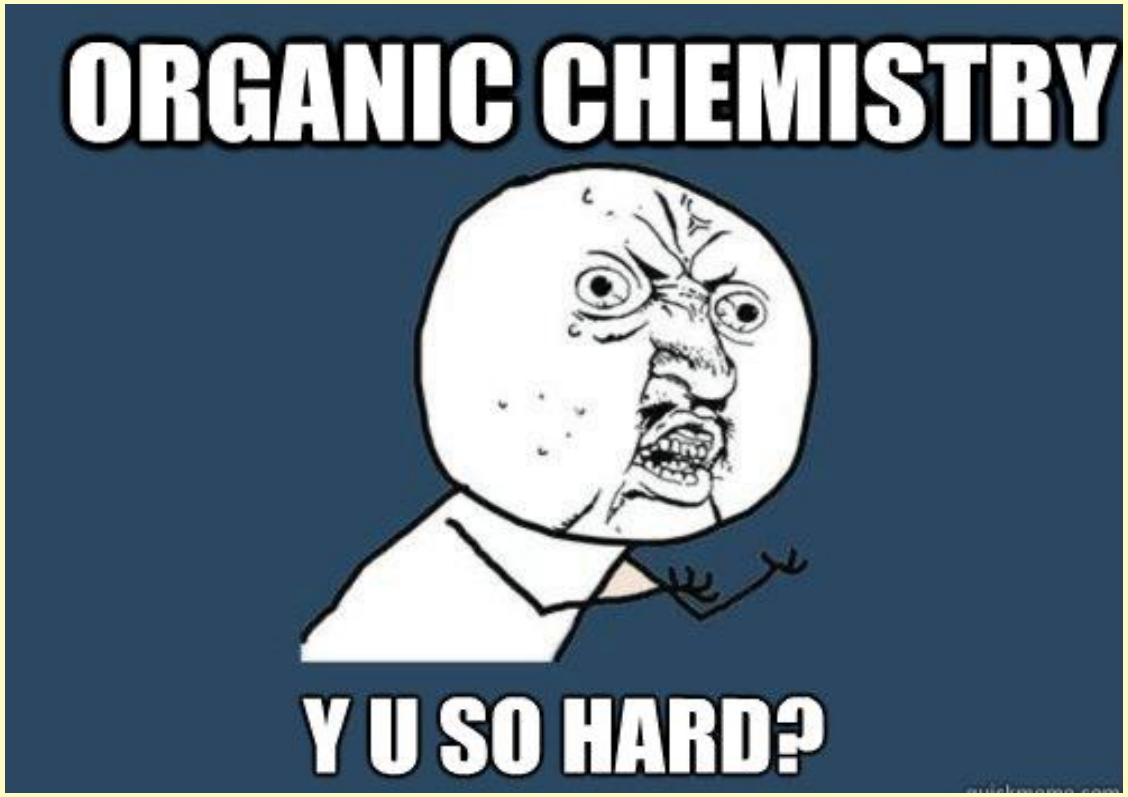
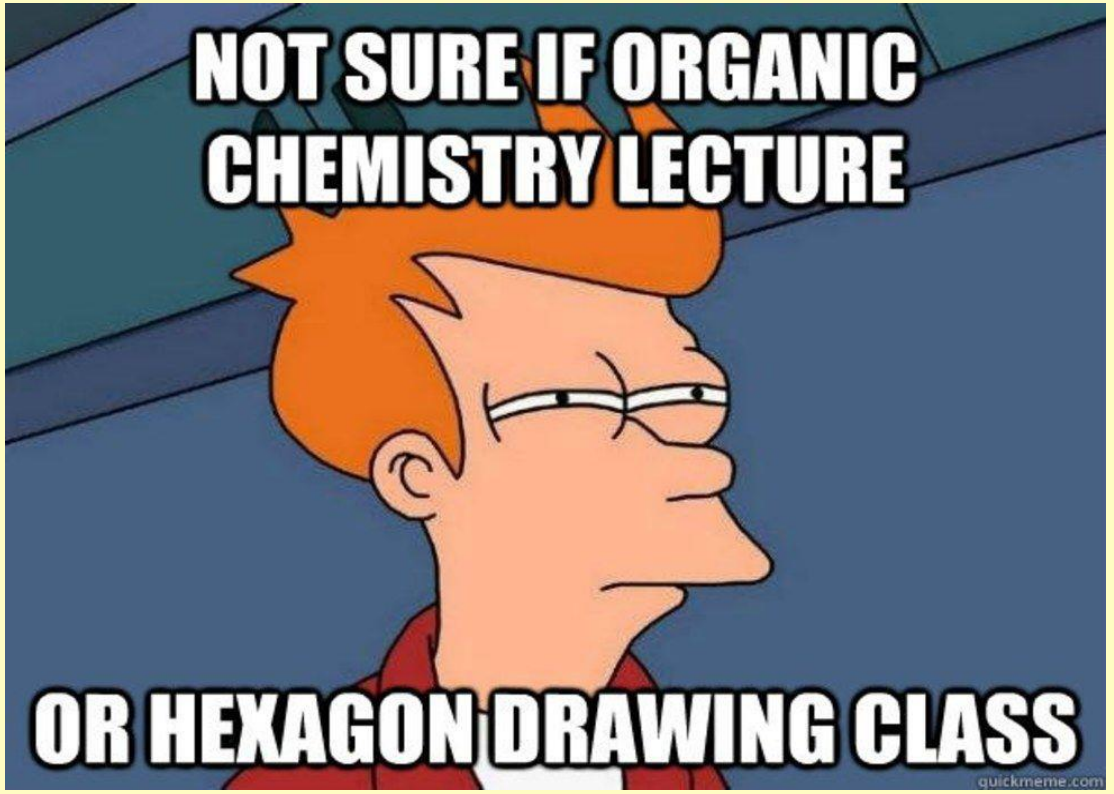




WELCOME TO ORGANIC CHEMISTRY

THE CHIRAL CENTER

How to Succeed in Organic Chemistry

- Believe **'You Can Do It!'**
- Read the Textbook Before Class
- Organize Notes Before Class
- Use Lecture to Refine Understanding
- Thoroughly Complete Homework
  - Compare first attempt to answer key and seek to learn from mistakes
- Practice Textbook Questions
  - Use Student Solution Manual to check answers and check understanding
- Ask Questions!



Resources to Help You Succeed

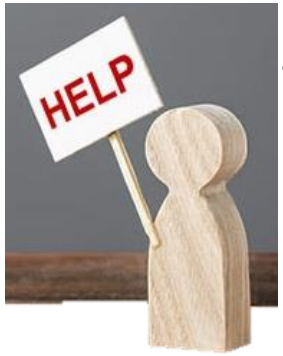


Let's start with the obvious:

- **Your Textbook** – Read the book for an organized layout of the material
  - The Student Solution Manual is available for rent as an eBook.
- **Your Homework** – This is an opportunity to judge your own understanding.
- **Office Hours** – Come *with questions*

Other resources:

- **Molecular Model Kits** – Free resources are available in the Skirball Science Learning Center or at Dr. Hamann's office.
  - You can purchase your own model kit. I recommend the one by Maruzen.
  - You can use an online version for convenience: Virtual Model Kit
- **YouTube's Organic Chemistry Tutor**
  - Great for practice with the basics. Note that this resource is not a textbook and is not 100% accurate, nor does it always reflect the material as taught by Dr. Hamann.
- **Chad's Online Prep Course** – Need even *more* help than we can provide?
  - This is a reputable online resource for Organic Chemistry.
  - There is a free trial available.





Your Recitation Grade



- **Determined by Homework Only**

  Due Date 1: Anytime before REC on relevant material

  - **You must submit completed HW to receive credit.**
  - During Recitation, we will:
    - Briefly review challenging concepts
    - Discuss general chapter concepts
    - Review challenging questions

  Due Date 2: Anytime before the *following* REC

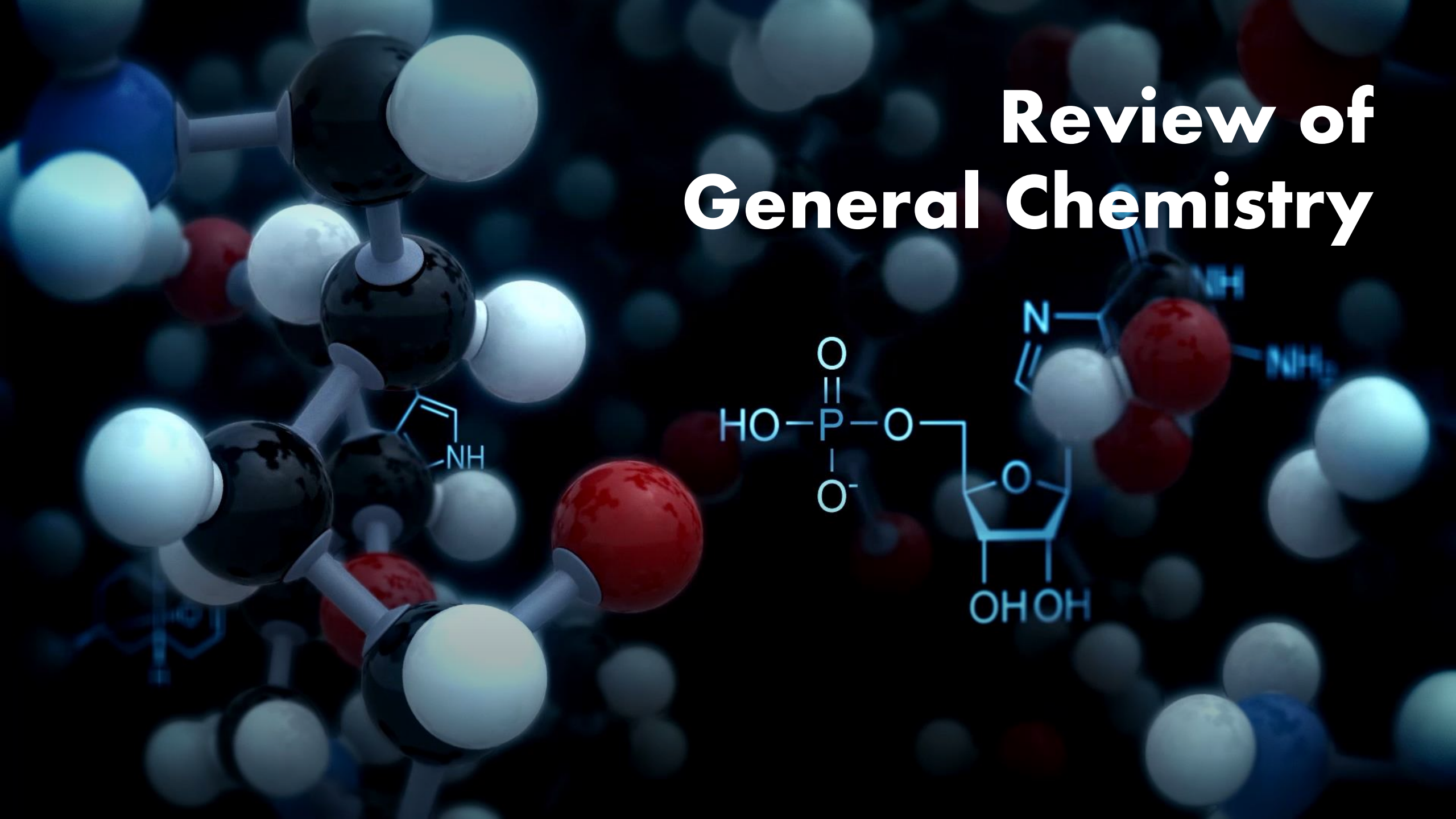
  - **The second submission is graded, based on my ability to see "learning on the page."**
  - The second submission must include corrections to first submission (in a different color)
    - Should include notes on what you learned by fixing the mistake
  - *Note*: Homeworks are still assignments and should not be full of scribbles and scratched-out answers. A human is grading them and if I cannot read the answers, I will not be as lenient.

- **Active Participation is Strongly Encouraged**

  For Pre-Health: I will not write evaluations for students who did not actively participate.



Review of General Chemistry



Isomers

- **Constitutional (Structural)**

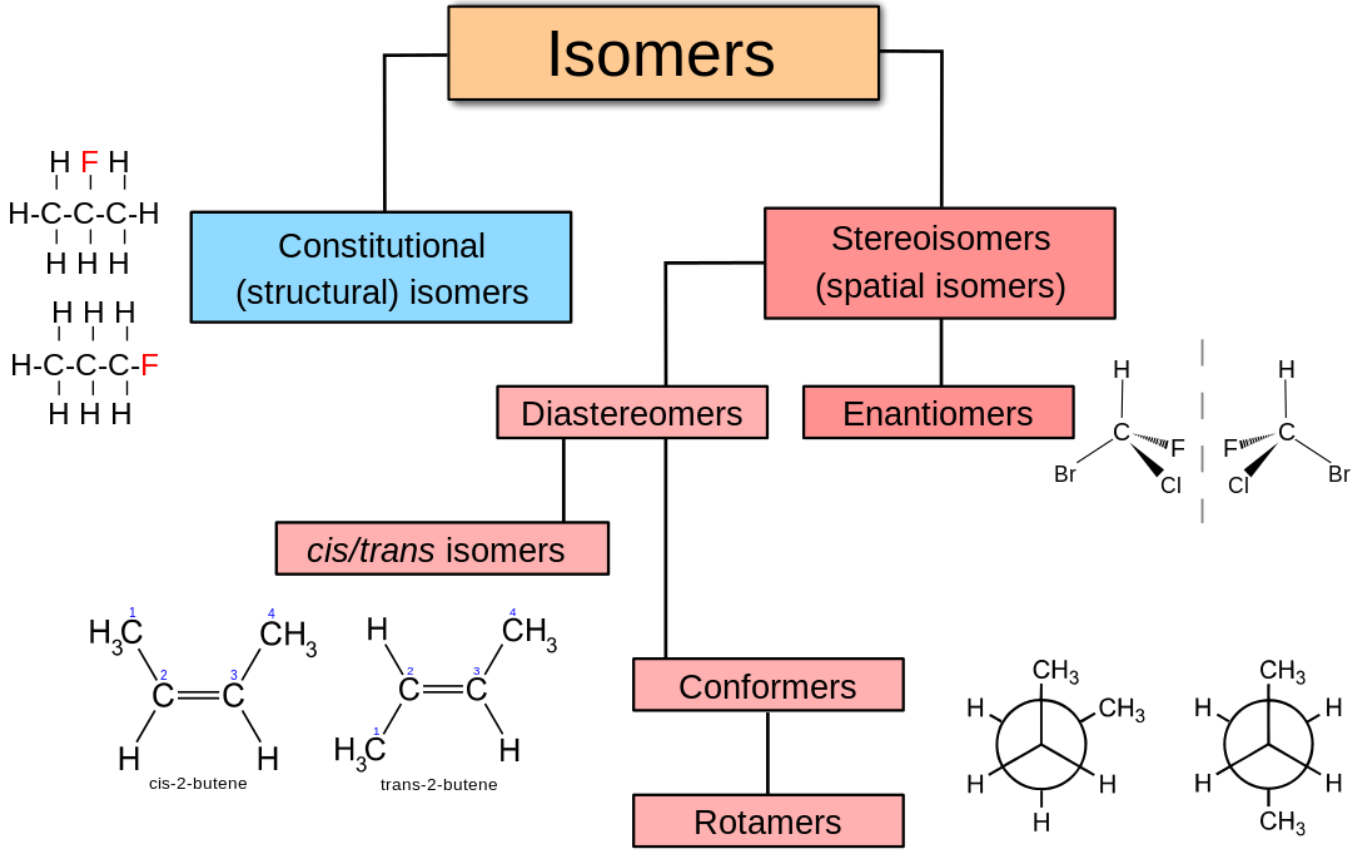
  Molecules that have the same molecular formula with different bonding patterns.

  – you should know these *now*

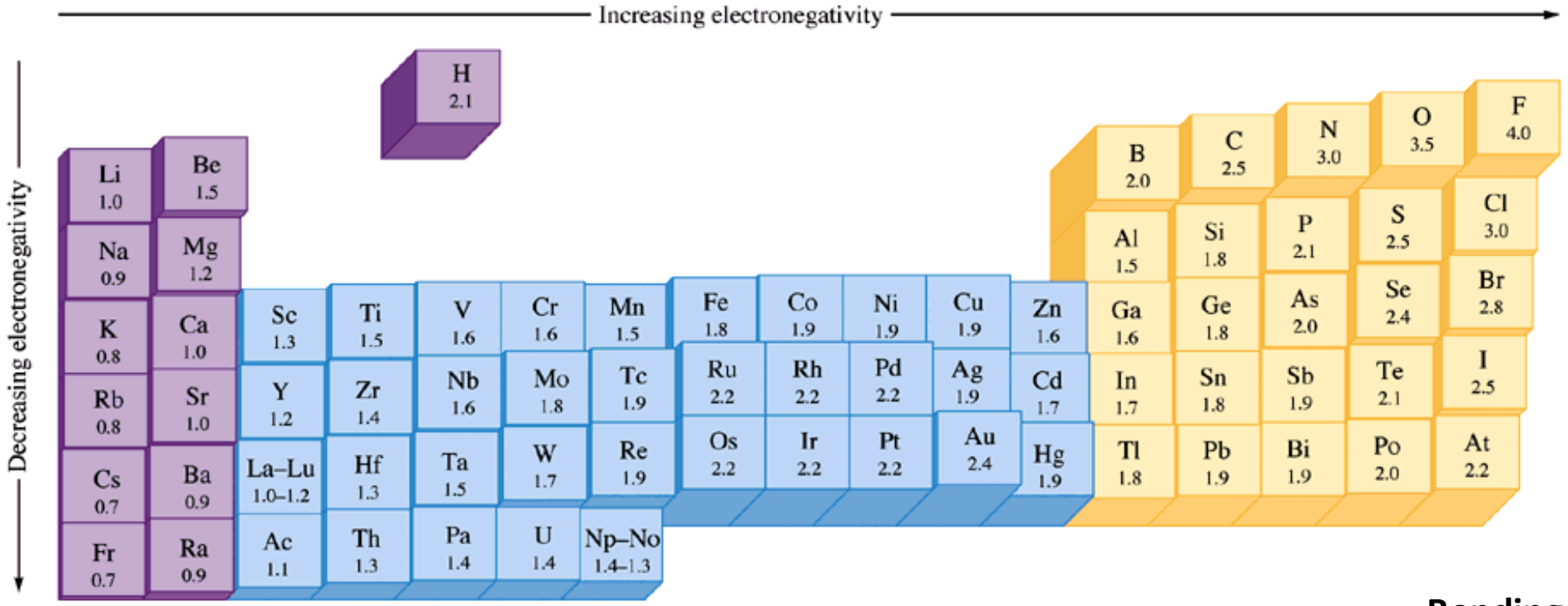
- **Stereoisomers**

  Molecules that have identical molecular formulas and bonding patterns. They differ only in spatial arrangement.

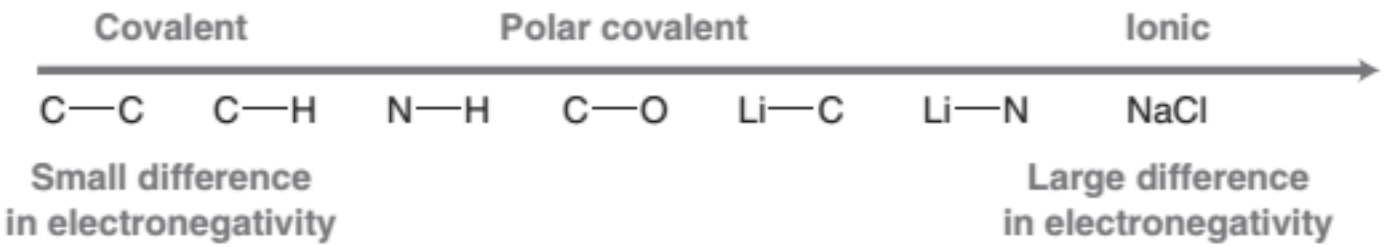
  – we will learn these *later*



Periodic Trends: Electronegativity



***Note*: Noble gases have an EN value of 0**



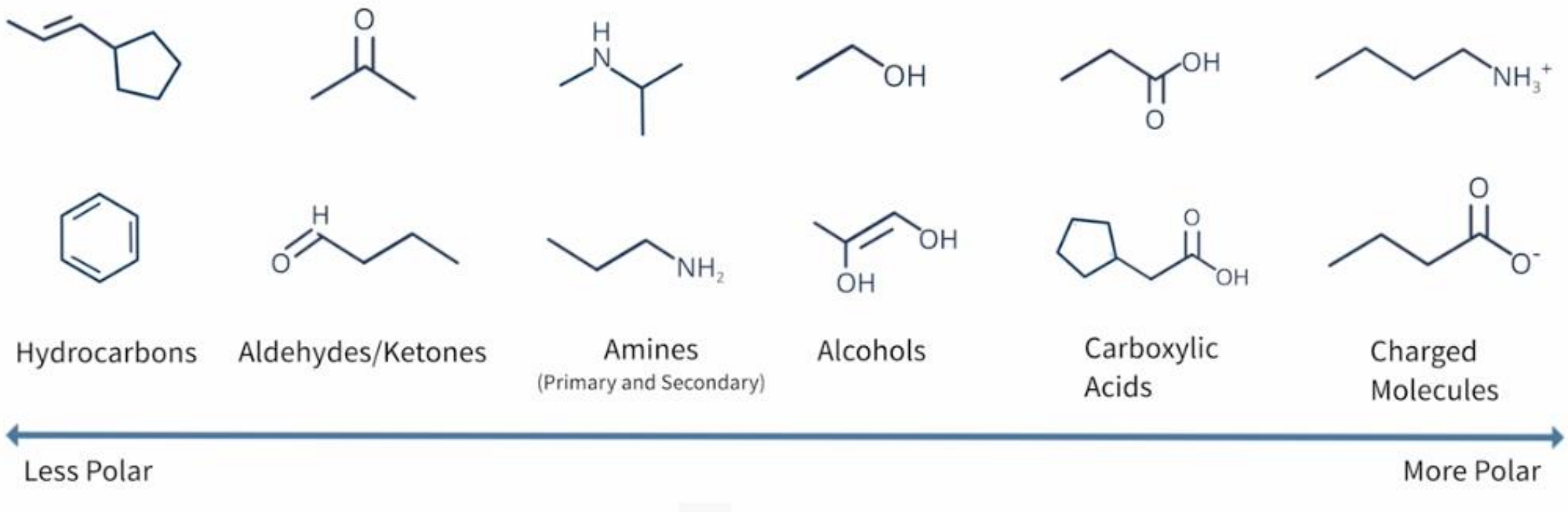
Bonding Type

ΔEN	**BOND**
< 0.5	non-polar covalent
0.5–1.7	polar covalent
> 1.7	ionic

Polarity

Polarity is the separation of electric charge (i.e. concentration of electrons) leading to an electric dipole moment, with a negative pole at one and a positive pole at the other end. This polarity is due to a difference in electronegativity (EN) between the bonded atoms.

The greater polarity exists when a formal charge is present.



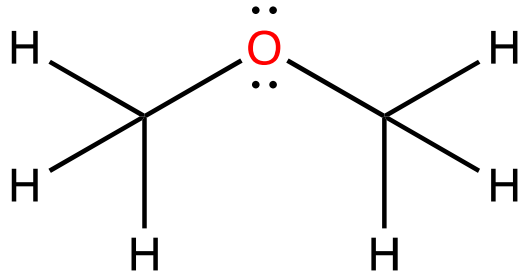
Partial Charge: δ^+ / δ^-

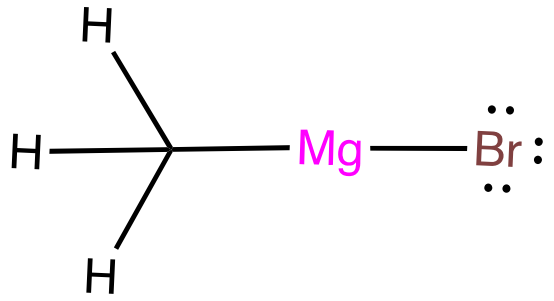
We can determine reactivity of certain compounds by observing the polarity of bonds.

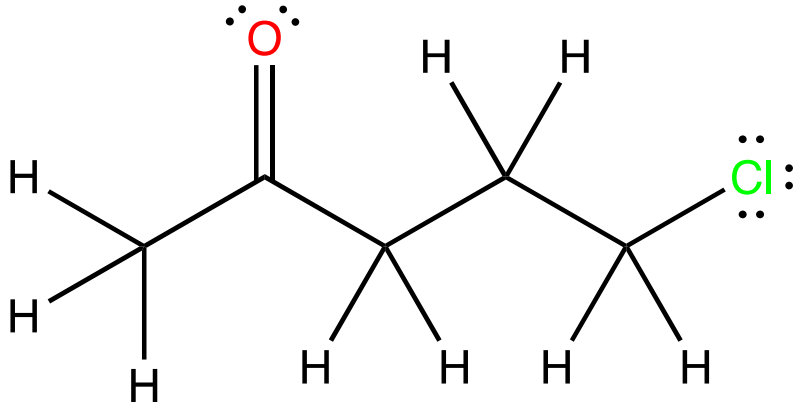
- **Ionic bonds** will result from significantly different electronegativity (EN) values, greater than 1.7.
- **Polar covalent bonds** result from a EN difference between 0.5-1.7. The more EN atom will withdraw the electron density from its neighbor, in a process called **induction**.
- **Covalent bonds** result from a EN difference that is <0.5.

We can label the molecules to reflect these partial dipoles, using the following symbols: δ^+ or δ^-

Note: Due to their opposite charges, strong δ^- are likely attack (i.e. react with) δ^+ regions on neighboring molecules.



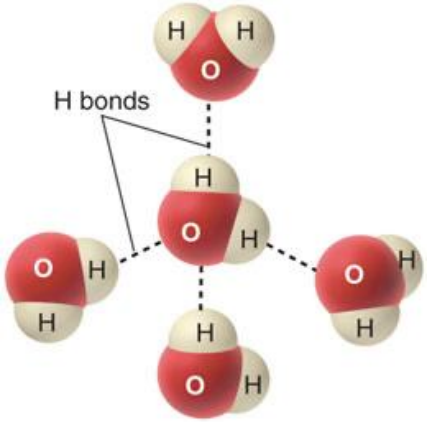


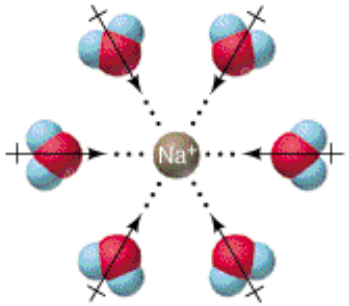


Intermolecular Forces (IMFs)

IMFs are forces that mediate interactions between molecules, including forces of attraction and repulsion between molecules and neighboring particles.

- Ion-Dipole
- H-Bonding (via N/O/F)
- Dipole-Dipole
  - – interactions between stable dipoles
- London Dispersion Forces
  - – instantaneous dipole induced as a result of e^- movement

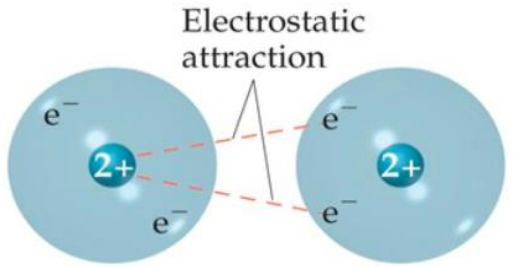




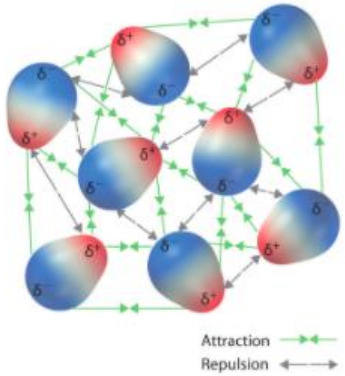
Ion-dipole

Properties of matter that depend on IMFs:

- BP / MP (and Vapor Pressure)
- Enthalpies of Phase Changes
- Solubility
- Viscosity



London Dispersion



Dipole-Dipole

Valence Electrons

Valence electrons are electrons in the highest occupied principal energy level of an atom. While they are associated with an atom, these electrons can also participate in a chemical bond (if the outer shell is not already complete).

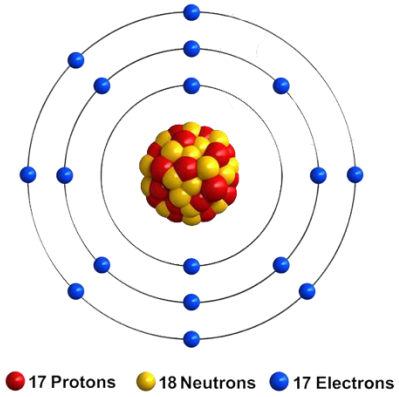
Some valence electrons are more easily removed (or gained) because an atom is closer to a noble gas configuration. Thus, the number of valence electrons determines the reactivity of the atom.

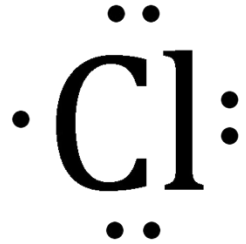
All electrons in lower energy sublevels are called **inner-shell electons**.

Note: Row 3 elements and below can expand their valence shell beyond 8 electrons!

The transition metals also have 2 valence e^-, but they have some unique properties due to an orbital subshell that is close to the valence or outer shell level

1																	2
1	2										3	4	5	6	7		8
1	2										3	4	5	6	7		8
1	2										3	4	5	6	7		8
1	2										3	4	5	6	7		8
1	2										3	4	5	6	7		8
1	2										3	4	5	6			





Calculating Formal Charge

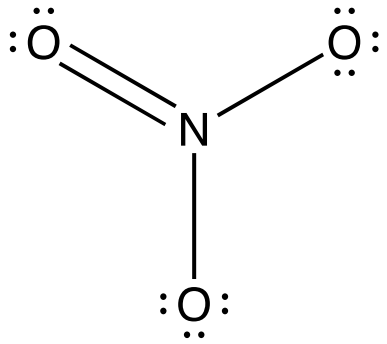
Formal Charge (FC) is a calculated number that can estimate the charge distribution within a molecule.

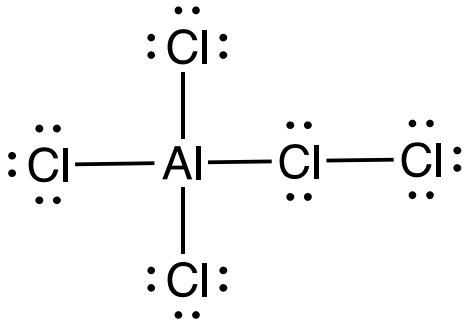
FC = (valence e^-) – (non-bonding valence e^-) – (bonding valence e^-)

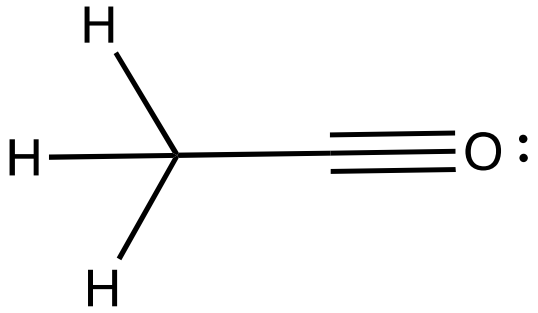
OR

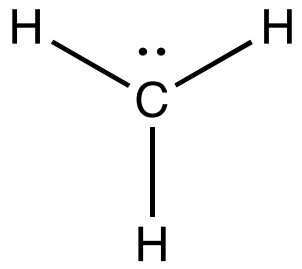
FC = (valence e^-) – "dots" – "lines"

Let's Practice.









Calculating Formal Charge

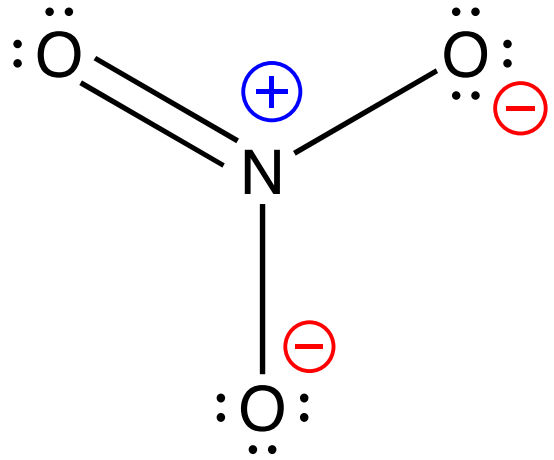
$FC_N = 5 - 0 - 4 = +1$

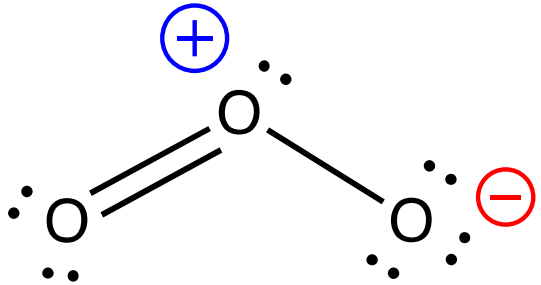
$FC_{O_2} = 6 - 4 - 2 = 0$

$FC_{O_2} = 6 - 4 - 2 = 0$

$FC_{O_1} = 6 - 6 - 1 = -1$

$FC_{O_1} = 6 - 6 - 1 = -1$





Lewis Structures

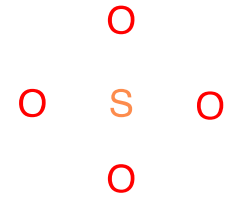
1. Count total electrons
2. Put least EN atom in center & surround w. more EN atoms (do not include H's and metal ions)
3. Create single bonds to central atom
4. Complete octets of peripheral atoms
5. Give remaining electrons to central atom (if applicable)
6. Calculate formal charge (FC)
7. Form double/triple bonds to minimize FC on neighboring atoms
8. Add H^+ or metal atoms to counter FC

- ***Remember***: Row 3 elements and below can have expanded octets!

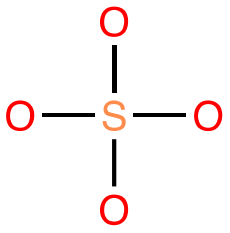
1 Ex. H_2SO_4

2H = 2(1) = 2
S = 6
4O = 4(6) = 24
Total = 32 e^-

2

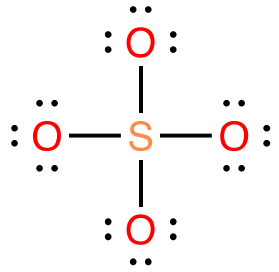


3



32 – 8 = 24 e^- remain

4

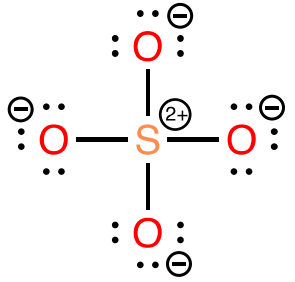


24 – 24 = 0 e^- remain

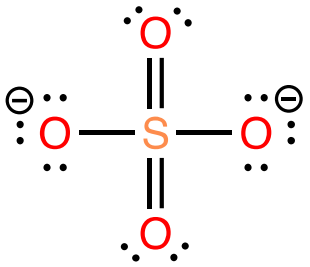
5

Skip this step

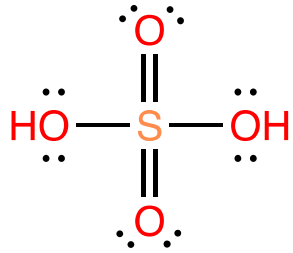
6



7



8



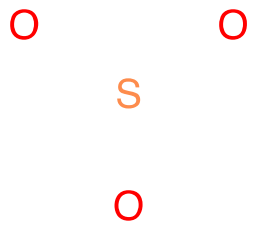
Ex. H_2SO_3

Answer: H_2SO_3

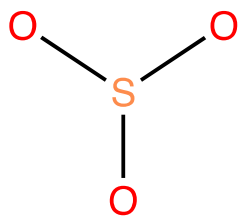
1 Count electrons

2H = 2
S = 6
3O = 3(6) = 18
Total = 26 e^-

2 Place least EN atom in center (except H and metals)

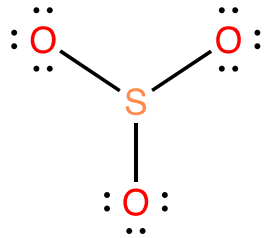


3 Form single bonds



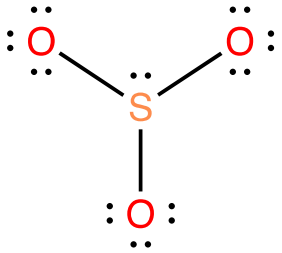
26 – 6 = 20 e^- remain

6 Complete peripheral octets.

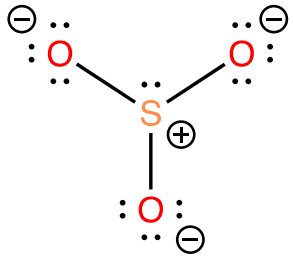


20 – 18 = 2 e^- remain

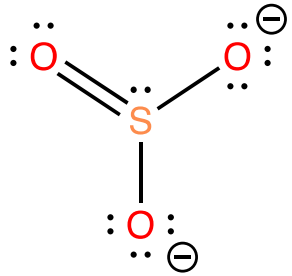
5 Fill central octet



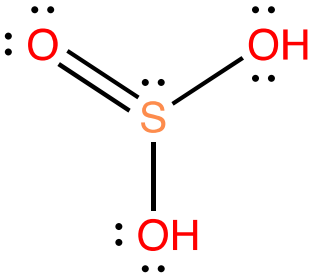
4 Calculate FC



7 Form bonds to minimize charge / complete octets



8 Add H's/metal ions to balance charge.



Ex. SO_3

Answer: SO_3

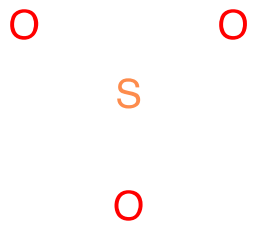
1 Count electrons

S = 6

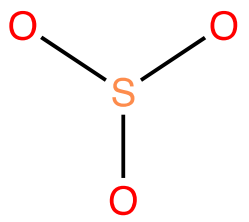
3O = 3(6) = 18

Total = 24 e^-

2 Place least EN atom in center (except H and metals)

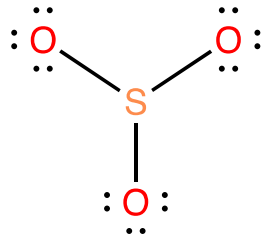


3 Form single bonds



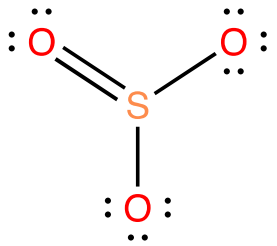
24 – 6 = 18 e^- remain

6 Complete peripheral octets.

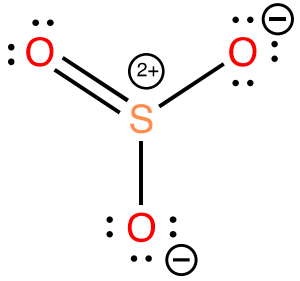


18 – 18 = 0 e^- remain

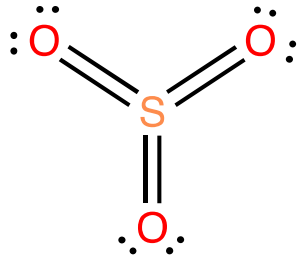
5 Fill central octet



4 Calculate FC



7 Form bonds to minimize charge / complete octets



8 Add H's/metal ions to balance charge.

Skip this step.

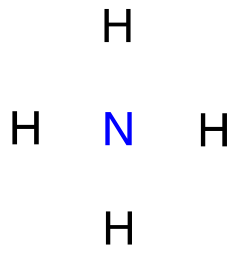
Ex. NH_4^+

Answer: NH_4^+

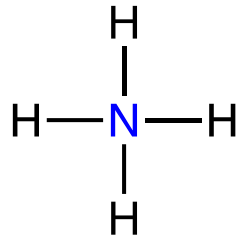
1 Count electrons

N = 5 e^-
4 H = 4(1) = 4 e^-
+ charge = –1 e^-
Total = 8 e^-

2 Place least EN atom in center



3 Form single bonds



8 – 8 = 0 e^- remain

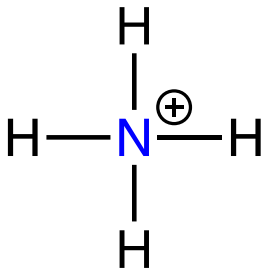
6 Complete peripheral octets.

Skip this step.

5 Fill central octet

Skip this step.

4 Calculate FC



7 Form bonds to minimize charge / complete octets

Skip this step.

8 Add H's/metal ions to balance charge.

Skip this step.

Ex. CO_2

Answer: CO_2

1 Count electrons

C = 4

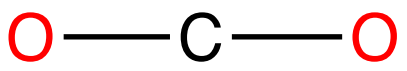
2O = 2(6) = 12

Total = 16 e^-

2 Place least EN atom in center

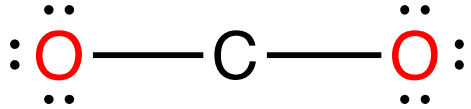
O C O

3 Form single bonds



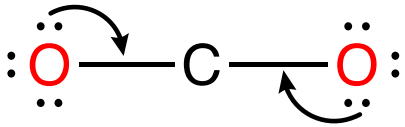
16 – 4 = 12 e^- remain

6 Complete peripheral octets.

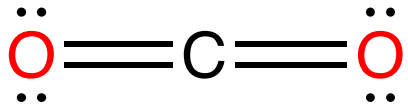


12 – 12 = 0 e^- remain

5 Fill central octet



4 Calculate FC



There is no FC

7 Form bonds to minimize charge / complete octets

Skip this step.

8 Add H's/metal ions to balance charge.

Skip this step.

Ex. NaN_3

Answer: NaN_3

1 Count electrons

Na = 1

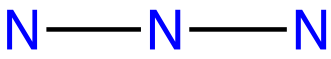
3N = 2(5) = 15

Total = 16 e^-

2 Place least EN atom in center (except H and metals)

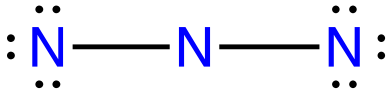
N N N

3 Form single bonds



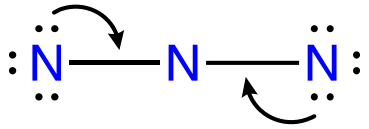
16 – 4 = 12 e^- remain

6 Complete peripheral octets.

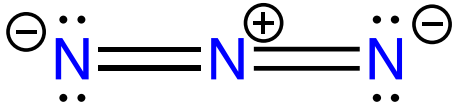


12 – 12 = 0 e^- remain

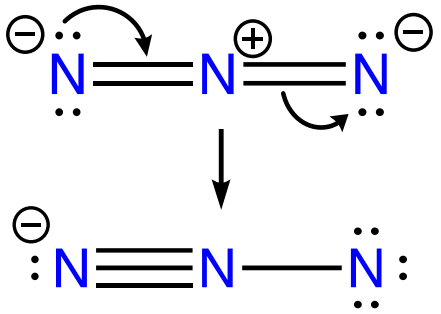
5 Fill central octet



4 Calculate FC



7 Form bonds to minimize charge / complete octets



8 Add H's/metal ions to balance charge.



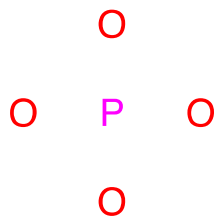
Ex. NaH_2PO_4

Answer: NaH_2PO_4

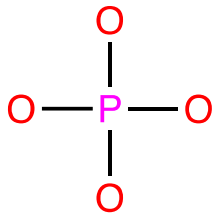
1 Count electrons

Na = 1
2H = 2(1) = 2
P = 5
4O = 4(6) = 24
Total = 32 e^-

2 Place least EN atom in center (except H and metals)

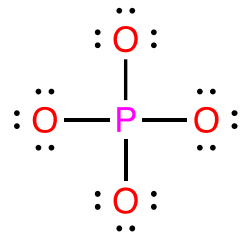


3 Form single bonds



32 – 8 = 24 e^- remain

6 Complete peripheral octets.

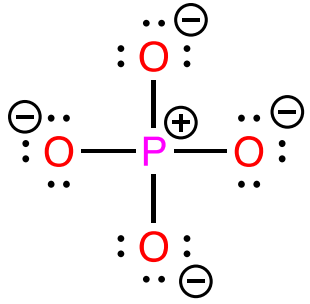


24 – 24 = 0 e^- remain

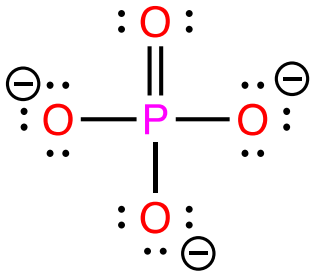
5 Fill central octet

Skip this step.

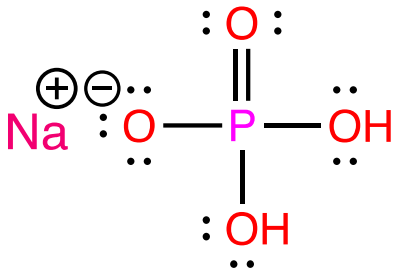
4 Calculate FC



7 Form bonds to minimize charge / complete octets

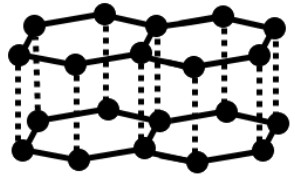


8 Add H's/metal ions to balance charge.

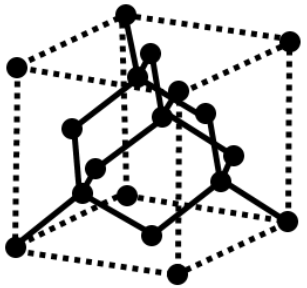


CARBON: "King of the Elements"

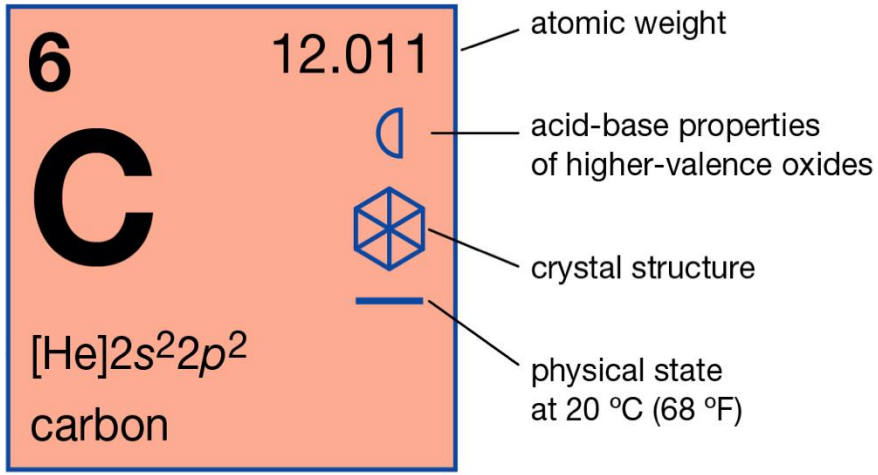
- Non-metallic tetravalent element
  - 15th most abundant element in Earth's crusts (0.025%)
  - 4th most abundant element in universe (by mass)
  - 2nd most abundant element in human body (by mass)
  - Forms more compounds than any other element (>10M)
- Allotropes
  - Graphite
  - Diamond
- Isotopes
  - C^{12} – 98.9% – most common form, used to define 1 mole
  - C^{13} – 1.1% – stable isotope, used for NMR
  - C^{14} – 0.0000000001% – radioisotope, used for carbon dating



Graphite Structure



Diamond Structure

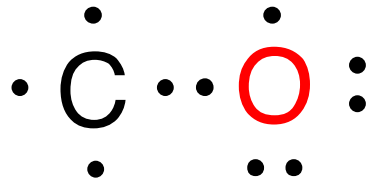




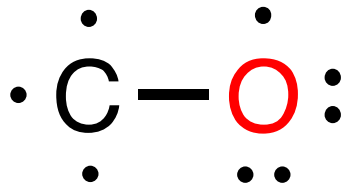
Ex. CH_2O

Answer: CH_2O

1. Draw Lewis Dot structures of the atoms (excluding H's) in the order presented in the molecular formula.

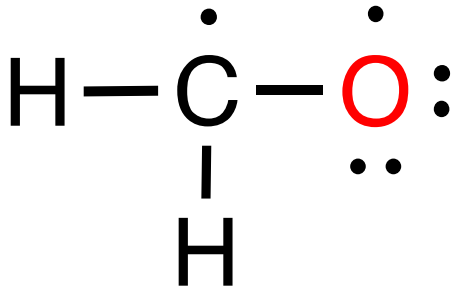


Connect atoms via single bond.



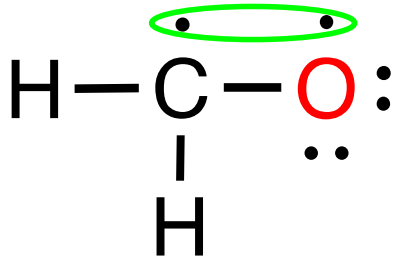
2. Connect all H atoms to the atom that precedes them in the molecular formula.

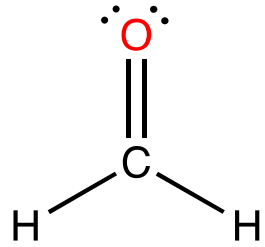
In this case, we will bond them to C.



3. Complete octets for all atoms that follow the octet rule.

Using this method, we can do this by forming double/triple bonds.





Formaldehyde

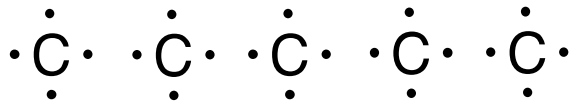
Ex. $CH_3CHCHCH_2CH_3$

What type of bonding is present in this molecule?

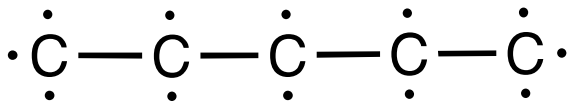
a) Covalent
b) Ionic
c) Hydrogen
d) Metallic

Answer: $CH_3CHCHCH_2CH_3$

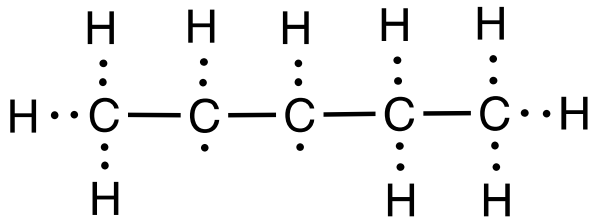
1 Write carbons out according to molecular formula.



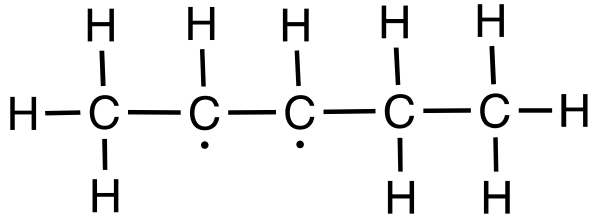
Connect carbons with single bonds.



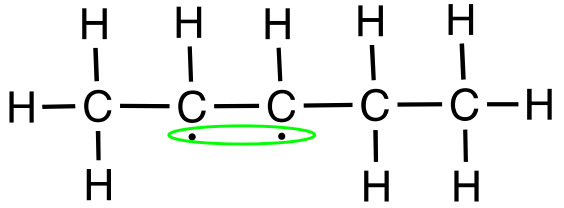
2 Connect all H atoms to the atom that precedes them in the molecular formula.

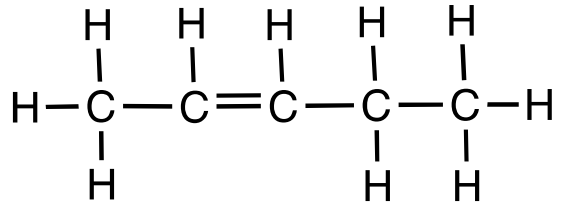


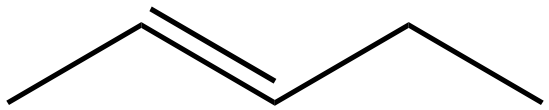
Connect hydrogens to carbon with single bonds.



3 If needed, form pi bonds to complete octets.



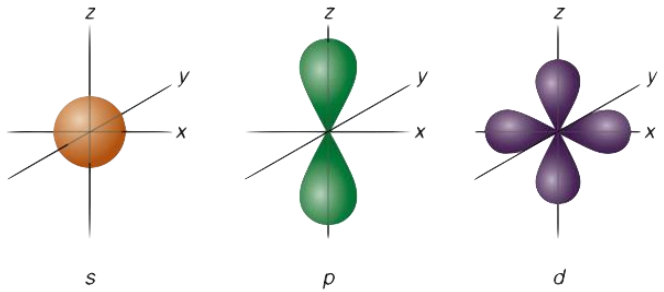




Atomic (AO) & Molecular Orbital (MO) Theory

AO Theory describes the probable locations of electrons in an atom.

- Types: s, p, d, f
- Each orbital can contain 2 electrons, with opposing spins
- Require *hybridization* to form bonds: sp, sp^2, sp^3, *etc*.

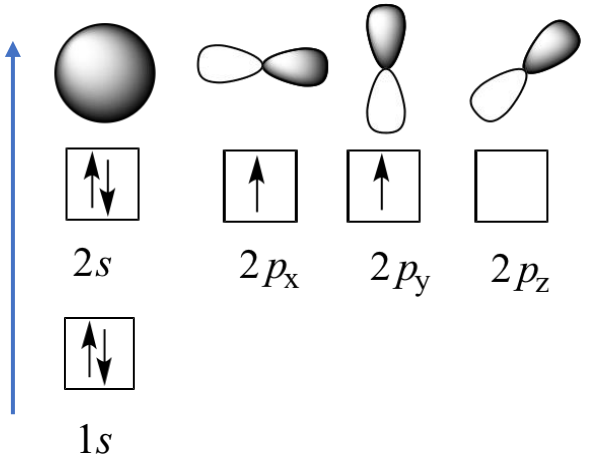


MO Theory describes the probable locations of electrons in a molecule. When atoms join to form molecules, atomic orbitals overlap and combine to become molecular orbitals.

Rule: The # of newly formed MOs equals the number of combined AOs.

Types: bonding, anti-bonding (*)

Ex. Carbon



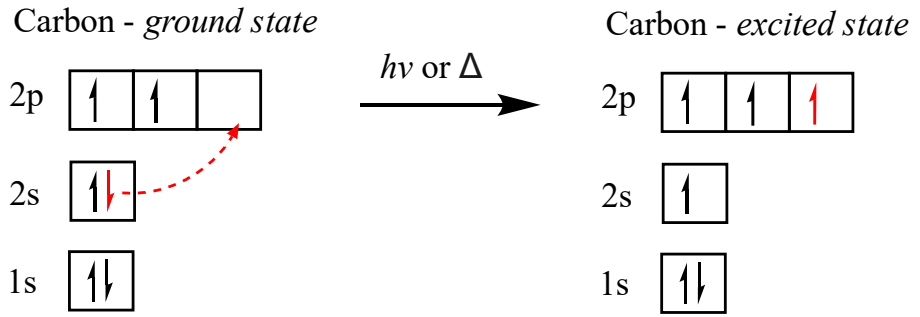
Q: How many bonds does C prefer to make?

a) 1
b) 2
c) 4
d) 6

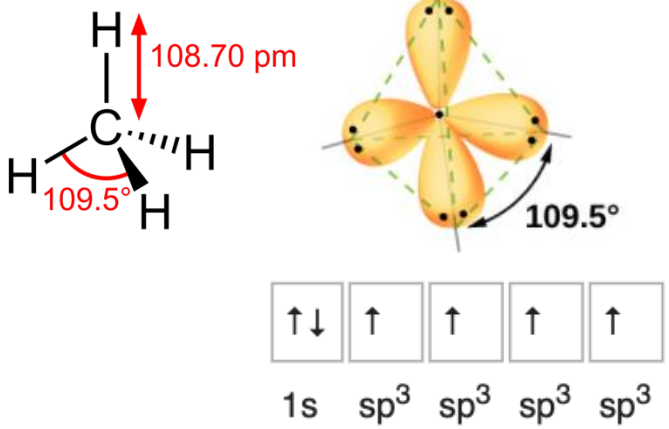
Orbital Hybridization

Hybridization of atomic orbitals occurs accommodate new electron density due to newly formed bonds. Different bonding patterns will result in different types of hybridization.

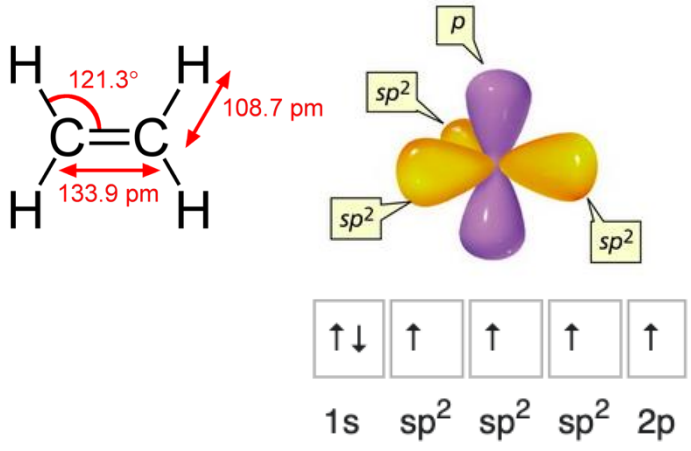
To form most carbon-containing compounds, a carbon atom's ground state must be "excited" from the s orbital to the empty p orbital. The energy released due to forming 2 more bonds compensates for excitation E required.



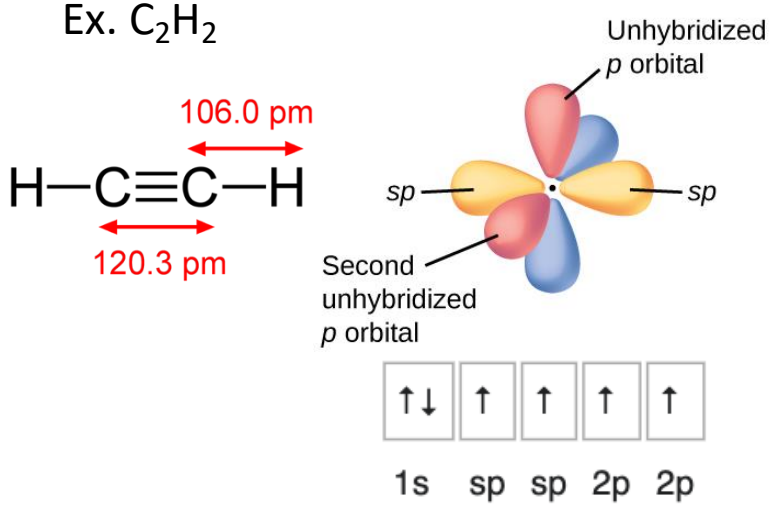
Ex. CH_4



Ex. C_2H_4



Ex. C_2H_2



Identifying Hybrid States

1. Find the **STERIC #**

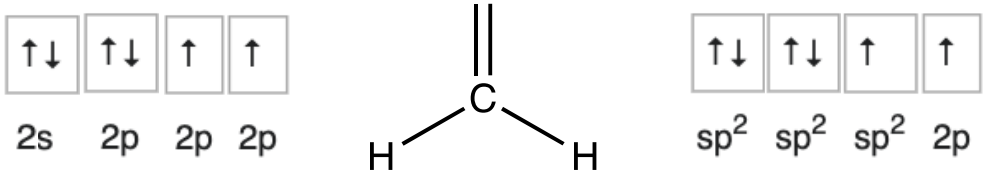
   Steric # = number of **atoms bonded** to and **lone pairs** on the central atom

2. You must hybridize the same number of atomic orbitals

Ex. Formaldehyde

CARBON	OXYGEN
Steric # = 3	Steric # = 3
• 3 bonded atoms • 0 lone pairs	• 1 bonded atom • 2 lone pairs

Hybridize 3 AO: 2s, 2p, 2p ⟶ sp^2, sp^2, sp^2

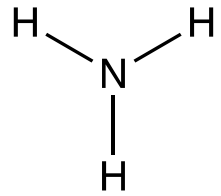


Q: Which atom is represented in the orbital diagram?

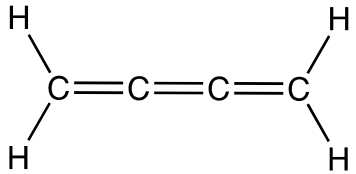
Ex. Ammonia

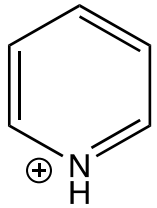
NITROGEN

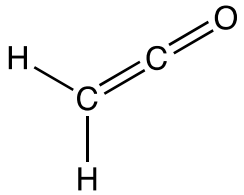
Steric # = ?



Hybridization Practice





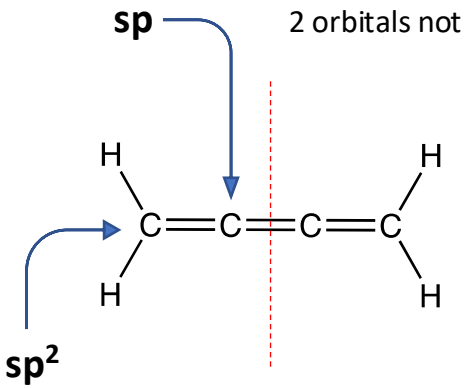




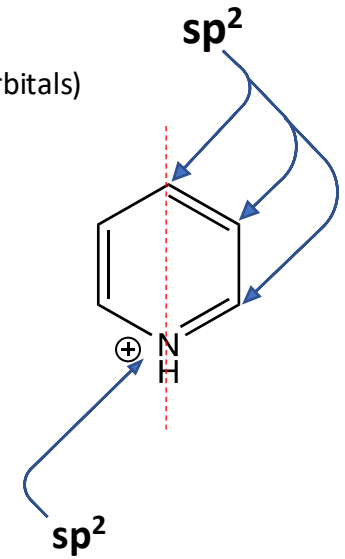
Answer: **Hybridization Practice**

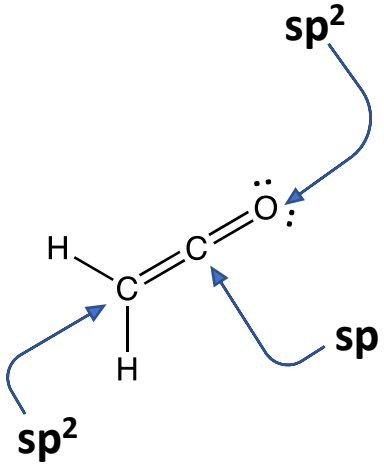
The first two molecules are symmetrical, so we only need to evaluate ½ of the molecule.

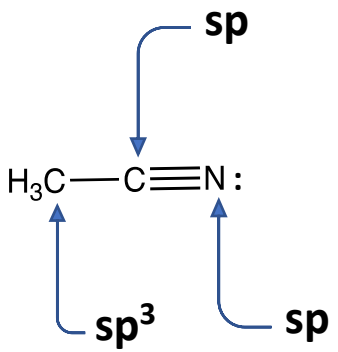
Steric # is 2 = 2 orbitals hybridized (1 s, 1 p)
2 orbitals not hybridized (2 p orbitals)



Steric # is 3 = 3 orbitals hybridized (1 s, 2 p)
1 orbital not hybridized (1 p orbital)







Steric # is 4 = 4 orbitals hybridized (1 s, 3 p)
0 orbital not hybridized