

## Lewis Structures & VSEPR

## Lewis Structure

- Lewis Structures – shows how the \_\_\_\_\_ are arranged among the atoms of a molecule
- There are rules for Lewis Structures that are based on the formation of a \_\_\_\_\_
- Atoms want to achieve a \_\_\_\_\_ configuration

## Octet & Duet Rules

- Octet Rule – atoms want to have \_\_\_\_\_ valence electrons
- Duet Rule – H is the exception. It wants to be like He & is stable with only \_\_\_\_\_ valence electrons

## Steps for drawing Lewis Structures

- Sketch a simple structure with a central atom and all attached atoms
- Add up all of the valence electrons for each individual atom
  - If you are drawing a Lewis structure for a negative ion add that many electrons to create the charge
  - If you are drawing a Lewis structure for a positive ion subtract that many electrons to create the charge

### Steps for drawing Lewis Structures

- Subtract 2 electrons for each bond drawn
- Complete the octet on the central atom & subtract those electrons
- Complete the octet on the surrounding atoms & subtract those electrons
- Get your final number
  - If 0 → you are done!
  - If + → add that many electrons to the central atom
  - If - → need to form multiple bonds to take away that many electrons

### Bond Types

- \_\_\_\_\_ bonds ( $\sigma$ ) – single covalent bond
- \_\_\_\_\_ bonds ( $\Pi$ ) – occur when multiple bonds are formed
- Single bond – sigma
- Double bond – 1 sigma & 1 pi
- Triple bond – 1 sigma & 2 pi

### Bond length & Strength

- As the number of bonds increases, the bond length \_\_\_\_\_
- The shorter the bond, the \_\_\_\_\_ the bond

### Examples

- $\text{CCl}_4$

### Examples

- HF

### Examples

- NH<sub>3</sub>

### Examples

- NO<sup>+</sup>

### Exceptions to the octet rule

- Sometimes the central atom violates the octet rule and has more or less than 8 valence electrons
- Keep using the same rules to draw Lewis Structures

Examples

- SF<sub>4</sub>

Examples

- ICl<sub>3</sub>

Examples

- XeF<sub>4</sub>

Examples

- ICl<sup>-</sup>

## Resonance

- When more than one Lewis Structure can be written for a particular molecule
- \_\_\_\_\_ – all possible Lewis structures that could be formed
- The actual structure is the \_\_\_\_\_ of all of the structures
- You MUST show all structures!

## Examples

- $\text{SO}_3$

## Examples

- $\text{NO}_2^-$

## Examples

- $\text{NO}_3^-$

## Formal Charge

- Formal charges can be used in 1 of 2 ways...
  - Suggest where the \_\_\_\_\_ are
  - Help select the most \_\_\_\_\_ structure from a set of resonance structures

## 1 - Suggest where the charges are

- Formal charge =  


## Example

- Calculate the formal charge on each element in the carbonate ion
- $\text{CO}_3^{2-}$

## 2 - Help select the most plausible structure from a set of resonance structures

- When choosing the most likely resonance structure
  - Most likely – All formal charges are \_\_\_\_\_
  - Next likely – All formal charges add up to \_\_\_\_\_
  - Next likely – Formal charges are closest to \_\_\_\_\_
  - Next likely – \_\_\_\_\_ charge is on most electronegative atom

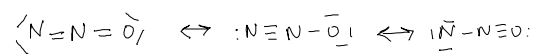
### Example

- Which of the following resonance structures is most likely for CH<sub>2</sub>O and why?



### Another Example

- Which is the most likely structure for N<sub>2</sub>O?



### VSEPR Theory

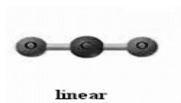
- VSEPR Theory  
\_\_\_\_\_
- The structure around a given atom is determined by minimizing the  
\_\_\_\_\_
- The electrons and elements bonded to the central atom want to be  
\_\_\_\_\_ as possible

### VSEPR Steps

1. Draw the Lewis structure for the molecule
2. Count the total number of things that are around the central atom to determine the **electron pair geometry**
3. Imagine that the lone pairs of electrons are invisible and describe the molecular shape

## 2 Electron Pairs

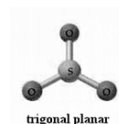
- If there are 2 things attached to the central atom, the shape is linear



- Bond angle =  $180^\circ$

## 3 Electron Pairs

- If there are 3 electron pairs the shape will be trigonal planar



- Bond angle =  $120^\circ$

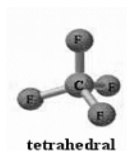
## 3 electron pairs

- Now imagine that you have 3 electron pairs, but one is just a lone pair (invisible) what would it look like then?



## 4 electron pairs

- If there are 4 electron pairs, the shape will be tetrahedral



- Bond angle =  $109.5^\circ$

### 4 electron pairs

- What if 1 of the electron pairs is a lone pair (invisible)? What would it look like then?

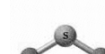


trigonal pyramidal

- Trigonal Pyramidal

### 4 electron pairs

- What if there are 2 lone pairs (invisible)? What would it look like then?



bent

- bent

### 5 electron pairs

- If there are 5 electron pairs the shape will be Trigonal Bipyramidal

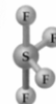


trigonal bipyramidal

- Bond angles =  $90^\circ$  &  $120^\circ$

### 5 electron pairs

- What if there is 1 lone pair (invisible)

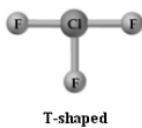


seesaw

- Seesaw

### 5 electron pairs

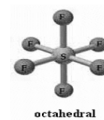
- What is there are 2 lone pairs (invisible)



- T-shaped

### 6 electron pairs

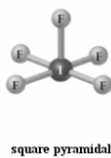
- If there are 6 electron pairs the shape will be octahedral



- Bond angle =  $90^\circ$

### 6 electron pairs

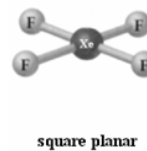
- What if there is 1 lone pair (invisible)?



- Square pyramidal

### 6 electron pairs

- What if there are 2 lone pairs (invisible)



- Square planar