Chapter 8: Bonding – General Concepts

8.1 Chemical Bond Formation
8.2 Covalent Bonding (Lewis Dot Structures)
8.3 Charge Distribution in Covalent Compounds
8.4 Resonance
8.5 Molecular Shapes (VSEPR)
8.7 Molecular Polarity
8.9 Bond Properties
Why Study Molecular Structure?

It can help us understand molecular behavior and physical properties in the macroscopic world.

Consider Two Molecules of Formula $\text{AB}_2$:

$\text{CO}_2$ & $\text{H}_2\text{O}$

Based on Molar Mass, which would be easier to vaporize?

But water is a liquid and $\text{CO}_2$ is a gas.
Valence Electrons

Valence Electrons are in outer shell and involved in bonding interactions

Core electrons or inner shell electrons do not become involved with bonding interactions and stay associated with the parent nucleus

Elements of the same group (family) have similar chemical properties because of isoelectronic valence shell electron configurations
## Representative Elements & Valence Shell Electron Configurations

<table>
<thead>
<tr>
<th>Group</th>
<th>Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>1A</td>
<td>ns$^1$</td>
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<tr>
<td>2A</td>
<td>ns$^2$</td>
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<tr>
<td>3A</td>
<td>ns$^2$np$^1$</td>
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<tr>
<td>4A</td>
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<tr>
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<td>ns$^2$np$^3$</td>
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<tr>
<td>6A</td>
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<tr>
<td>7A</td>
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</tr>
<tr>
<td>8A</td>
<td>ns$^2$np$^6$</td>
</tr>
</tbody>
</table>
Group 1A: \((ns^1)\)  \(\overset{1}{A}\)°

Group 2A: \((ns^2)\)  \(\overset{2}{A}\)°

Group 3A: \((ns^2np^1)\)  \(\overset{2}{A}\)°

Group 4A: \((ns^2np^2)\)  \(\overset{2}{\overset{2}{A}}\)•
Representative Elements & Lewis Dot Structures

Group 5A: \((ns^2np^3)\)

Group 6A: \((ns^2np^4)\)

Group 7A: \((ns^2np^5)\)

Group 8A: \((ns^2np^6)\)
8.1 Chemical Bonds

1. **Ionic Bonds** - Ionic compounds between metals and nonmetals consisting of [+] cations and [-] anions.

2. **Covalent Bonds** - Classical “molecules” where valence electrons are shared between two nonmetals.

3. **Polar Covalent Bonds** - Covalent bonds with ionic character in that the electrons are not equally shared.

4. **Metallic Bonds** - Pure metals and alloys where delocalized free electrons hold together the positive nuclei.
How do We Distinguish between Ionic, Polar Covalent and Covalent Bonds?
Electronegativity

**Electronegativity** - “the relative ability of a bonding atom to attract electrons to itself.”
Electronegativity

Least

Li 1.0
Mg 1.2
Na 0.9
K 0.8
Rb 0.8
Cs 0.7
H 2.1
Be 1.5
Ca 1.0
Sc 1.3
Ti 1.5
Zr 1.4
Hf 1.3
Ta 1.5
V 1.6
Cr 1.6
Mn 1.5
Mo 1.8
W 1.7
Re 1.9
Fe 1.8
Co 1.9
Ni 1.9
Cu 1.9
Zn 1.6
Ga 1.6
Ge 1.8
As 2.0
Se 2.4
Br 2.8
Cl 3.0
N 3.0
O 3.5
F 4.0

Most

He
Ne
Ar
Kr
Xe
Rn
## Delta Electronegativities

<table>
<thead>
<tr>
<th>Bond Type</th>
<th>$\Delta EN$</th>
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</thead>
<tbody>
<tr>
<td>Covalent</td>
<td>0.0 - 0.5</td>
</tr>
<tr>
<td>Polar Covalent</td>
<td>0.5 - 2.0</td>
</tr>
<tr>
<td>Ionic</td>
<td>4.0 - 2.0</td>
</tr>
</tbody>
</table>
8.2 Covalent Bonds

Consider bonds to be result of atoms sharing electron pairs

2 Types of Electron Pairs
- lone pairs
- bonding pairs
Lewis Dot Structures

- Uses dots to represent valence electrons
- Ignores core electrons since they are not involved in bonding
- Usually use 4 orbitals (2 e\(^-\) each) for a max of 8 e\(^-\) (octet rule), note there are important exceptions to the octet rule

\[
\text{Sodium} = \text{Na} \cdot \\
\text{Sodium ion} = \text{Na}^+ \\
\text{Chlorine} = \cdot \ddot{\text{Cl}} : \\
\text{Chloride} = [ \cdot \ddot{\text{Cl}} : ]^-
\]
Exceptions to the Octet Rule

1. Atoms with odd # of electrons (free radicals)

2. Atoms which can have less than an octet
   - Hydrogen can only have 2 e⁻
   - Beryllium can have 4 valence e⁻
   - Aluminum can have 6 valence e⁻

3. Atoms which can have expanded octets
   Any atom of the 3rd or higher period has d orbitals which can allow for more than 8 valence electrons
Covalent Bonding - Diatomics

1. Hydrogen - Always has a **Duet**
   - Is an exception to the Octet Rule

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

or \( \text{H} - \text{H} \)

2. Chlorine: (Halogens)

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

or \( \text{:Cl} - \text{Cl}: \)
Covalent Bonding - Diatomics

3. **Oxygen** - has a **Double Bond**

\[
\text{O} + \text{O} \rightarrow \text{O} = \text{O}
\]

4. **Nitrogen** - has a **Triple Bond**

\[
\text{N} + \text{N} \rightarrow \text{N} \equiv \text{N}
\]
Lewis Dot Structures - Polyatomics

- Represent molecular orbitals with bonding and lone pair electrons
- Accounts for all valence electrons contributed to the molecule by each of its atoms
- Resonance structures arise when there is more than one way to arrange the electrons
- Formal Charge correlates a Lewis Dot Structures electron locations to number of valence electrons an atom donates to a molecule
Lewis Dot Structures: Rules

- These “Rules” can serve as a guide in determining the correct Lewis dot structure and will work for all compounds given in this class.

- You need to practice this.
1. Determine Total Number of Available Electrons.

- Each atom of the representative elements donates the number of electrons equivalent to group number (IA – VIIA).

- Sum all available electrons, the molecule (Lewis dot structure) must have this number of electrons.

- If there is an odd number of electrons, the lone electron exists in a nonbonding orbital or a multiple bond.

- If the compound is an ion, add one electron for each negative charge and subtract one electron for each positive charge.
2. Determine Structural Connectivity.

• First atom is central atom (unless Hydrogen, which can never be central)

• If more than 2 types of atoms, set up in order of formula (Acids are a notable exception).

• Organic compounds often contain multiple carbons connected to each other, you may need to be able to identify functional groups.

• Draw a line between each bonded atom (representing two electrons in a covalent bond)
3. Draw Hypothetical structure

• Draw hypothetical structure with each atom surrounded by 8 electrons in pairs.

• For each atom not surrounded by an octet of electrons (4 bonds), add lone pairs.

• Never give hydrogen an octet, its 1s valence shell can only contain 2 electrons.
4. Compare Electrons in Hypothetical Structure to the Available Number

- Count electrons in hypothetical structure.
- Compare to the available number of electrons (from step 1).

Three Possibilities

i. Hypothetical structure has same number of electrons: Your structure is correct

ii. Hypothetical structure has more electrons than are available: Go to step 5.

iii. Hypothetical structure has less electrons than it needs: Go to step 7.
5. Can Central Atom Have Less Than an Octet?

- Hydrogen always has a duet (one bond – 2 electrons)
- Group IIA can have a quartet (4 electrons)
- Group IIIA can have a sexter (6 electrons)

If not, go to step 6 and make multiple bonds
6. Can Central Atom Have Less Than an Octet?

- Place multiple bonds into structure by replacing two loan pairs from two adjacent atoms with a bonding pair.

- Repeat this until the structure has the same number of electrons as determined in step 1.

- Note, resonance structures arise if you have more than one choice in making the multiple bond.
7. Can Central Atom Have Expanded Octet?

- Atoms with d orbitals in their valence shell (3\textsuperscript{rd} period or greater) can have an “expanded octet”, that is more than 8 valence electrons.

- Add lone pairs to these atoms until the number of electrons equals the number from step 1.
Class Problems

1. Ammonia
2. Ammonium Ion
3. Methane (\text{CH}_4)
4. Sulfite
Draw Lewis Dot Structures for:

1. NO₂

2. BCl₃

3. ClF₃

4. BrF₅
Lewis Structure - Carbon Dioxide

# Valence e\(^{-}\) = 4 + 2(6) = 16

This figure has 20 e\(^{-}\), we only have 16, What can we do?

Note, if we replace 2 adjacent “lone pairs” with a bonding pair, we have reduced the total # of e\(^{-}\) by 2 (while still maintaining an octet), so CO\(_2\) needs 2 multiple bonds.
Lewis Structure - Carbon Dioxide

Note, sometimes there are several ways we can add multiple bonds

\( \ddot{\text{O}} = \text{C} = \ddot{\text{O}} \)

\( \ddot{\text{O}} - \text{C} \equiv \ddot{\text{O}} \)

\( \dddot{\text{O}} \equiv \text{C} - \dddot{\text{O}} \)

Experiments show us that the carbon-oxygen bonds in CO\(_2\) are identical, so the first structure is the best one.
Resonance Structures - 2 or More Lewis Dot Structures That Satisfy the Octet Rule.

Consider Ozone (O₃)

Neither Structure Is Correct As All O-O Bonds Are Identical, the True Structure Is an Average of the Resonance Structures
Group Problems

Draw Sulfur Trioxide, Include All Resonance Structures

Compare these to sulfite
8.3 Formal Charge

How do we determine which structure is best if more than 1 Lewis structure is possible?

**Formal Charge:**

Way of determining if Lewis dot structure places electrons near the atom which donates those electrons.
Formal Charge

FC = Group # - (Number of lone pair electrons + 1/2 Number of bonding pair electrons)

Look at the Formal Charges of each atom
-The best Lewis structure usually minimizes the formal charge of each atom
-If structures are assigned nonzero formal charges, the more electronegative atom should be assigned the negative formal charge
-Polyatomic ions have atoms with non-zero formal charges
Group Problem

Write the Lewis Dot Structure for a Molecule of Hydrogen, Nitrogen and Carbon

\[
\begin{align*}
\text{H} & \equiv \text{C} & \equiv & \equiv \text{N} \\
\text{H} & \equiv & \equiv \text{N} & \equiv \equiv \text{C}
\end{align*}
\]

Are These Resonance Structures?
**Group Problem**

Which Is Best?

![Structural Isomers](https://example.com/structure.png)

- These are structural isomers.
- They are isolatable and different chemical species (resonance structures are not “real” isolatable species)
- Use two arrows for isomers
- Use a double headed arrow for resonance structures

**NO**
\[ \text{H} - \text{C} \equiv \text{N} ! \]

\[ \text{FC}_\text{H} = 1 - [0 + 1/2(2)] = 0 \]
\[ \text{FC}_\text{C} = 4 - [0 + 1/2(8)] = 0 \]
\[ \text{FC}_\text{N} = 5 - [1 + 1/2(8)] = 0 \]

\[ \text{H} - \text{N} \equiv \text{C} ! \]

\[ \text{FC}_\text{H} = 1 - [0 + 1/2(2)] = 0 \]
\[ \text{FC}_\text{C} = 4 - [1 + 1/2(8)] = -1 \]
\[ \text{FC}_\text{N} = 5 - [0 + 1/2(8)] = 1 \]
Consider Resonance Structures of \( \text{NO}_2 \)

\[
\text{\begin{tabular}{c}
\text{\(\text{O} = \text{N} - \text{O}\) :} \\
\text{\(\text{O} = \text{\(\text{\(N\) - \text{O}}\)}\) .}
\end{tabular}}
\]

Which is Best?
Shortcomings of Lewis Structures

1. Resonance Structures

\[ \overset{\text{O}}{=\text{O}} - \overset{\text{O}}{\text{O}} \leftrightarrow \overset{\text{O}}{\text{O}} - \overset{\text{O}}{\text{O}} = \overset{\text{O}}{\text{O}} \]

Incorrectly Predicts Different Bond Lengths

2. Molecular Geometry

Lewis Dot Structures Can Not Predict Molecular Geometries

\[ \overset{\text{O}}{\text{O}} = \overset{\text{C}}{\text{C}} = \overset{\text{O}}{\text{O}} \quad \text{H} - \overset{\text{O}}{\text{O}} - \text{H} \]
How Do We Determine Molecular Geometries?

Two Step Process

1. Determine Valence Electron Orbital Geometry
   - Note Two Types of Valence Orbitals
     A. Bonding Orbitals
     B. Lone Pair Orbitals

2. Determine Molecular Geometry Based only on Bonding Orbitals
What Theory Describes How Negatively Charged Electron Orbitals Orient Around a Positively Charged Nucleus?
8.5 VSEPR

Valence Shell Electron Pair Repulsion Theory

-Electrons in Valence Orbitals Interact to Minimize -/- Repulsion and Maximize +/- Attraction

Most Stable Geometry Maximizes Distance Between Orbitals While Maintaining Constant Distance From the Nucleus
Steric Number (SN)

SN = the Number of Electron Domains

Different Electronic Geometries Result from Different Steric Numbers

Note: Multiple Bonds are Considered a Single Electron Domain
VSEPR Geometries
Two Electron Pairs (SN = 2)

- Linear
- $180^\circ$ angle
VSEPR Geometries
Three Electron Pairs (SN = 3)

- Trigonal Planar
- 120° angle
VSEPR Geometries
4 Electron Pairs (SN = 4)

- Tetrahedral
- 109.5° angle
VSEPR Geometries

- Trigonal
- Bipyramidal

5 Electron Pairs (SN=5)

- 90° axial
- 120° equatorial
VSEPR Geometries

5 Electron Pairs (SN=5)

- Trigonal Bipyramidal
- 90° axial
- 120° equatorial
VSEPR Geometries

6 Electron Pairs  (SN=6)

- Octahedral

- 90°
Prediction of Molecular Geometries

1. Use VSEPR to Identify Electronic Structure Based on 2-6 Electron Pairs From the Lewis Dot Structure.

-Multiple Bonds Count as a Single Electron Pair

2. Visualize the Molecular Structure From the Bonding Orbitals of the Electronic Structure
Two Electron Pairs

(Linear Electronic Geometry)

A. 2 bonding, 0 nonbonding orbitals

BeH₂
Two Electron Pairs

(Linear Electronic Geometry)

A. 2 bonding, 0 nonbonding orbitals

BeH$_2$
Three Electron Pairs

(Trigonal Planar Electronic Geometry)

A. 3 bonding, 0 nonbonding orbitals
B. 2 bonding, 1 nonbonding orbitals
Three Electron Pairs
(Trigonal Planar Electronic Geometry)

A. 3 bonding, 0 nonbonding orbitals

$\text{BF}_3$
Three Electron Pairs
(Trigonal Planar Electronic Geometry)

A. 3 bonding, 0 nonbonding orbitals

\[
\text{BF}_3
\]

Trigonal Planar Geometry

120° F-B-F bond angle
Three Electron Pairs
(Trigonal Planar Electronic Geometry)

B. 2 bonding, 1 nonbonding orbitals

CH₂
Three Electron Pairs
(Trigonal Planar Electronic Geometry)

B. 2 bonding, 1 nonbonding orbitals

\( \text{CH}_2 \)

Bent Geometry

120° H-C-H angle
Four Electron Pairs
(Tetrahedral Electronic Geometry)

A. 4 bonding, 0 nonbonding orbitals
B. 3 bonding, 1 nonbonding orbitals
C. 2 bonding, 2 nonbonding orbitals
Four Electron Pairs
(Tetrahedral Electronic Geometry)

A. 4 bonding, 0 nonbonding orbitals

METHANE CH₄
Four Electron Pairs
(Tetrahedral Electronic Geometry)

A. 4 bonding, 0 nonbonding orbitals

METHANE CH$_4$

-Tetrahedral Geometry
-109.5° H-C-H angle
Four Electron Pairs
(Tetrahedral Electronic Geometry)

B. 3 bonding, 1 nonbonding orbitals

Ammonia  \( \text{NH}_3 \)
Four Electron Pairs
(Tetrahedral Electronic Geometry)

B. 3 bonding, 1 nonbonding orbitals

Ammonia $\text{NH}_3$

- Trigonal Pyramidal geometry
- $107^\circ \text{H-N-H bond angle}$
- Lone Pair Takes up More Space than Bonding Pairs
Four Electron Pairs
(Tetrahedral Electronic Geometry)

c. 2 bonding, 2 nonbonding orbitals

H₂O
Four Electron Pairs
(Tetrahedral Electronic Geometry)

c. 2 bonding, 2 nonbonding orbitals

$\text{H}_2\text{O}$

Bent Geometry

$104.5^\circ \text{ H-O-H bond angle}$
Five Electron Pairs (SN=5)
(Trigonal BiPyramidal electronic geometry)

A. 5 bonding, 0 nonbonding orbitals
B. 4 bonding, 1 nonbonding orbitals
C. 3 bonding, 2 nonbonding orbitals
D. 2 bonding, 3 nonbonding orbitals
Five Electron Pairs
(Trigonal Bipyramidal Electronic Geometry)

A. 5 bonding, 0 nonbonding orbitals
Five Electron Pairs
(Trigonal Bipyramidal electronic geometry)

A. 5 bonding, 0 nonbonding orbitals
Five Electron Pairs
(Trigonal Bipyramidal electronic geometry)

B. 4 bonding, 1 nonbonding orbitals

- Lone Pair takes equatorial Position

See-Saw Geometry or Distorted Tetrahedron

\[
\text{SF}_4
\]
Five Electron Pairs
(Trigonal Bipyramidal Electronic Geometry)

C. 3 bonding, 2 nonbonding orbitals
Five Electron Pairs
(Trigonal BiPyramidal Electronic Geometry)

C. 3 bonding, 2 nonbonding orbitals

BrF$_3$

T-Shaped
Five Electron Pairs
(Trigonal BiPyramidal electronic geometry)

D. 2 bonding, 3 nonbonding orbitals

$\text{XeF}_2$

Linear
Six Electron Pairs

(Octahedral Electronic Geometry)

A. 6 bonding, 0 nonbonding orbitals
B. 5 bonding, 1 nonbonding orbitals
C. 4 bonding, 2 nonbonding orbitals
Six Electron Pairs
(Octahedral Electronic Geometry)

A. 6 bonding, 0 nonbonding orbitals

$\text{SF}_6$
Six Electron Pairs
(Octahedral electronic Geometry)

A. 6 bonding, 0 nonbonding orbitals

\( \text{SF}_6 \)

Octahedral Geometry

90° F-S-F bond angle
Six Electron Pairs
(Octahedral electronic geometry)

B. 5 bonding, 1 nonbonding orbitals

\[ \text{IF}_5 \]

Square Pyramidal Geometry
Six Electron Pairs

(Octahedral electronic geometry)

C. 4 bonding, 2 nonbonding orbitals

XeF$_4$

Square Planar Geometry
Geometry of Organic Compounds

VSEPR Determines the Geometry of Molecules With More Than One Central Atom

What is the geometry of each carbon in the following compound?

\[ \text{HC} \equiv \text{CCH}_2\text{COOH} \]
8.7 Polarity of Polyatomic Molecules

**Polar Bond** - Covalent Bond Between 2 Atoms of Different Electronegativities

**Bond Dipole** - Measure of Bond Polarity

*How do we Determine if a Molecule is Polar?*
The Molecular Dipole Is the Vectoral Sum of All the Individual Bond Dipoles Present in a Molecule.

Therefore We Need to Know:

1. The Number and Types of Bonds
2. The Polarity of Each Type of Bond
3. The Orientation of Each Bond
Is Carbon Dioxide Polar?

\[ \text{NO!} \]

The C-O bond dipoles are equal and cancel each other
Molecular Symmetry and Polarity

Is Water Polar?
Non Polar Molecules with Polar Bonds

Symmetric cancellation of Bond Dipoles
Bond Polarity & Dipole Moment

What types of geometries would have cancellation of bond dipoles?
What Molecular Geometric Structures are Nonpolar

A-central atom, B,C,D- atoms, N-nonbonding electron pairs

1. Linear(l): $\text{AB}_2$, $\text{AB}_2\text{N}_3$
2. Trigonal Planar(tp): $\text{AB}_3$
3. Tetrahedral: $\text{AB}_4$
4. Square Planar(sp): $\text{AB}_4\text{N}_2$, $\text{AB}_2\text{C}_2\text{N}_2$
5. Trigonal Bipyramidal: $\text{AB}_5$, $\text{AB}_2\text{C}_3$
6. Octahedral: $\text{AB}_6$, $\text{AB}_4\text{C}_2$, $\text{AB}_2\text{C}_2\text{D}_2$
Linear Cancellation of Bond Dipoles
Trigonal Planar Cancellation of Bond Dipoles
Tetrahedral Cancellation of Bond Dipoles
What Molecular Geometric Structures are Nonpolar

A-central atom, B,C,D- atoms, N-nonbonding electron pairs

1. Linear(l): $AB_2$, $AB_2N_3$
2. Trigonal Planar(tp): $AB_3$,
3. Tetrahedral: $AB_4$
4. Square Planar(sp): $AB_4N_2$, $AB_2C_2N_2$
5. Trigonal Bipyramidal: $AB_5$, $AB_2C_3$
6. Octahedral: $AB_6$, $AB_4C_2$, $AB_2C_2D_2$
8.9: Covalent Bonds
Ionic Bonds

Coulomb’s Law

\[ E = \kappa \frac{Q_1 Q_2}{d} \]

- Charge of anion
- Charge of cation
- Distance between nuclei

\[ k = 2.31 \times 10^{-19} \text{ J-nm} \]
Crystal Lattice Energies

\[ E = \kappa \frac{z_1 z_2}{d} \] (kJ/mol)

<table>
<thead>
<tr>
<th></th>
<th>F^-</th>
<th>Cl^-</th>
<th>Br^-</th>
<th>O^-2</th>
</tr>
</thead>
<tbody>
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<td>Li^+</td>
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<td>-807</td>
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<td>-5492</td>
<td>-5361</td>
<td>-15916</td>
</tr>
</tbody>
</table>
Covalent Bonds
Polar Bonds

$\delta^+ \quad \delta^-$

H  (EN=2.1)  Cl  (EN=3.0)

$\Delta$EN=.9
# Strength of Covalent Bonds

## Bond Dissociation Energies (kJ/mol)

<table>
<thead>
<tr>
<th>Bond</th>
<th>Energy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C=O</td>
<td>732</td>
</tr>
<tr>
<td>O=O</td>
<td>498</td>
</tr>
<tr>
<td>H-H</td>
<td>436</td>
</tr>
<tr>
<td>C-F</td>
<td>450</td>
</tr>
<tr>
<td>H-C</td>
<td>410</td>
</tr>
<tr>
<td>C-Cl</td>
<td>330</td>
</tr>
<tr>
<td>C-C</td>
<td>350</td>
</tr>
<tr>
<td>C-Br</td>
<td>270</td>
</tr>
<tr>
<td>C=C</td>
<td>611</td>
</tr>
<tr>
<td>C-I</td>
<td>240</td>
</tr>
<tr>
<td>C=C</td>
<td>835</td>
</tr>
<tr>
<td>H-O</td>
<td>460</td>
</tr>
</tbody>
</table>

- Energy Required to Split the Atoms Apart
## Strength of Covalent Bonds

The Greater the Polarity the Stronger the Bond of Similar Bonds

(kJ/mol)

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<td>240</td>
</tr>
<tr>
<td>H-O</td>
<td>460</td>
</tr>
</tbody>
</table>

**Strength Order:** C-F > C-Cl > C-Br > C-I
### Strength of Covalent Bonds

Multiple Bonds are Stronger Than Single Bonds

(kJ/mol)

<table>
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<td>C-I</td>
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</tr>
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<td>C≡C</td>
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</tr>
<tr>
<td>H-O</td>
<td>460</td>
</tr>
</tbody>
</table>

C≡C > C=C > C-C
Bond Order

The higher the bond order, the stronger the bond

\[ \text{BO:} \quad (3) > (2) > (1) \]

\[ \text{C≡C > C=C > C-C} \]
## Bond Energy & Bond Length

For similar types of bonds, the greater the Energy, the shorter the bond strength:

<table>
<thead>
<tr>
<th>bond</th>
<th>Energy (kJ/mol)</th>
<th>length (nm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C-C</td>
<td>350</td>
<td>14.3</td>
</tr>
<tr>
<td>C=C</td>
<td>611</td>
<td>12.3</td>
</tr>
<tr>
<td>C=C</td>
<td>835</td>
<td>11.3</td>
</tr>
</tbody>
</table>
Reaction Enthalpies from Bond Dissociation Energies

\[ \Delta H_{rxn} \text{ can be determine from bond enthalpies} \]

\[ \Delta H_{rxn} = \sum \text{(bond dissociation energies of bonds broken)} - \sum \text{(bond dissociation energies of bonds formed)} \]
Calculate $\Delta H_{\text{rxn}}$ for the combustion of methane ($\text{CH}_4$)

$$\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$$

$$\Delta H_{\text{rxn}} = 4D_{\text{C-H}} + 2D_{\text{O=O}} - \{2D_{\text{C=O}} + 4D_{\text{H-O}}\}$$

$$\Delta H_{\text{rxn}} = 4(410) + 2(498) - \{2(732) + 4(460)\}$$

$$= -668 \text{ kJ/mol}$$