**BONDING**

**Chemical Bond**
Attraction that holds atoms together

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

**Electrostatic Attraction**
Two objects near each other with different / opposite electrical charges

- Forces of attraction that allow atoms to bond... +/- ions attracted to each other in an ionic bond OR the positive nuclei attracted to the shared e- in covalent

**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**

- How to show using **Electron Dot Structures**:
  1) Draw the dot structure for each element in the ionic compound
  2) Determine which element will lose e- and which will gain e- and how many
  3) Use arrows to show the e- being transferred to an empty space on the anion’s dot structure
  4) Continue until each element has a complete octet... add more of each element as needed

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

Wants to lose **ONE** electron!  
Wants to gain **ONE** electron!
**Ionic Bonding**

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

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

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)!!

\[
\text{NaCl}
\]

**Ionic Bonding**

- **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!!!!**

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

Wants to lose **TWO** electrons!

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

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

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

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

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

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

\[ \text{Ca}^{2+} \quad \text{Ca}_3\text{P}_2 \]

Ionic Bonding

**MORE EXAMPLES:**
Show the electron dot structures for each pair and show the transfer of electrons using arrows. Be sure to include the charge on each ion after the transfer and write the formula unit.

- **LITHIUM** and **FLUORINE**
- **ALUMINUM** and **SULFUR**
- **MAGNESIUM** and **CHLORINE**
- **POTASSIUM** and **OXYGEN**

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!**)

Can be used to help determine properties of ionic compounds...
Lattice Energy

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

<table>
<thead>
<tr>
<th>Cation</th>
<th>F⁻</th>
<th>Cl⁻</th>
<th>Br⁻</th>
<th>I⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li⁺</td>
<td>1016</td>
<td>853</td>
<td>797</td>
<td>767</td>
</tr>
<tr>
<td>Na⁺</td>
<td>923</td>
<td>707</td>
<td>704</td>
<td>689</td>
</tr>
<tr>
<td>K⁺</td>
<td>821</td>
<td>715</td>
<td>662</td>
<td>648</td>
</tr>
<tr>
<td>Ba²⁺</td>
<td>3905</td>
<td>3203</td>
<td>2914</td>
<td>2868</td>
</tr>
<tr>
<td>Mg²⁺</td>
<td>2997</td>
<td>2956</td>
<td>2640</td>
<td>2527</td>
</tr>
<tr>
<td>Ca²⁺</td>
<td>2630</td>
<td>2256</td>
<td>2176</td>
<td>2074</td>
</tr>
<tr>
<td>Al³⁺</td>
<td>5215</td>
<td>5462</td>
<td>5316</td>
<td>5216</td>
</tr>
</tbody>
</table>

**SMALLER** ions of the same charge = **STRONGER** bond!

**GREATER** charge = **STRONGER** bond!

*Lattice Energy*

Calculate the ΔH₂ for Sodium Chloride:

\[ \text{Na}(s) + \frac{1}{2} \text{Cl}_2(g) \rightarrow \text{NaCl}(s) \]

\[ \text{Lattice Energy} = -786 \text{ kJ/mol} \]

\[ \text{Ionization Energy for Na} = 495 \text{ kJ/mol} \]

\[ \text{Electron Affinity for Cl} = -349 \text{ kJ/mol} \]

\[ \text{Bond energy of Cl}_2 = 239 \text{ kJ/mol} \]

\[ \text{Enthalpy of sublimation for Na} = 109 \text{ kJ/mol} \]

\[ \text{Na}(g) \rightarrow \text{Na}(g) + 109 \text{ kJ} \]

\[ \text{Na}(g) \rightarrow \text{Na}⁺(g) + e⁻ + 495 \text{ kJ} \]

\[ \frac{1}{2} \text{Cl}_2(g) \rightarrow \text{Cl}⁻(g) + \frac{1}{2}(239 \text{ kJ}) \]

\[ \text{Cl}⁻(g) + e⁻ \rightarrow \text{Cl}⁻(g) - 349 \text{ kJ} \]

\[ \text{Na}⁺(g) + \text{Cl}⁻(g) \rightarrow \text{NaCl}(s) - 786 \text{ kJ} \]

\[ \text{Na}(s) + \frac{1}{2} \text{Cl}_2(g) \rightarrow \text{NaCl}(s) - 412 \text{ kJ/mol} \]

*Lattice Energy*

*Lattice Energy*

*Lattice Energy*

*Lattice Energy*

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

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
Properties of Ionic Compounds

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)

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!

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!
**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)

**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!!

Bond Length

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)

Molecular Dot Structures

• 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

DO NOT overfill an atom’s valence shell... Ex: H only wants two electrons!!

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

Atoms still need a FULL valence!!

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, P, etc. 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

\[
\begin{align*}
\text{H} & \quad \rightarrow \quad \text{H-O-H} \\
\text{O} & \quad \rightarrow \quad \text{N=\text{N}}
\end{align*}
\]

**Resonance Structures**

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

- Found in many **DOUBLE** bonded molecules
- Ex: \( \text{CO}_2^2- \) or \( \text{C}_6\text{H}_6 \)
- 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 (more than one possibility) Lewis structures... assumes perfect sharing of electrons

\[
\text{FC} = \frac{\# \text{ Valence e}^- \text{ in Free Atom} - \text{Total } \# \text{ Nonbonding e}^- - \frac{\text{Total } \# \text{ Bonding e}^-}{2}}{\text{Group } \# \text{ (Ex: 4A)}}
\]

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

**Experiments show**
that the molecules with resonance have **EQUAL** bond lengths and strengths!
**Formal Charge**

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

\[
\text{FC} = \frac{\# \text{ Valence } e^{-} \text{ in Free Atom} - \frac{\text{Total # Nonbonding } e^{-}}{\text{Group # (Ex: 4A)}} - \frac{\text{Total # Bonding } e^{-}}{2}}{\text{Or just # of bonds}}
\]

- 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

**Example:**

Which is the correct structure for sulfuric acid?

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

**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⁻?

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

**Bond Polarity**

*Covalent bonds involve the sharing of electrons but not all atoms share equally!*

*Electrons are in a "TUG OF WAR" between atoms*

**Nonpolar Covalent Bonds**

*EQUAL sharing of electrons*

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

**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>

**Polarity**
- EXAMPLE:
  - H = electronegativity of 2.1
  - Cl = electronegativity of 3.0
  - POLAR BOND  $\rightarrow$ 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 ($\delta^+$, $\delta^-$)
- Written as:
  $\begin{array}{c}
  \delta^+ \\
  H \rightarrow Cl
  \end{array}$
  or
  $\begin{array}{c}
  \delta^-
  H \rightarrow Cl
  \end{array}$

Having a polar bond and the right shape make HCl a POLAR MOLECULE and it is a DIPOLE (molecule with two poles)!!

**Polar Molecules**
- For a molecule to be POLAR it must have:
  - At least one polar bond
  - Asymmetrical shape (lone pairs on center...) sometimes helps ID this) or asymmetrical atoms

**Polar Molecules**
- EXAMPLE:
  - 2 Polar Bonds (O-H)
  - Highly electronegative O pulls the e- away from the H
  - Asymmetrical shape

Just having polar bonds does NOT make a molecule polar!!
Polar Molecules

- EXAMPLE:

\[
\text{CO}_2
\]

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

NONPOLAR MOLECULE

\[
O=\overset{\delta^-}{\text{C}}=\overset{\delta^+}{\text{O}} = \overset{\delta^-}{\text{O}}=\overset{\delta^+}{\text{C}}=\overset{\delta^-}{\text{O}}
\]

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)

- H-H 432
- H-F 565
- H-Cl 427
- H-Br 363
- H-I 295
- C-H 413
- C-C 347
- C-N 305
- C-O 358
- C-F 465
- C=O 397
- C=Cl 339
- C=Br 276
- C=S 259
- N-H 391
- N-N 160
- N-F 272
- N-Cl 200
- N-O 201
- O-H 467
- O-Cl 476
- O-Cl 676
- O=O 146
- O=F 190
- O=S 238
- O=Cl 203
- O=C 234
- O=O 495
- C-C 218
- C-Br 193
- C-Cl 149
- C-I 208
- C-N 193
- C-O 607
- N=O 146
- S=O 203
- Si=Si 340
- Si=O 452
- Higher energy = STRONGER bond!

Bond Enthalpy

- How to Calculate for a Reaction:
  1) Draw the dot structures for all the substances in order to see which bonds will break and form
  2) Sum the energies of ALL of the bonds that are broken and subtract the sum of the energies of ALL of the bonds that are formed
  3) Amounts of each substance (from the reaction equation) need to be considered when calculating

EX:

\[
2 \text{H}_2\text{O}_2 \rightarrow 2 \text{H}_2\text{O} + \text{O}_2
\]

Higher energy = STRONGER bond!
**Bond Enthalpy**

- **EXAMPLE:** Calculate the change in enthalpy for the reaction using bond energies.

\[
H_2 + CO_2 \rightarrow H_2O + CO
\]

\[
CO + 2 H_2 \rightarrow CH_3OH
\]

**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

- **LINEAR:**
  - **BeH₂**
  - 180° bond angle
  - 2 atoms attached
  - 0 lone pairs

**Common Molecular Shapes**

- **LINEAR (another):**
  - **XeF₂**
  - Type: \(AB₂\)
  - 180° bond angle
  - 2 atoms attached
  - 3 lone pairs
Common Molecular Shapes

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

  Type: \( AB_2E_2 \)
  
  104.5° bond angle
  
  2 atoms attached
  
  2 lone pairs (or just 1)

- **TRIGONAL PLANAR**:
  
  \[ \text{BF}_3 \]

  Type: \( AB_3 \)
  
  120° bond angle
  
  3 atoms attached
  
  0 lone pairs

- **TRIGONAL PYRAMIDAL**:
  
  \[ \text{NH}_3 \]

  Type: \( AB_3E \)
  
  107.3° bond angle
  
  3 atoms attached
  
  1 lone pair

- **T-SHAPED**:
  
  \[ \text{ClF}_3 \]

  Type: \( AB_3E_2 \)
  
  90°/180° bond angle
  
  3 atoms attached
  
  2 lone pairs

- **TETRAHEDRAL**:
  
  \[ \text{CH}_4 \]

  Type: \( AB_4 \)
  
  109.5° bond angle
  
  4 atoms attached
  
  0 lone pairs

- **SEESAW**:
  
  \[ \text{SF}_4 \]

  Type: \( AB_4E_1 \)
  
  90°/120°/180° angles
  
  4 atoms attached
  
  1 lone pair
Common Molecular Shapes

- **SQUARE PLANAR**: XeF$_4$
  - Type: $AB_4E_2$
  - $90^\circ$/$180^\circ$ bond angle
  - 4 atoms attached
  - 2 lone pairs

Common Molecular Shapes

- **TRIGONAL BIPYRAMIDAL**: PCl$_5$
  - Type: $AB_5$
  - $90^\circ$/$120^\circ$ bond angle
  - 5 atoms attached
  - 0 lone pairs

Common Molecular Shapes

- **SQUARE PYRAMIDAL**: ClF$_5$
  - Type: $AB_5E$
  - $90^\circ$/$180^\circ$ bond angle
  - 5 atoms attached
  - 1 lone pair

Common Molecular Shapes

- **OCTAHEDRAL**: SF$_6$
  - Type: $AB_6$
  - $90^\circ$ 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)
**Orbital Hybridization Theory**

*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*

**Orbital Hybridization Theory**

• Let's look at CARBON:

\[ \text{C} + \text{H}_2 \rightarrow \text{CH}_3 \]

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

\[ \text{C} + \text{H}_2 \rightarrow \text{CH}_3 \]

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

- 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 \( \text{sp}^3 \) hybridized orbitals

\[ \text{C} + \text{H}_2 \rightarrow \text{CH}_3 \]

New orbitals have slightly *MORE* energy than the 2s and slightly *LESS* energy than the 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: BeF$_2$

Shape Hybridizations

- LINEAR (2 attached): sp
- TRIGONAL PLANAR (3 attached): sp$^2$
- TETRAHEDRAL (4 attached): sp$^3$
- TRIGONAL BIPYRAMIDAL (5 attached): sp$^3$d
- OCTAHEDRAL (6 attached): sp$^3$d$^2$
  
  *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)
  - Single bonds
  - sp
  - sp$^2$
  - sp$^3$

- Double Bonds
  - sp$^2$

- Triple Bonds
  - sp

Sigma ($\sigma$) Bonds

Electrons are located around the central axis of a bond

- Single bonds
- Ex: H$_2$

Pi ($\pi$) 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: O$_2$ / N$_2$