Dougherty Valley • AP Chemistry

Bonding, Molecular Structure & Hybridization

S-52

STUDY LIST From Paul Groves

Valence Electrons & Lewis Symbols

I can…

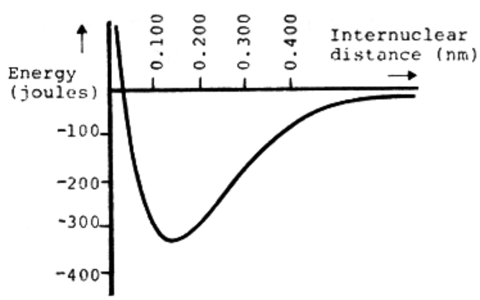
- ☐ State the number of valence electrons for any atom.
- ☐ Draw Lewis Dot Symbols for any atom or ion.
- ☐ Explain that families II, III, and IV have both a ground state and promoted state form of the Lewis symbol.
- ☐ Draw the Lewis symbol for a simple ion such as Na^+ or Cl^-.

Bonding

- ☐ State the type of bond (ionic, covalent, metallic) formed between any two atoms.

metal-metal	metallic bond
metal-nonmetal	ionic bond
nonmetal-nonmetal	covalent bond

- ☐ Explain (using attractions and repulsions) why the formation of a bond lowers the potential energy of a molecule.
- ☐ Use the following diagram to determine the bond length and bond energy of a bond.



- ☐ State that a covalent bond usually forms between two atoms with half-filled orbitals.

Lewis Dot Symbols

- ☐ Draw Lewis dot symbols to show a covalent bond between atoms in a molecule.

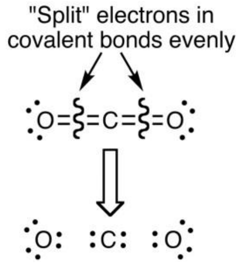
- ☐ Identify "lone pair" electrons vs. "shared pair" electrons in a Lewis structure.
- ☐ Draw molecules with double and triple bonds.
- ☐ Note that only C, N, O, and sometimes S form multiple bonds.
- ☐ Use Mr. Groves' method of "take away a pair, take away a pair, make these guys share" to draw molecules with multiple bonds while maintaining the octet of electrons for each atom.
- ☐ State that many atoms gain, lose, or share electrons until they are surrounded by eight electrons. This is called the "octet rule".
- ☐ Memorize the Lewis symbols for the seven diatomic molecules. N_2 has a triple bond. O_2 has a double bond.
- ☐ Draw examples of molecules that do not follow the octet rule because the atoms have **less** than an octet. (e.g., CaH_2, H_2, Families I, II, III)
- ☐ Draw Lewis symbols for polyatomic ions.
- ☐ Draw Lewis symbols for molecules and ions that exhibit resonance.
- ☐ Memorize some of the more common molecules and ions that exhibit resonance [e.g., NO_3^-, CO_3^{2-}, SO_2, NO_2, O_3, C_6H_6, $C_2H_3O_2^-$].
- ☐ Draw Lewis symbols for molecules and ions that violate the octet rule by using their "p" orbitals for extended valence shells. [e.g., SF_6, XeF_2, XeF_4, IBr_3, PCl_5]
- ☐ Explain why P can form PF_3 and PF_5, but N (same family) can form NF_3, but not NF_5.

Bond Energies

- ☐ Define bond energy.
- ☐ Write a chemical equation to show bond energy of any bond. For example, the F-F bond in F_2 is $F_2(g) + \text{energy} \rightarrow 2F(g)$
- ☐ Determine the bonds broken and bonds formed during a chemical reaction by drawing the Lewis structures of the reactants and products.
- ☐ Use a chart of bond energies to calculate the Enthalpy of a reaction (ΔH).
- ☐ Explain that this method does not give exactly the same answer as Hess's Law because bond energies are **average** bond energies that differ slightly from molecule to molecule.

Formal Charge & Oxidation Number

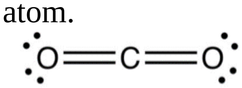
☐ Define formal charge as the charge on an atom if all shared electrons are shared equally.



☐ Determine the formal charge of any atom in a Lewis Structure and use these formal charges to determine the best arrangements of atoms.

☐ State that the best structures have minimal formal charges and the more electronegative atoms have the negative formal charges.

☐ Contrast formal charge with oxidation states in which shared electrons are assigned to the more electronegative atom.



	O	C	O
Formal Charge	0	0	0
Oxidation State	2–	4+	2–

Shapes of Molecules

☐ Define Steric Number (SN) as the # of bonded atoms plus the # of lone pairs on an atom.

☐ State the Steric Number (SN) of the central atom in any Lewis structure.

☐ Use VSEPR to state the shape and bond angle associated with each Steric Number.

2	linear	180°
3	trigonal planar	120°
4	tetrahedral	109.5°
5	trigonal bipyramidal	90° & 120°
6	octahedral	90°

☐ State the shape of a molecule (arrangement of the atoms). [AKA "Molecular Geometry"]

☐ State the type of orbital hybridization used with each steric number.

1	2	3	4	5	6
s	sp	sp^2	sp^3	sp^3d	sp^3d^2

☐ Explain that non-hybridized orbitals remain as p-orbitals. For example:

$$\mathbf{s} + \mathbf{p} + p + p \rightarrow \mathbf{sp} + \mathbf{sp} + p + p$$

$$\mathbf{s} + \mathbf{p} + \mathbf{p} + p \rightarrow \mathbf{sp^2} + \mathbf{sp^2} + \mathbf{sp^2} + p$$

Electronegativity

☐ Use the difference in electronegativity values (ΔEN) of any two atoms to classify the bond.

ionic	$\Delta EN > 1.7$
polar covalent	$0.5 \leq \Delta EN \leq 1.7$
nonpolar covalent	$\Delta EN < 0.5$

☐ State the positive and negative end of any polar bond.

☐ Judge from the molecular shape whether the molecule is polar if the bonds are polar.

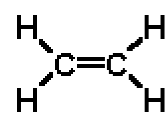
☐ State the electronegativity values for C, N, O, F, P, S, and Cl from their positions on the table.

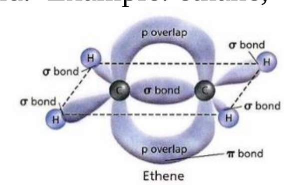
Multiple Bonds, Bond Order, & Resonance

☐ Define bond order as the number of pairs of electrons holding two atoms together in a covalent bond.

single bond	B.O. = 1	longer	weaker
double bond	B.O. = 2		
triple bond	B.O. = 3	shorter	stronger

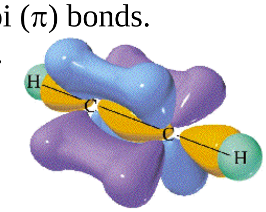
☐ Describe a double bond as an atom using sp^2 hybridization (SN=3) and utilizing the p-orbital to form a pi (π) bond. Example: ethane, C_2H_4.



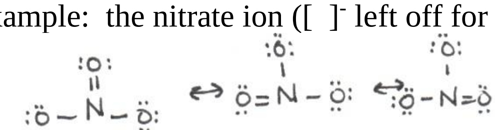


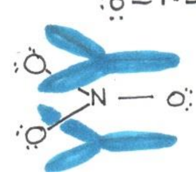
☐ Describe a triple bond as an atom using sp hybridization (SN=2) and utilizing the two p-orbitals to form two pi (π) bonds. Example: ethyne, C_2H_2.

H:C:::C:H



☐ Explain that when resonance occurs, each atom involved uses sp^2 hybrid orbitals and each of the p-orbitals blends into a pi bond. Example: the nitrate ion ([ ]$^-$ left off for clarity)





Notice that the bond order of the N-O bond is 1.33.
When two resonance structures are involved, the bond order is 1.5.