Today's Lecture Topics

- Lewis Structures
- Formal Charges
- Resonance Structures
- Shapes of Molecules
- Valence Bond Theory/ Hybridization

Review sessions: Monday & Thursday 3:30pm

Drawing Lewis Structures

1. You must know how the atoms are connected
2. Count up the total valence e-‘s.
3. Connect the bonded atoms by placing an e- pair between each pair of bonded atoms
4. Place additional e- pairs around atoms as non-bonding pairs to achieve an octet
5. Move non-bonded e- pairs into bonding positions to form multiple bonds as needed (if necessary) to achieve octet

ex: CH₄

1. C is central atom
2. C: 1×4e- = 4
   H: 4×1e- = 4
   8e- total

ex: H₂O

O: 1×6e- = 6
H: 2×1e- = 2
8e- total

ex: CO₂

We know O-C-O is arrangement of atoms

C: 1×4e- = 4
O: 2×6e- = 12
16e- total

Octet rule not satisfied for C
Assigning formal charges
- If atom has same # of e−s as it came in with, it is neutral.

Ex: H₂O

O − 6e−
Each H − 1e+

H − O − H

O has 6 e−, so neutral
H has one e−, so neutral

Ex: Diazomethane


N came in with 5 valence e−s

Octet rule isn't always followed:

Ex: B₃H₃

B 1x3e− − 3e−
H 3x1e− 13e−

6e− total, not 8! (electron deficient)

Carbon can NEVER have more than 8 e−.
Greater than 8 e− can only happen for 3rd row or higher elements.

Ex: PCl₅

P − 1x5e− − 5
Cl 5x7e− 35

40e− total

Resonance Structures - Very important concept in organic chemistry!

Ex: Diazomethane


In reality, neither form actually exists. Diazomethane looks like a hybrid of these two forms:

Electrons are delocalized
Acetic acid:

\[
\begin{align*}
\text{CH}_3\text{C}=\text{O} & \leftrightarrow \text{H}_3\text{C}-\text{C}=\text{O} \\
\text{H}_3\text{C}-\text{C}=\text{O} & \leftrightarrow \text{H}_3\text{C} - \text{O} \\
\text{H}_3\text{C} - \text{O} & \rightarrow \text{H}_3\text{C} - \text{O}
\end{align*}
\]

Hints for drawing Resonance Structures:
1) Obey normal rules of valence.
2) Resonance forms only differ in the positions of multiple bonds and non-bonding e−s (atoms and single bonds don't move).

**Shapes of Molecules**

Methane, CH₄ \( \text{H}-\text{C}-\text{H} \) is not actually planar. In 3-D, it is a perfect Tetrahedral (pyramid w/ Δ base).

VSEPR - electrons repel each other, so they try to be as far away from each other as possible.

For CH₄, tetrahedral is shape where e− are furthest apart.

**Valence Shell Electron Pair Repulsion Theory**

1) Draw Lewis Structure
2) Count total e− pairs around central atom (Ignore multiple bonded pairs)
3) Determine geometry of e− pairs based on VSEPR
4) Ignore non-bonded pairs to predict shape of molecule

**Example:**

H−O−H → 4 pairs of e− \( \leftrightarrow \) Td (tetrahedral) → H \( \text{O} \) \( \text{H} \)

<table>
<thead>
<tr>
<th># of atoms bonded to central atom</th>
<th>Shape</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>Linear</td>
</tr>
<tr>
<td>3</td>
<td>Trigonal planar</td>
</tr>
<tr>
<td>4</td>
<td>Tetrahedral (Td)</td>
</tr>
<tr>
<td>5</td>
<td>Trigonal bi-pyramidal</td>
</tr>
<tr>
<td>6</td>
<td>Octahedral</td>
</tr>
</tbody>
</table>