

Helpful Hints for Drawing Covalent Lewis Dot Structures

- H is always a terminal atom → ALWAYS connected to only 1 other atom.
- Lowest electronegativity is central atom in molecule.
- If drawing the Lewis structure for a polyatomic ion,
 - For positive ions, subtract electron(s) from the central atom.
 - For negative ions, add electron(s) to the central atom.
 - Enclose the dot structure in square brackets and include the ion's charge outside the brackets.
- Not all elements can form double or triple bonds: **only C, N, O, P, and S!** (Think CNOP-S)

Exceptions to the octet rule

- Elements that will have less than 8 valence electrons and are stable.
 - Hydrogen, 2 electrons (1 bond)
 - Beryllium, 4 electrons (2 bonds)
 - Boron, 6 electrons (3 bond)
- Elements that will have more than 8 valence electrons and are stable. ⇒ expanded octet!
 - Elements in period (row) 3 through 7 can often expand their octet and can form more than 4 bonds (can have up to 12 electrons, 6 bonds)
 - This is only possible between periods 3 through 7 because they can hold electrons in their empty d sublevel.
 - If you are unsure where to put extra lone pairs, check to see if the central atom can have an expanded octet (check to see if the element is in periods 3 through 7)





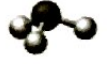

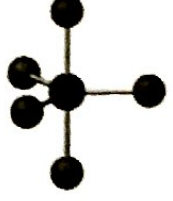
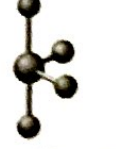





Sigma (σ) and Pi (π) Bonding

Sigma (σ) bond: covalent bond formed by orbitals overlapping end to end. The electron density is concentrated between the nuclei of the two atoms involved in a bond.

Pi (π) bond: covalent bond formed by orbitals overlapping side by side. The electron density is concentrated above and below the nuclei of the two atoms involved in a bond.

Bond Type	Made of	Length / Strength
single bond	<u>1</u> sigma bond	longest/ weakest
double bond	<u>1</u> sigma bond + <u>1</u> pi bond	medium length/ strength
triple bond	<u>1</u> sigma bond + <u>2</u> pi bonds	shortest/ strongest

VSEPR: Memorize Shapes and Bond Angles!

Electron Groups*	Bonding Groups	Lone Pairs	Electron Geometry	Molecular Geometry	Approximate Bond Angles	Example
2	2	0	Linear	Linear	180°	$\ddot{\text{O}}=\text{C}=\ddot{\text{O}}:$ 
3	3	0	Trigonal planar	Trigonal planar	120°	$\begin{array}{c} \text{:F:} \\ \\ \text{:F}-\text{B}-\text{F:} \\ \\ \text{:F:} \end{array}$ 
3	2	1	Trigonal planar	Bent	≈118°	$\ddot{\text{O}}=\ddot{\text{S}}-\ddot{\text{O}}:$ 
4	4	0	Tetrahedral	Tetrahedral	109.5°	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{H} \\ \\ \text{H} \end{array}$ 
4	3	1	Tetrahedral	Trigonal pyramidal	≈107°	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{N}-\text{H} \\ \\ \text{H} \end{array}$ 
4	2	2	Tetrahedral	Bent	≈105°	$\text{H}-\ddot{\text{O}}-\text{H}$ 
5	5	0	Trigonal bipyramidal	Trigonal bipyramidal	120° (equatorial) 90° (axial)	$\begin{array}{c} \text{:Cl:} \\ \\ \text{:Cl}-\text{P}-\text{Cl:} \\ \\ \text{:Cl:} \end{array}$ 
5	4	1	Trigonal bipyramidal	Seesaw	≈118° (equatorial) ≈88° (axial)	$\begin{array}{c} \text{:F:} \\ \\ \text{:F}-\text{S}-\text{F:} \\ \\ \text{:F:} \end{array}$ 
5	3	2	Trigonal bipyramidal	T-shaped	≈86°	$\begin{array}{c} \text{:F:} \\ \\ \text{:F}-\text{Br}-\text{F:} \\ \\ \text{:F:} \end{array}$ 
5	2	3	Trigonal bipyramidal	Linear	180°	$\text{:F}-\text{Xe}-\text{F:}$ 
6	6	0	Octahedral	Octahedral	90°	$\begin{array}{c} \text{:F:} \\ \\ \text{:F}-\text{S}-\text{F:} \\ \\ \text{:F:} \end{array}$ 
6	5	1	Octahedral	Square pyramidal	≈88°	$\begin{array}{c} \text{:F:} \\ \\ \text{:F}-\text{Br}-\text{F:} \\ \\ \text{:F:} \end{array}$ 
6	4	2	Octahedral	Square planar	90°	$\begin{array}{c} \text{:F:} \\ \\ \text{:F}-\text{Xe}-\text{F:} \\ \\ \text{:F:} \end{array}$ 

Hybridization

Hybrid Orbitals: orbitals of equal energy created by blending two or more valence orbitals on the same atom

- Hybridization can be determined by counting regions of electron density: electron domains!

# of Electron Domains	# of Hybrid Orbitals	Electron Geometry	Hybridization
2	2	Linear	sp
3	3	Trigonal planar	sp^2
4	4	Tetrahedral	sp^3
5	5	Trigonal bipyramidal	sp^3d
6	6	Octahedral	sp^3d^2

Not AP tested!

Isomers vs Resonance Hybrids

Isomers: molecules with the same molecular formula but different spatial arrangement of atoms

- Same # of atoms of each element
- Different arrangement of their atoms in space
- Different physical properties (boiling point, viscosity, etc) depending on their IMFs!

Resonance structures: when two or more Lewis structures can validly represent a molecule (or ion)

- Same position of atoms in space (+ same # of atoms)
- Different electron arrangement
- The actual structure, the resonance hybrid, is intermediate between the two or more resonance structures.
 - All possible dot structures contribute to the real structure, BUT more stable ones (↓ formal charge) contribute more.
 - Resonance often occurs in structures with a double or triple bond.

Formal Charge

Formal Charge: a way to identify the best Lewis dot structure when more than one valid dot structure exists

- Formal charges are hypothetical charges assigned to each element in the dot structure

Formal Charge = # of valence electrons - # non-bonding electrons (lone) - $\frac{1}{2}$ # bonding electrons

You do NOT need to show work for formal charge calculations!!! 😊

Formal Charge Rules (Which dot structure is best?)

- Small (or even better, 0) formal charges are more stable.
- formal charge on the more electronegative atoms.
- + formal charge on the less electronegative atoms.
- Sum of all formal charges must equal the charge of the molecule.