
<tr>
<td>C-C</td>
<td>347</td>
</tr>
<tr>
<td>C-N</td>
<td>305</td>
</tr>
<tr>
<td>C-O</td>
<td>358</td>
</tr>
<tr>
<td>C-F</td>
<td>485</td>
</tr>
<tr>
<td>C-Cl</td>
<td>339</td>
</tr>
<tr>
<td>C-I</td>
<td>240</td>
</tr>
<tr>
<td>C-S</td>
<td>259</td>
</tr>
</tbody>
</table>

Higher energy = STRONGER bond!

Examples:

1. Calculate the change in enthalpy for the reaction using bond energies.
   \[ \text{H}_2 + \text{CO}_2 \rightarrow \text{H}_2\text{O} + \text{CO} \]

2. Calculate the change in enthalpy for the reaction using bond energies.
   \[ \text{CO} + 2 \text{H}_2 \rightarrow \text{CH}_3\text{OH} \]
**Bond Enthalpy**

**EXAMPLE:**
The formation of ethyl butanoate, one of the compounds that give pineapple its flavor, is produced according to the reaction below. Calculate the change in enthalpy for the reaction using bond energies.

\[
\text{CH}_3\text{CH}_2\text{CH}_2\text{CO}_2\text{H} + \text{HO} = \text{CH}_2\text{CH}_2\text{OH} \rightarrow \text{CH}_3\text{CH}_2\text{CH}_2\text{CO}_2\text{H} + \text{HO} = \text{H}
\]

(Butanoic acid) (Ethanol) (Ethyl butanoate) (Water)

---

**VSEPR Theory**

*Valence Shell Electron Pair Repulsion*

Valence electron pairs attempt to get as FAR APART as possible, thus changing the 3D shape of the molecule!!!

- Molecules are really 3D, not 2D
- Unshared (LONE) pairs are held closer to the atom and they repel other lone pairs and bonding pairs which pushes them closer together

---

**Common Molecular Shapes**

- **LINEAR:**
  
  \[ \text{BeH}_2 \]

  Type: \( \text{AB}_2 \)
  
  180° bond angle
  
  2 atoms attached
  
  0 lone pairs

---

- **LINEAR (another):**
  
  \[ \text{XeF}_2 \]

  Type: \( \text{AB}_2\text{E}_3 \)
  
  180° bond angle
  
  2 atoms attached
  
  3 lone pairs

---

- **BENT:**
  
  \[ \text{H}_2\text{O} \]

  Type: \( \text{AB}_2\text{E}_2 \)
  
  104.5° bond angle
  
  2 atoms attached
  
  2 lone pairs (or just 1)
Common Molecular Shapes

- **TRIGONAL PLANAR:**
  \[ \text{BF}_3 \]
  \[ \text{Type: AB}_3 \]
  \[ 120^\circ \text{ bond angle} \]
  \[ 3 \text{ atoms attached} \]
  \[ 0 \text{ lone pairs} \]

- **TRIGONAL PYRAMIDAL:**
  \[ \text{NH}_3 \]
  \[ \text{Type: AB}_3\text{E} \]
  \[ 107.3^\circ \text{ bond angle} \]
  \[ 3 \text{ atoms attached} \]
  \[ 1 \text{ lone pair} \]

- **T-SHAPED:**
  \[ \text{ClF}_3 \]
  \[ \text{Type: AB}_3\text{E}_2 \]
  \[ 90^\circ/180^\circ \text{ bond angle} \]
  \[ 3 \text{ atoms attached} \]
  \[ 2 \text{ lone pairs} \]

- **TETRAHEDRAL:**
  \[ \text{CH}_4 \]
  \[ \text{Type: AB}_4 \]
  \[ 109.5^\circ \text{ bond angle} \]
  \[ 4 \text{ atoms attached} \]
  \[ 0 \text{ lone pairs} \]

- **SEESAW:**
  \[ \text{SF}_4 \]
  \[ \text{Type: AB}_4\text{E}_1 \]
  \[ 90^\circ/120^\circ/180^\circ \text{ angles} \]
  \[ 4 \text{ atoms attached} \]
  \[ 1 \text{ lone pair} \]

- **SQUARE PLANAR:**
  \[ \text{XeF}_4 \]
  \[ \text{Type: AB}_4\text{E}_2 \]
  \[ 90^\circ/180^\circ \text{ bond angle} \]
  \[ 4 \text{ atoms attached} \]
  \[ 2 \text{ lone pairs} \]
Common Molecular Shapes

• **TRIGONAL BIPYRAMIDAL**:
  - PCl$_5$
  - Type: AB$_5$
  - 90°/120° bond angle
  - 5 atoms attached
  - 0 lone pairs

• **SQUARE PYRAMIDAL**:  
  - ClF$_5$
  - Type: AB$_5$E
  - 90°/180° bond angle
  - 5 atoms attached
  - 1 lone pair

• **OCTAHEDRAL**:  
  - SF$_6$
  - Type: AB$_6$
  - 90° bond angle
  - 6 atoms attached
  - 0 lone pairs

Electronic vs. Molecular

• **ELECTRONIC DOMAINS**: based on the electron groups on the central atom (ignores shared vs. lone pairs) / helps with bond angle

• **MOLECULAR GEOMETRY**: used to describe the shape of the actual molecule

• EXAMPLE: H$_2$O
  - 4 Electron Domains... Tetrahedral Electron Geometry! (4 attached)
  - But, a Bent Molecular Geometry! (2 attached and 2 lone pairs)

*NOTE: Double or triple bonds in a molecule often give linear or planar shapes!!*
Orbital Hybridization Theory

Combines information about molecular bonding and molecular shape

Pug + Beagle = Puggle

s orbital + p orbital = sp orbital

Orbital Hybridization Theory

Let's look at CARBON:

Can you see a problem with this?

Carbon only has TWO electrons available for bonding... That is not enough!

So what’s going on here??

Orbital Hybridization Theory

First thought was that carbon promoted one of its 2s electrons to the empty 2p orbital

• Now try to bond to four hydrogen atoms to form methane (CH₄)

Orbital Hybridization Theory

• Three of the C-H bonds would involve a 2p matched with a 1s, but a fourth bond would be between a 2s and a 1s

• This one bond would have slightly LESS energy than the other bonds in methane... this is NOT what chemists observe though!

Orbital Hybridization Theory

Measurements show that all four bonds are EQUAL so they must be HYBRIDIZED

• For methane, the s orbital combines with the three p orbitals to create four equal sp³ hybridized orbitals

New orbitals have slightly MORE energy than the 2s and slightly LESS energy than the 2p!

Orbital Hybridization Theory

Notice the energy is in between the 2s and 2p!
Orbital Hybridization Theory

Combines information about molecular bonding and molecular shape

- Covalent bonds form when atomic orbitals overlap
- Mixing of s, p, and sometimes d orbitals that allows bonds to form... HYBRID ORBITALS
- Ex: BeF₂

Shape Hybridizations

- LINEAR (2 attached): sp
- TRIGONAL PLANAR (3 attached): sp²
- TETRAHEDRAL (4 attached): sp³
- TRIGONAL BIPYRAMIDAL (5 attached): sp³d
- OCTAHEDRAL (6 attached): sp³d²

*Each lone pair of e⁻ counts as an “attached” group too!!
Depends on which types of orbitals are being mixed!

Multiple Bond Hybridizations

Single Bonds

(Just COUNT number of bonds to determine)

Double Bonds

sp²

Triple Bonds

sp

Sigma (σ) Bonds

Electrons are located around the central axis of a bond

- Single bonds
  - Ex: H₂

Pi (π) Bonds

Electrons are located above and below the axis of a bond

- Weaker than sigma bonds
- Double bonds = 1 sigma / 1 pi
- Triple bonds = 1 sigma / 2 pi
- Ex: O₂ / N₂