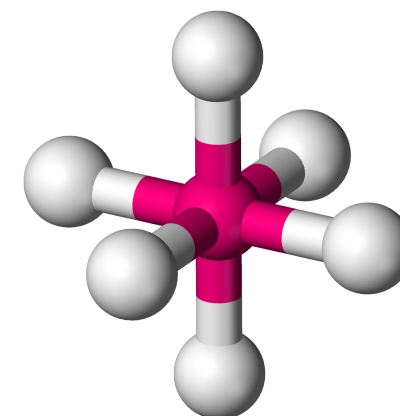


Unit 3 Exam Review

BONDING, PART ONE

How to succeed on this exam:

- **Conceptual Questions are important**
 - There is no great practice for conceptual questions: study the lecture material and reviews (use gchem as a supplement)
- **Practice A LOT:**
 - The Sapling homework has some really good problems that will prepare you for the exam
 - Canvas extra practice
 - Gchem > Bonding Unit > Additional Content > Helpsheets and Worksheets
 - Another good exercise is drawing out the Lewis Structures of ALL polyatomic ions:
 - Draw the structure (might have resonance!)
 - Determine the formal charges on the atoms for your acceptable structures
 - Determine the VSEPR geometry and bond angles
 - Check your work with wikipedia



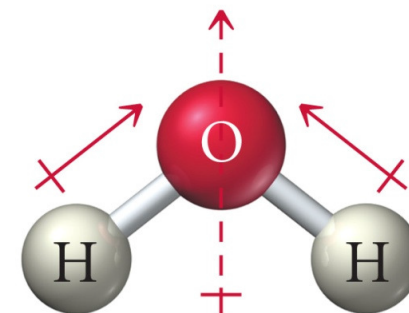
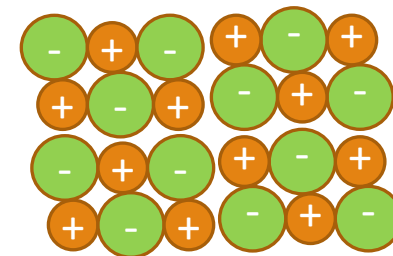
Don't waste
your time with:
Chromate,
dichromate,
permanganate,
or peroxide

Learning Objectives – Unit 3

1. Identify metals and non-metals, and predict the types of compounds (ionic/covalent) that will form from different elements.
2. Distinguish between molecules, ions, and atoms.
3. Predict the anion or cation that a main-group element is likely to form.
4. Relate Coulomb's law to ionic radii, ionic charge, and lattice energy.
5. Describe the distance dependence of the potential energy of a covalent bond.
6. Predict and explain relative bond strength and lengths in a compound.
7. Name and write formulas for covalent compounds.
8. Interpret line drawings of chemical compounds with implicit hydrogens, carbons, and lone pairs.
9. Rank the polarity of covalent bonds based on relative electronegativity.
10. Define dipole moment and identify polar bonds.
11. Draw the best Lewis structure (including any resonance structures) for a molecule or polyatomic ion.
12. Apply formal charges to structures and use them to predict the most likely structure.
13. Recognize and apply exceptions to the octet rule.
14. Apply the VSEPR model to determine a molecule's electronic and molecular geometry based on its Lewis dot structure.
15. Assess if a molecule is polar based on polar bonds and its molecular geometry.

Ionic vs. Covalent Bonds

- **Ionic bonds** are the electrostatic interaction between discrete fully positive and fully negative charges.
 - Ionic bonds are quantified by lattice energy
 - Ionic compounds are almost always a metal (+) and nonmetal (-)
 - The ions can be elements or polyatomic ions (which are, themselves, covalent compounds)
 - Results in an alternating lattice of positive and negative charges
- **Covalent bonds** are the sharing of electrons between two species
 - Covalent bonds are quantified by bond length, bond strength, bond order, and polarity ($\Delta\epsilon_n$)
 - Covalent bonds result from the electrostatic interaction between partial charges (δ^+ , δ^-)
 - Results in molecules



Lattice Energy

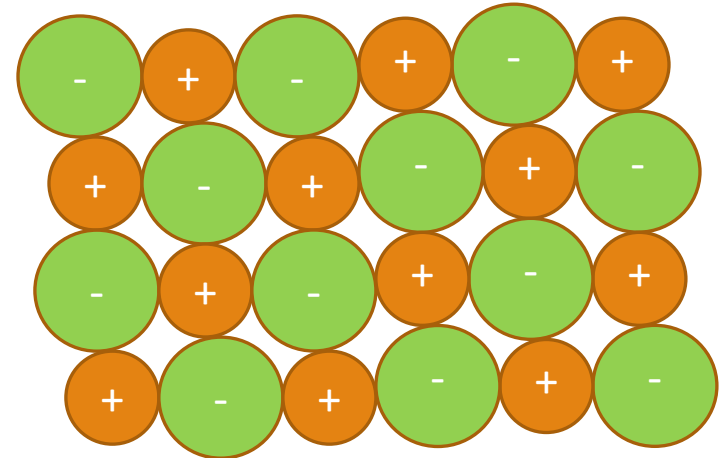
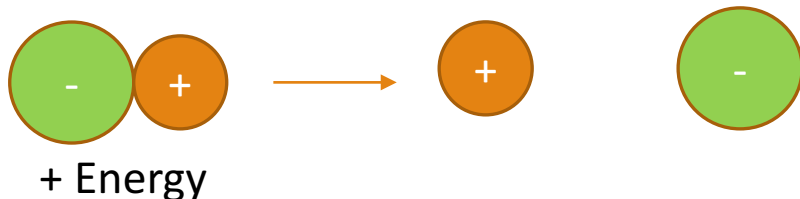
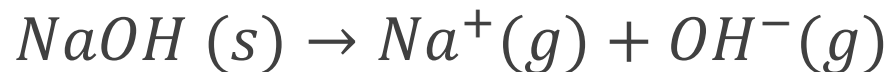
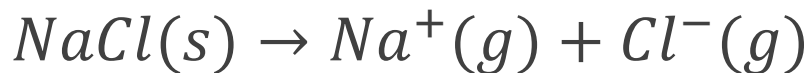
Is lattice energy positive or negative?

- Depends on how you're measuring it. But use the absolute value.
- **Lattice energy indicates the amount of energy necessary to overcome the negative potential energy binding the charges of an ionic compound**

- Lattice Energy: is directly proportional to the charges of the ions (q_1 and q_2) and inversely proportional to the size of the ions (r)

$$\Delta H_{lattice} \propto \frac{q_1 q_2}{r}$$

- This describes the energy required for the following reaction:

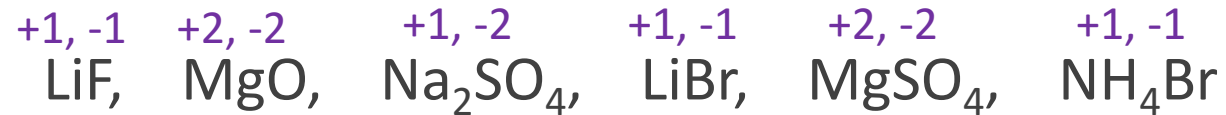


When Ranking Lattice Energies:

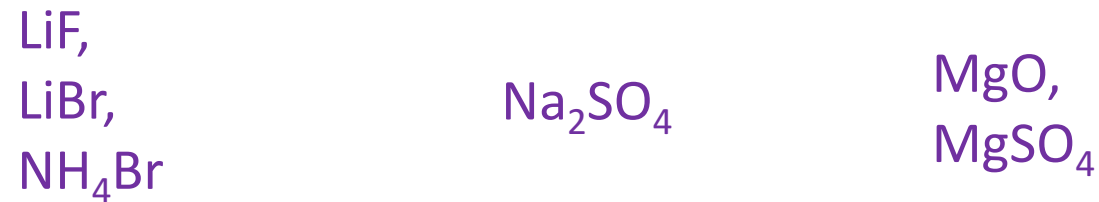
1. Prioritize **charge first**
2. If two ionic compounds have the same charge, the **smaller one will have a higher energy**
3. Polyatomic ions are **big**

Lattice Energy

List the following in terms of increasing lattice energy:



1. Rank based on charge:

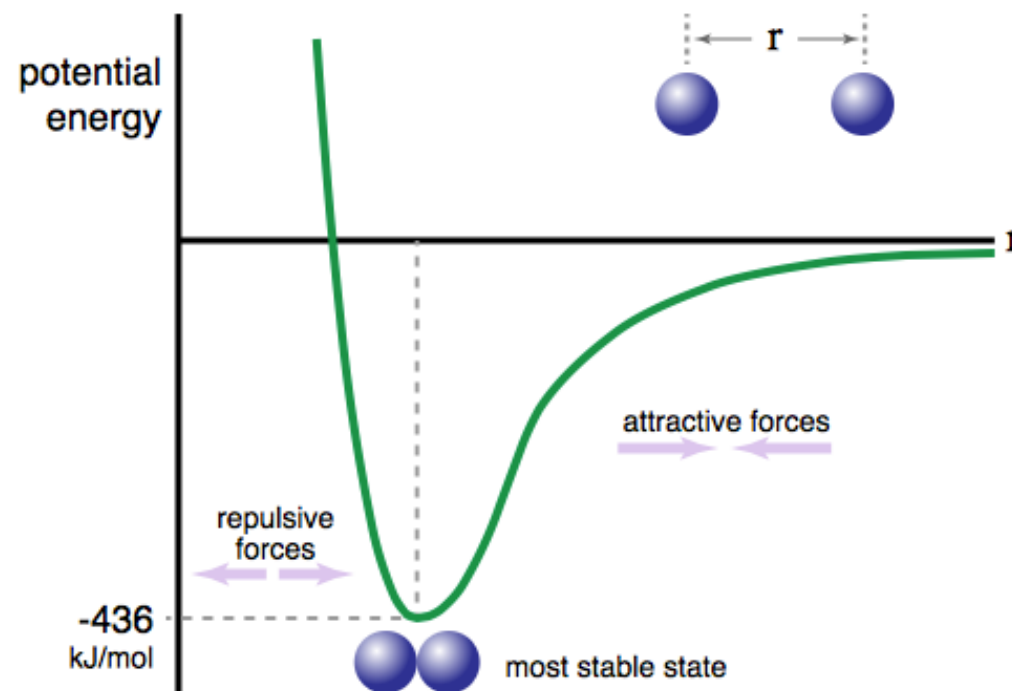


2. Settle everything else based on size:

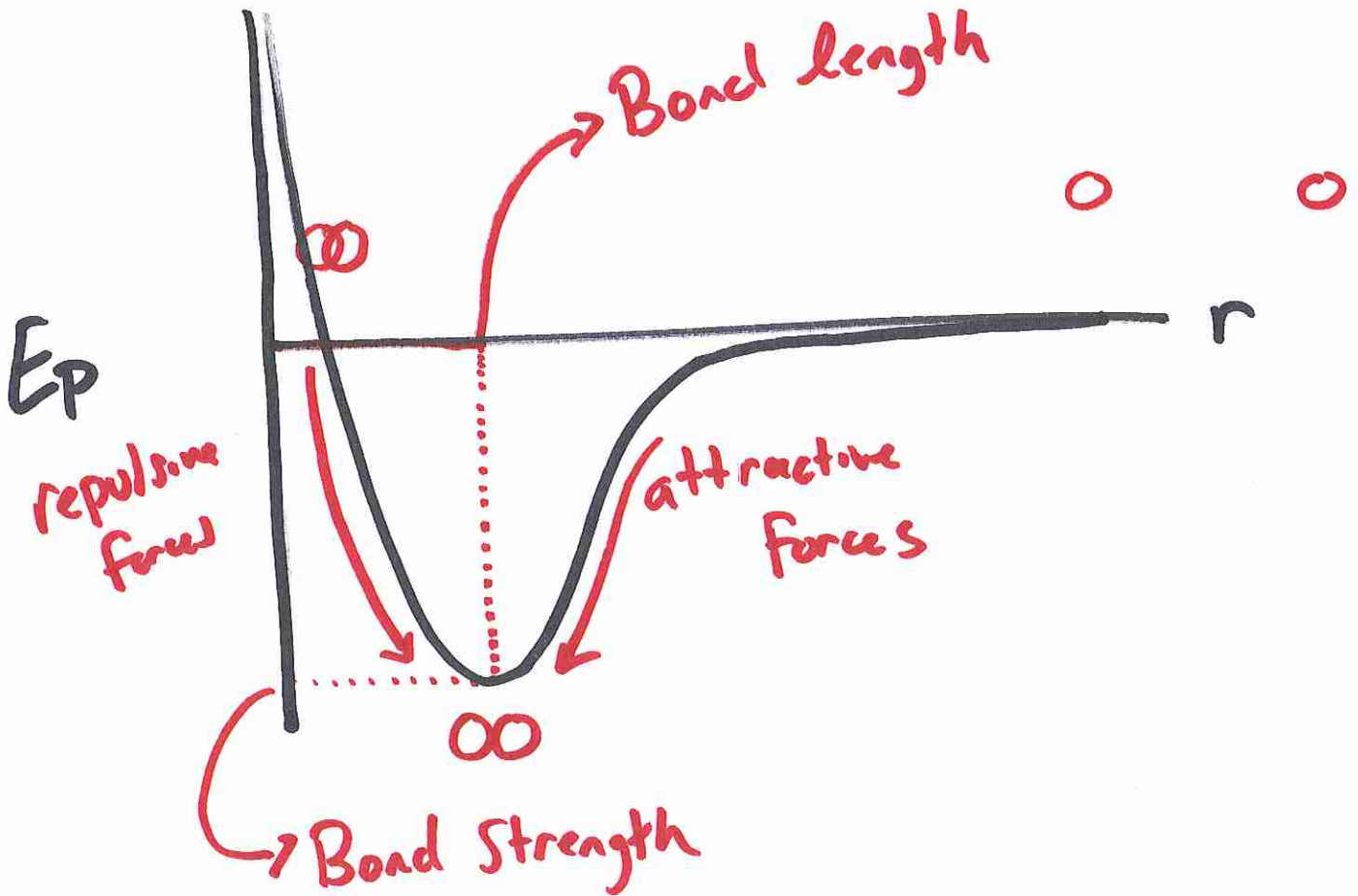
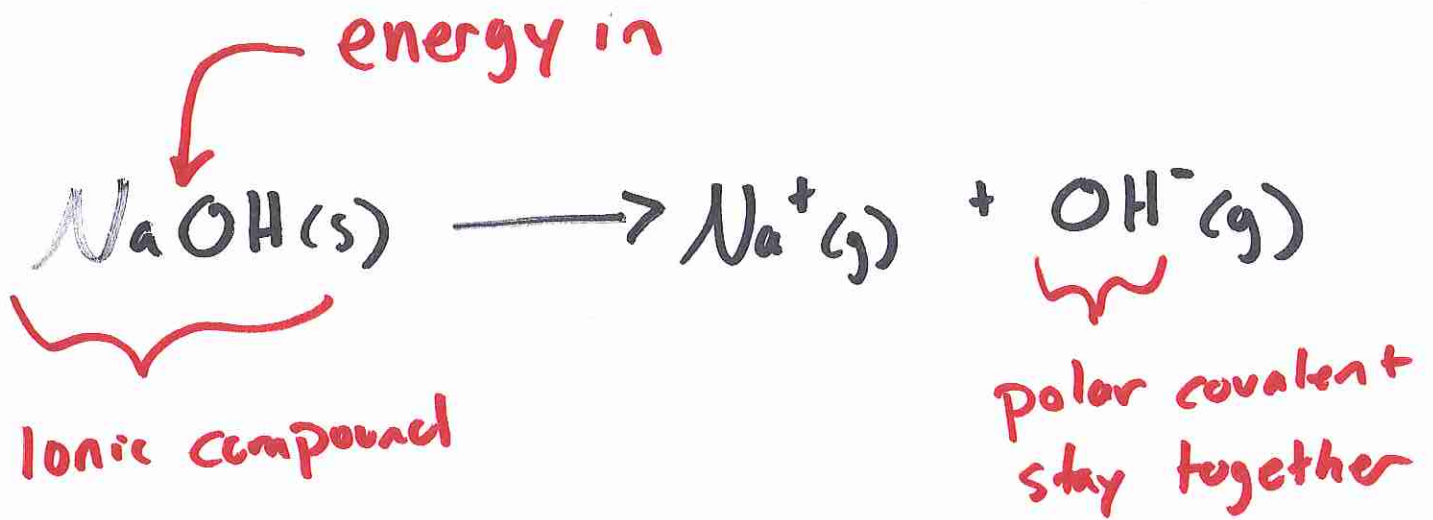


Covalent Bond Strength and Stability

- **Bond strength**: the energy required to break a covalent bond (also known as Bond Dissociation Enthalpy)
- **Bond order**: a measurement of the average number of chemical bonds between two atoms in a species (1= single bond, 2 = double bond, 3= triple bond)
- **Bond length**: the inter-nuclear distance between atoms in a bond
- **KEY: stronger bonds are shorter; higher order bonds are stronger (and thus shorter)**



Notice how when a covalent bond gets just right, your potential energy is minimized just like lattice energy



Covalent Bonds and Stability

- Bond strength and bond order are important indicators of stability
- Single bonds are the weakest, lowest order, and the longest.
- Triple bonds are the strongest, highest order, and the shortest.
- Resonance gives rise to fractional bond orders, which follow the same trend
 - A 1.5 bond is stronger than a single bond and weaker than a double bond

Bond Type (Carbon-Oxygen)	Bond Energy (kJ/mol)	Bond Length (pm)
Single	358	143
Double	799	120
Triple	1072	113

Lewis Structure Formal Charge Foundations

Element	-1 Charge	Neutral	+1 Charge
Hydrogen		1 bond	
Carbon*	3 bonds, 1 LP (rare)	4 bonds	
Nitrogen	2 bonds, 2 LP (rare)	3 bonds, 1 LP	4 bonds
Oxygen	1 bond, 3 LP	2 bonds, 2 LP	3 bonds, 1 LP
Halogens**	0 Bonds, 4 LP	1 bond, 3 LP	

This chart represents a “best guess” to draw out simple Lewis Structures. The formal charges are calculated based on a full octet and no exceptions.

*Carbon almost always has four bonds. Some notable exceptions are CO, CN⁻, CNO⁻

**Represents halogens as a peripheral atom, not a central atom. In other words, this applies to the fluorine in ClF₃ but not the chlorine.

Lewis Structures

- Lewis structures are the main form of drawing molecules. There are two important equations to know when drawing Lewis Structures:

$$S = N - A$$

- The number of shared electrons is equal to the total electrons **needed** to fill the valence (2 for hydrogen, 8 for other non-exception elements) minus the electrons **available** (valence electrons). Account for positive charge by subtracting from available; account for negative charge by adding to available.
- You can then calculate the number of bonds by dividing the shared electrons by two (bonds = S/2)**

$$FC = \text{Valence} - (\text{lone electrons} + \text{bonds})$$

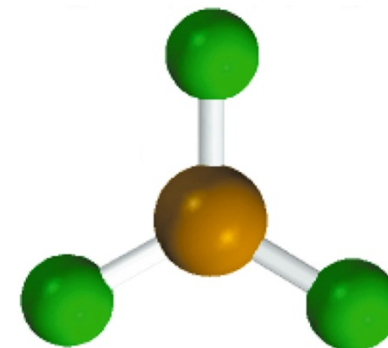
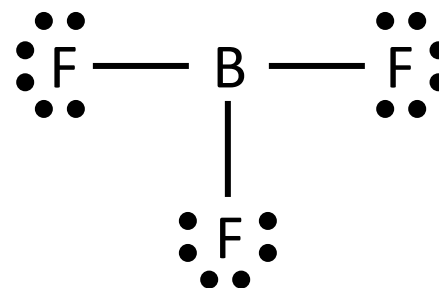
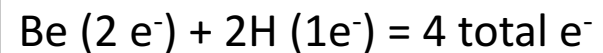
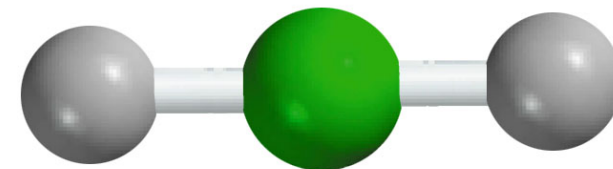
- The formal charge of any atom in a molecule is equal to the the valence electrons of the atom (as seen on the periodic table) minus the “things it’s touching” in the Lewis structure (the lone electrons and the number of bonds)**

$$FC = \text{Valence} - (\text{“things it’s touching”})$$

Incomplete Octet

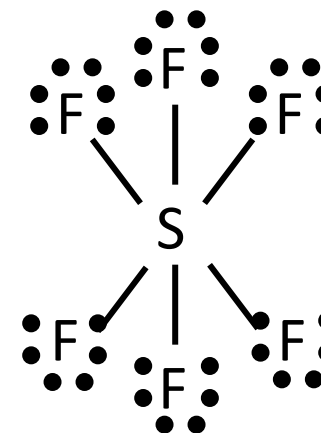
Exam Tip: there are only a few good examples of the incomplete octet: BeX_2 and BX_3 , where X is either Hydrogen or a Halogen.

- The octet rule will also be useless for central atoms with fewer than 4 valence electrons (ex: **Boron** and **Beryllium**).
- Beryllium (Be) is satisfied making **2 bonds** (4 valence e^-).
- Boron is satisfied making **3 bonds** (6 valence e^-)
- Boron and beryllium don't have many valence electrons and they have low electronegativities, so we consider them to be **electron-poor** elements with little desire to fill their octets.



Lewis Structures: Expanded Octet

1. Find the number of electrons **available** (valence electrons plus the charges – subtract for positive charge, add for negative).
2. Create your skeletal structure, with one bond to each peripheral atom
3. If necessary, add more bonds to minimize the total number of atoms with formal charges*
4. Add lone pairs to your peripheral atoms
5. If you have any electrons left based on your **available** electrons from step 1, add them to your central atom.
6. Confirm your molecule's structure by checking the formal charge

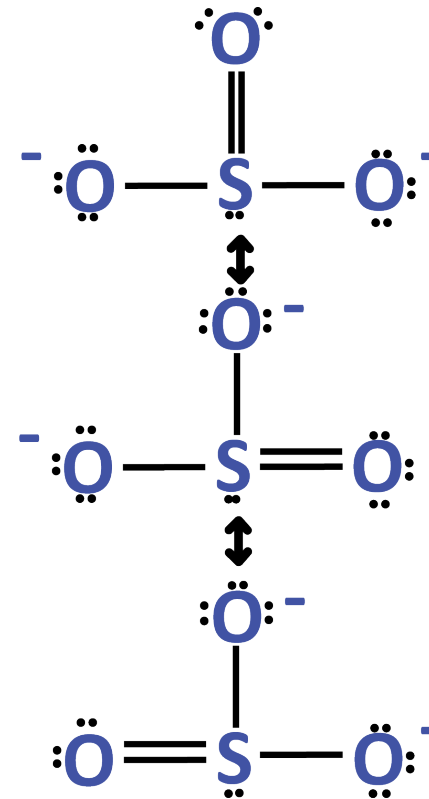


Expanded octets become possible for the central atom after Aluminum, due to the valence shell's access to empty 3d orbitals

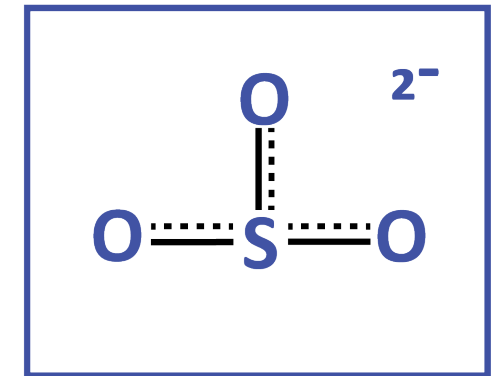
$$B.O. = \frac{\text{total \# of bonds in resonance}}{\text{total \# of bonding regions in resonance}}$$

Bonding and Lewis Structures

- **Resonance:** multiple acceptable Lewis Structures are available for a given compound, meaning that the compound actually exists as the average of all acceptable structures
- **Delocalized electrons:** in the case of resonance, electrons are not confined to a single bond. Instead the electron charge is distributed over multiple bonds. We call this delocalization.
- **If you have multiple acceptable Lewis Structures in resonance, your actual structure will have the bonding character of the average. Example: Sulfite (SO_3^{2-})**



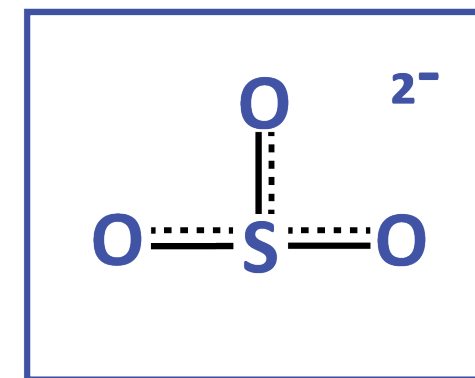
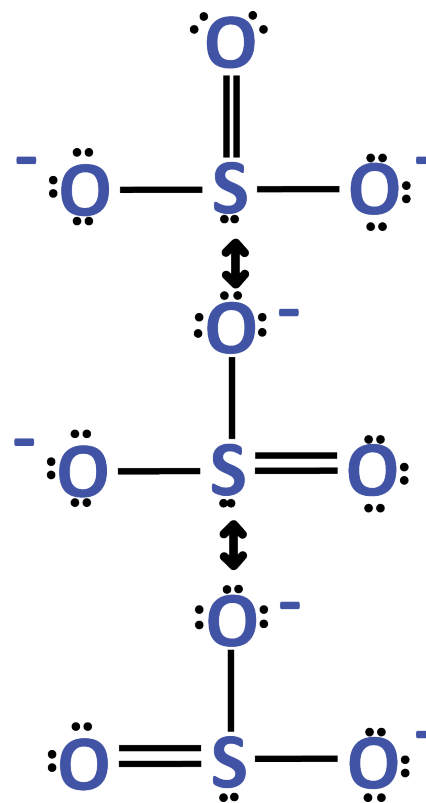
Correct Structure: three
1.33 bonds



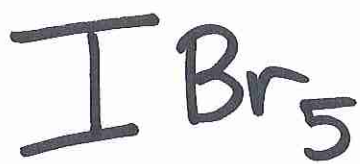
“Acceptable” Structures: double and single bonds

Rules for an Acceptable Structure

- What constitutes an “acceptable” Lewis Structure?
 1. The negative charge is placed on only the electronegative or peripheral atoms, unless no better structure is possible (for example, CO and CN⁻). Positive charges on the central atom are possible; however....
 2. Any individual charge should not exceed +/- 1
 3. Put your charges on as few atoms as possible
 4. Carbon, nitrogen, oxygen, and fluorine cannot disobey the octet rule
 5. **Your structure accurately reflects the number of available electrons in your atoms**



:

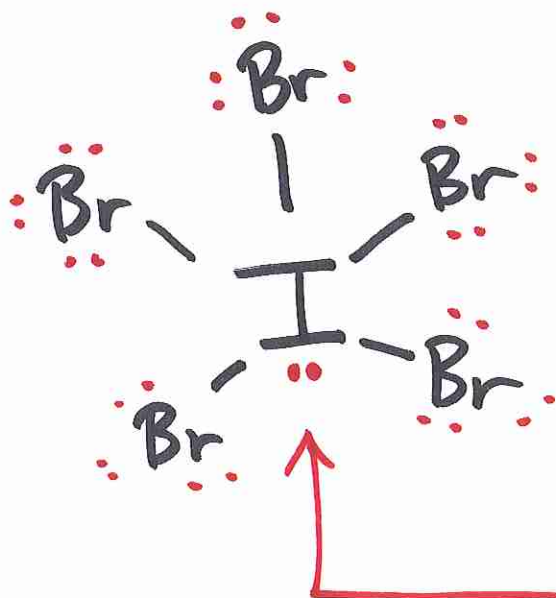


Available:

$$7 + 7(5) = 42e^-$$

$$\text{skeletal: } 40e^-$$

2e⁻ remaining

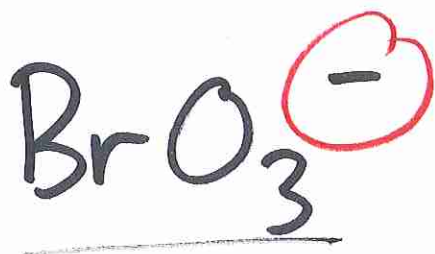


e⁻ geometry: octahedral

molecular: square pyramidal

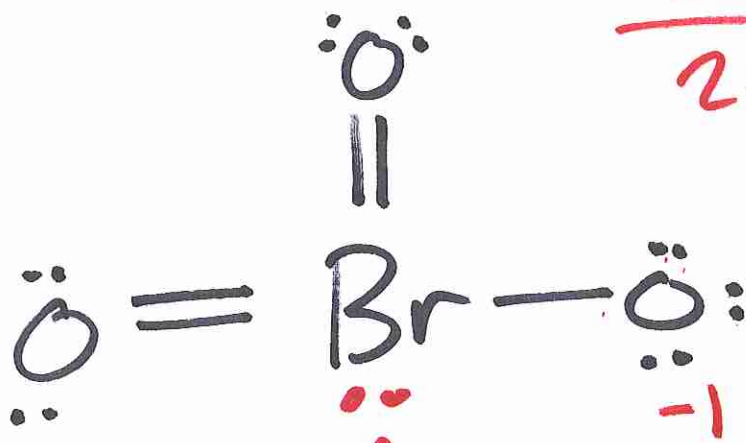
Bond angles: 90° , 180°

Polar? Sure



Available: $7 + 3(6) + 1$

$$\begin{array}{r} 26e^- \\ 24e^- \\ \hline 2e^- \end{array}$$



Resonance:

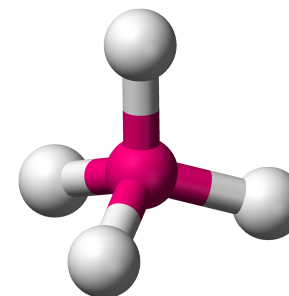
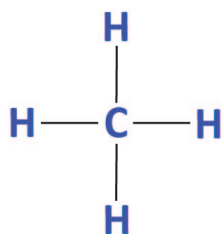


$B_o = \frac{5}{3} = 1.67$

Note: none of this means anything if your Lewis Structure is drawn incorrectly!

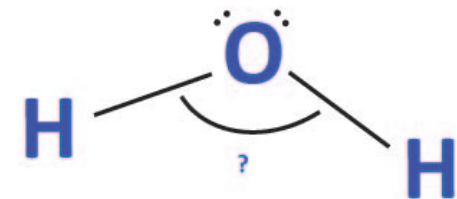
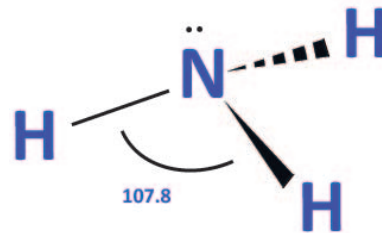
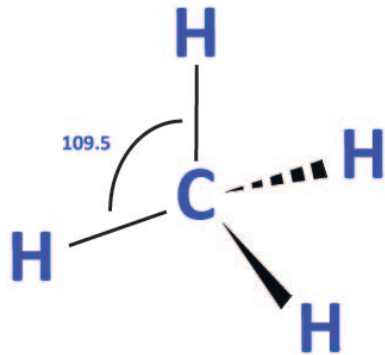
Valence Shell Electron Pair Repulsion

- The basis of VSEPR is that electron-rich regions will achieve the lowest possible energy state by minimizing repulsions (maximizing the angles to spread out)
- VSEPR gives us two different 3D geometries: **electronic geometry and molecular geometry**.
 - **Electronic geometry**: the shape of all electron-rich regions around a central atom, determined by counting the bonding AND lone pair regions
 - **Molecular geometry**: the shape of only the bonding regions around a central atom, determined by the electronic geometry and the (only) bonding regions
- The VSEPR shape is essential for the “actual” shape of the molecule, which gives us insight into the molecule’s bond angles and polarity.



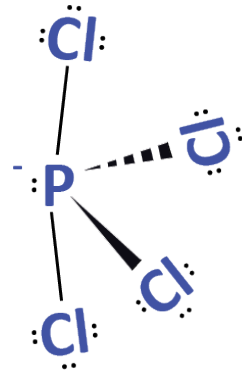
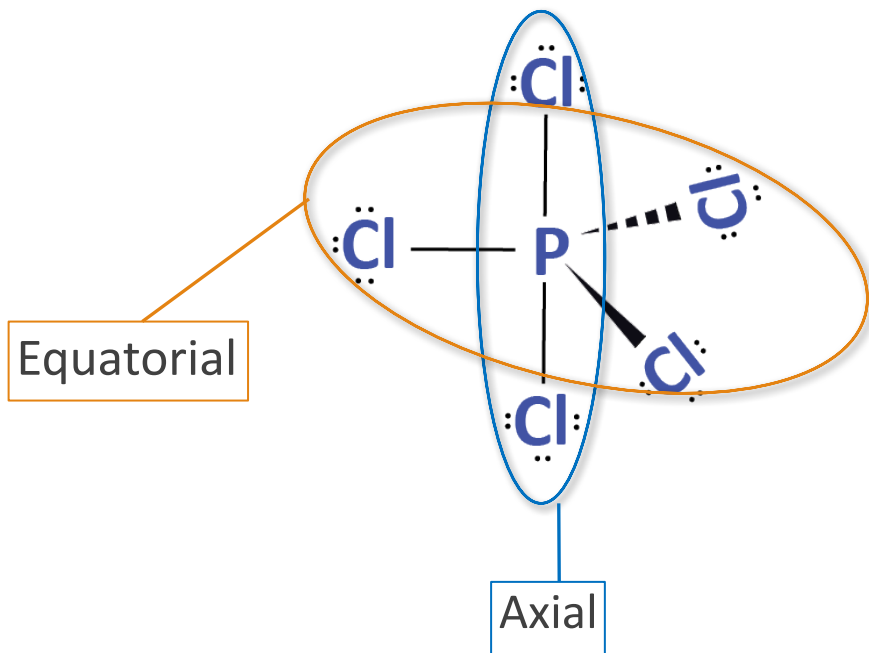
The Effect of Lone Pairs

- Lone pairs occupy more space on the central atom than a bonding pair of electrons, **meaning they end up being more repulsive than bonding regions**
- The electron density of lone pairs is **ALL on the one central atom** – not shared between two – as in a conventional bond
- **The extra repulsive strength of the lone pair tweaks the bond angles of the central atom, making them smaller than expected for a pure geometry**

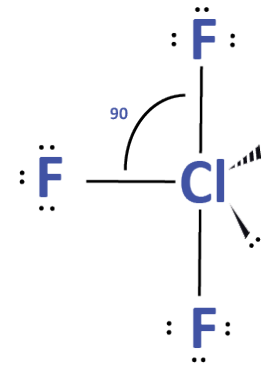


VSEPR Geometries: 5 e- Regions

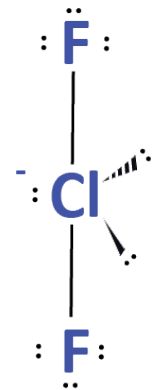
- The trigonal bipyramid geometry is the one base electronic geometry where lone pairs actually have a preference on position. We pick on this one for conceptual questions.



Molecular:
Seesaw



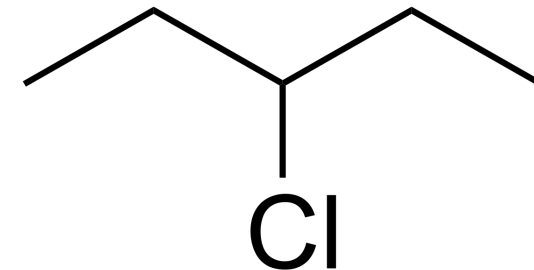
Molecular:
T-Shaped



Molecular:
Linear

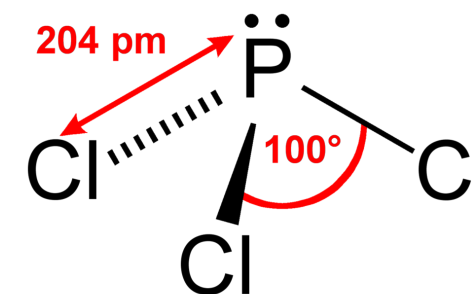
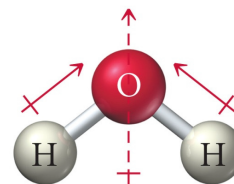
<u>Bonding Regions</u>	<u>L.P.'s</u>	<u>Molecular Geometry</u>
4	0	Tetrahedral
2	2	bent
4	1	seesaw
3	3	T-shape
5	1	Square Pyramidal
2	3	Linear

Determining Molecular Polarity



- Polarity in a **molecule** occurs when there is a permanent dipole moment on the molecule. This can be identified by a few rules:
 - **A perfectly symmetrical molecule will always be nonpolar if its bonds are to the same atom.**
 - Examples: CO_2 , SO_3 , SF_6 , PCl_5 , CCl_4 are non-polar molecules even though they have polar bonds
 - **A central atom bound to different peripheral atoms will always be polar, as long as there is at least one polar bond.**
 - Examples: CH_3Cl , CHF_3 , PCl_4F
 - **Molecular geometries with lone pairs are polar (even if the bonds are non-polar), except in linear and square planar geometries**
 - Examples: H_2O , NH_3 , H_2S , O_3

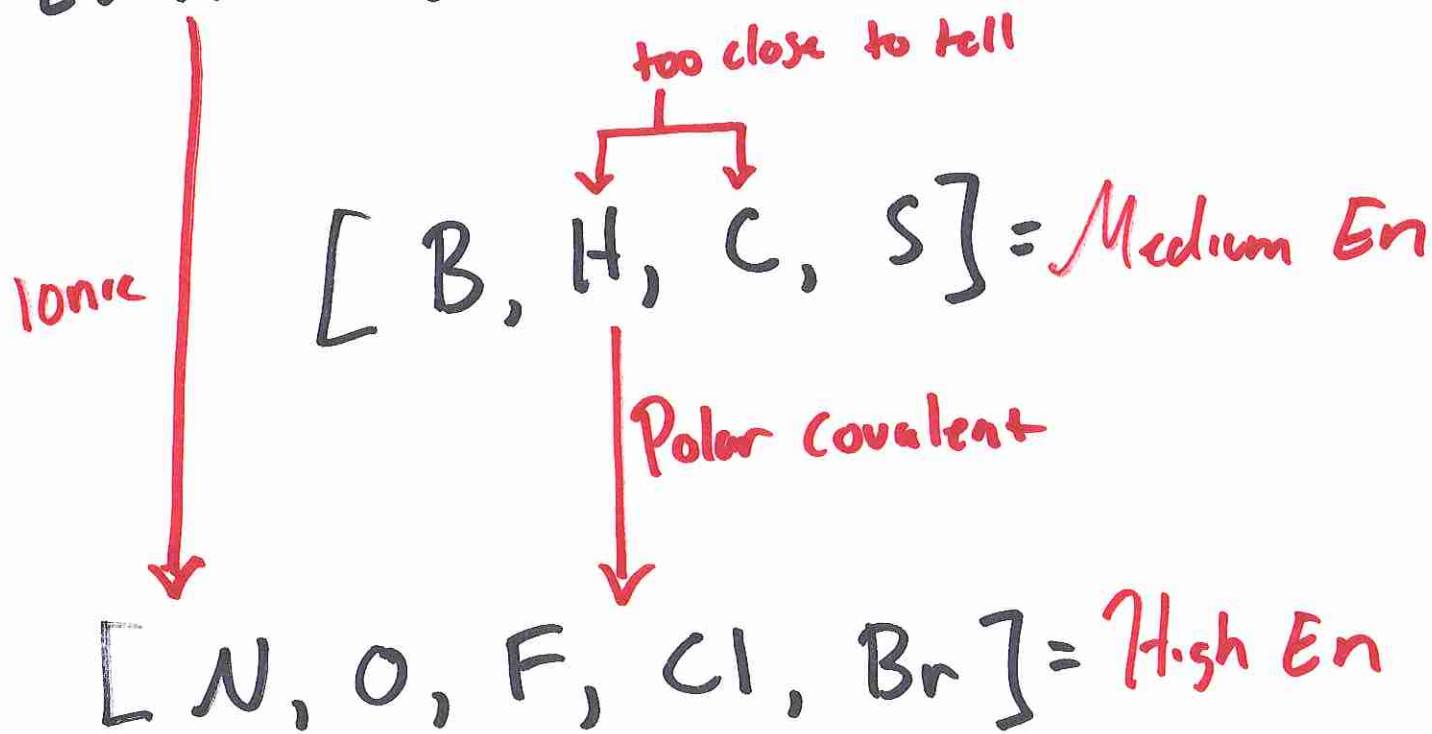
Net dipole moment


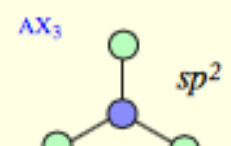
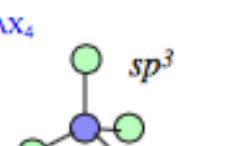
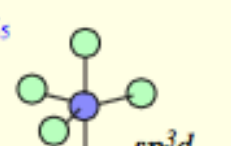



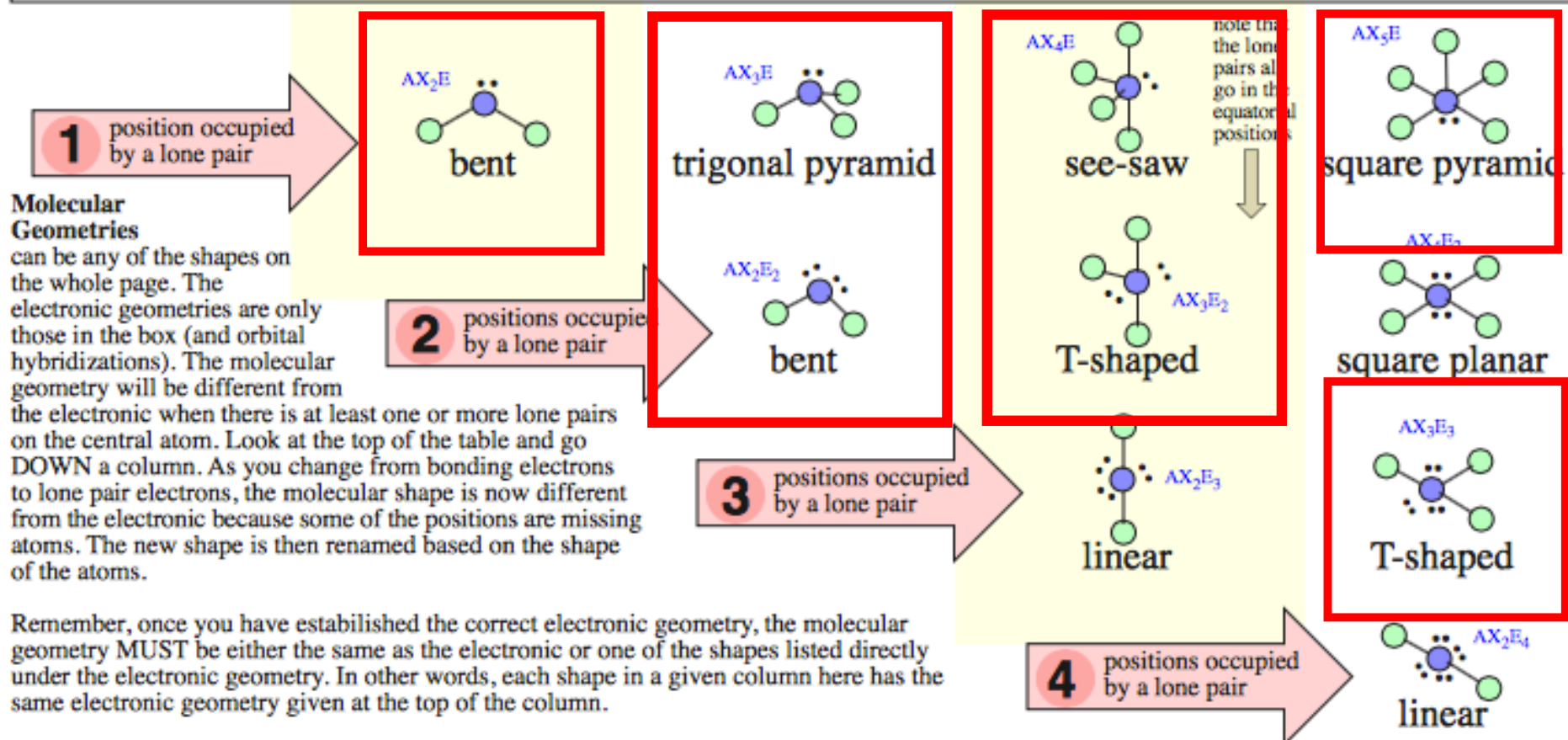
More than just polar bonds: To determine the symmetry of the molecule and the presence of lone pairs, you must draw the Lewis Structures and consider the VSEPR geometry (electronic AND molecular).

Polarity Cheat Sheet

[Metals] = Low En



2 electron regions	3 electron regions	4 electron regions	5 electron regions	6 electron regions
AX_2  <i>sp</i>	AX_3  <i>sp²</i>	AX_4  <i>sp³</i>	AX_5  <i>sp³d</i>	AX_6  <i>sp³d²</i>
linear	trigonal planar	tetrahedral	trigonal bipyramid	octahedral



Molecular Geometries can be any of the shapes on the whole page. The electronic geometries are only those in the box (and orbital hybridizations). The molecular geometry will be different from the electronic when there is at least one or more lone pairs on the central atom. Look at the top of the table and go DOWN a column. As you change from bonding electrons to lone pair electrons, the molecular shape is now different from the electronic because some of the positions are missing atoms. The new shape is then renamed based on the shape of the atoms.

Remember, once you have established the correct electronic geometry, the molecular geometry **MUST** be either the same as the electronic or one of the shapes listed directly under the electronic geometry. In other words, each shape in a given column here has the same electronic geometry given at the top of the column.

Polarity
If all the positions on the electronic geometry are the same (have the same atoms surrounding the central atom), the molecule is **NOT** polar because of the symmetry. Any of the other molecular geometries (except square planar and linear) under the box will be polar.

Basic rules for molecular polarity:

1. Look for asymmetry in connectivity.
2. Look for lone pair asymmetry.

All red shapes are polar

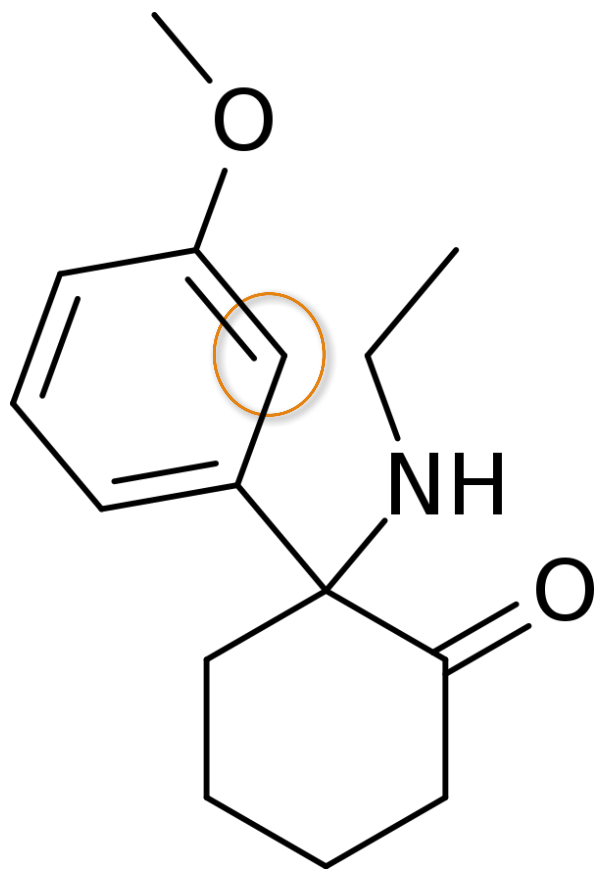
Molecule

Non-Polar

Polar

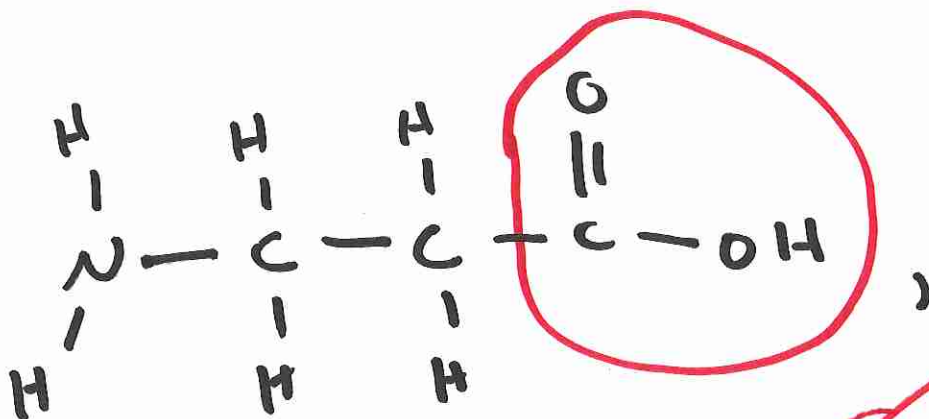
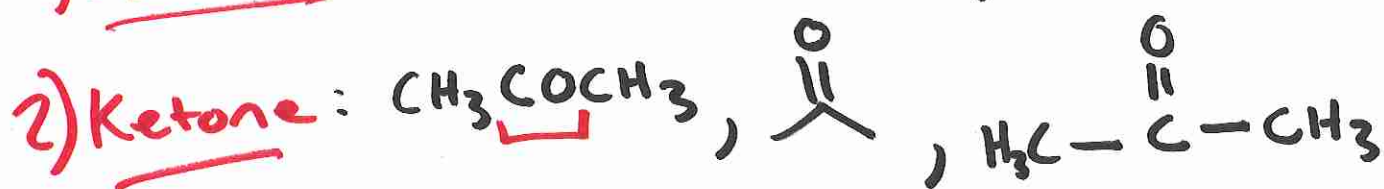
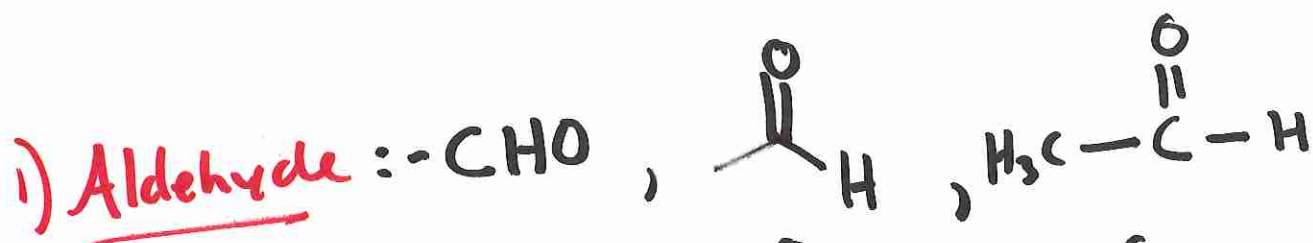


Big Molecule Question Checklist



Can you:

- Write the empirical formula for this compound? $C_{15}H_{21}NO_2$
- Determine the electronic and molecular geometries around any given central atom?
- Determine the bond angles around any given central atom?
- Let's work the opposite direction for a simple carbon chain molecule: Can you draw $NH_2CH_2CH_2COOH$?



Helpful, not
required
for exam 3