

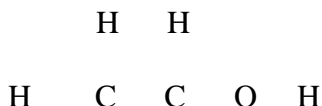
## STEPS FOR DRAWING LEWIS STRUCTURES

1. Given the molecular formula, add up the valence electrons in the atoms.

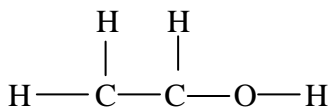
Example:  $C_2H_4O$                        $2(4) + 4(1) + 1(6) = 18$

*Note: In the case of an ion if it has a "+" charge subtract that number from the total valence electrons. If it has a "-" charge, add that number to the total valence electrons. Example:  $NH_4^+$  has 8 valence electrons ( $5 + 4 - 1$ ) and  $NO_3^-$  has 24 valence electrons ( $5 + 18 + 1$ ).*

2. Arrange the atoms with one or two central atoms and the others around it.
- H atoms are always on the outside because they always only have 1 bond.
  - O atoms are usually on the outside (not central) unless combined with H like  $H_2O$
  - C atoms are always a central atom

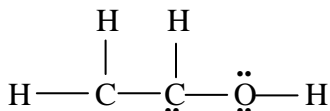


3. Connect the atoms with lines (bonds) to the central atom.



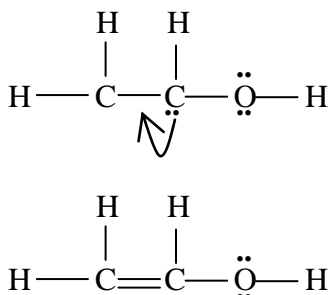
4. Subtract the bonded electrons (2 per line) from the total valence electrons:  $18 - 6(2) = 6$

5. Place these extra electrons around outside atoms until they have a complete octet, then put the remainder on inside atoms. In this case, put the 4 electrons around the oxygen atom and two electrons on one carbon atom. Always place electrons around atoms in pairs.



6. Count electrons around each atom to be sure there is a complete octet, except for H which only has a duet. One of the carbon atoms does not have a complete octet.

7. If one or more of the central atoms do not have a complete octet then use some of the non-bonding electron pairs to make a bonding pair (i.e., make multiple bonds) so that that all atoms end up with complete octets.




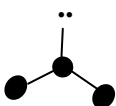
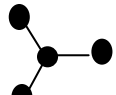

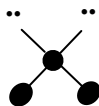
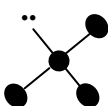
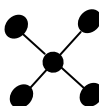


8. Recount electrons around atoms to be sure all have a complete octet. Also count bonding and non-bonding electrons and make sure they equal the total number of valence electrons you started with.  
All atoms have complete octets.

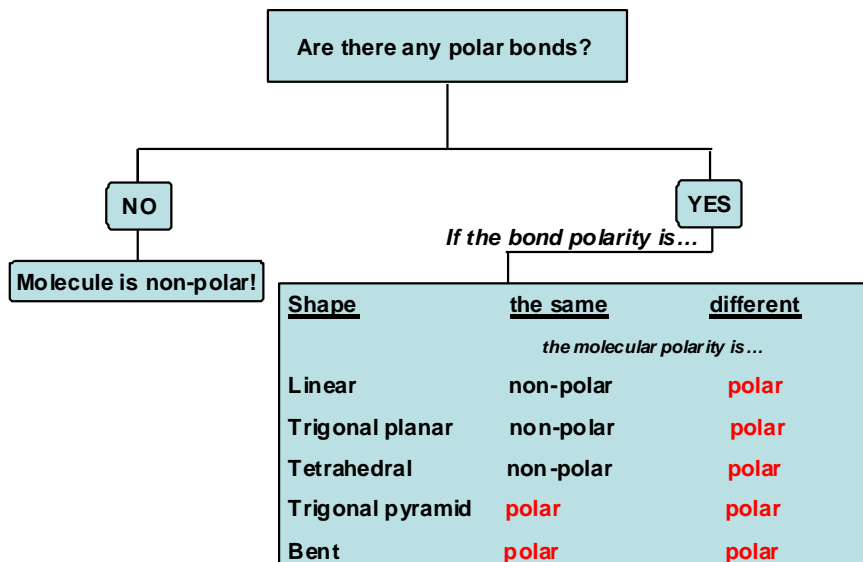
$$\begin{array}{r} \text{bonding electrons} = 14 \\ \text{lone pairs} = \quad \underline{4} \\ \text{total} = \quad \quad 18, \text{ the same as the valence electrons.} \end{array}$$

9. The Lewis structure is finished and it is correct!
10. Some other considerations are the usual bonding patterns for various elements:
- H always has 1 bond and never any lone pairs. It only has a duet of electrons.
  - C always has 4 bonds. They can be 4 single, 1 double and 2 single, 1 triple and 1 single. Carbon is ALWAYS a central atom and NEVER has any lone pairs.
  - F always has 1 bond and 3 lone pairs and is ALWAYS an outside atom.
  - Cl, Br and I usually have 1 bond and 3 lone pairs but can have other bonding arrangements depending on the compound.
  - O usually has 2 bonds and 2 lone pairs. The two bonds can be 2 single bonds or 1 double bond. It is usually an outside atom.
  - N usually has 3 bonds and 1 lone pair. The three bonds can be 3\2 single bonds or 1 double and 1 single bond or 1 triple bond. It is usually an outside atom.

## SUMMARY OF VSEPR MODEL

TOTAL ELECTRON GROUPS <i>(Electron Group Geometry)</i>			
 2 linear Bond angle: 180°	 2 linear 180°  CO <sub>2</sub>		
 3 trigonal planar Bond angle: 120°	 2 bent 120°  NO <sub>2</sub> <sup>-</sup>	 3 trigonal planar 120°  BF <sub>3</sub>	
 4 tetrahedral Bond angle: 109.5°	 2 bent 109.5°  H <sub>2</sub> O	 3 trigonal pyramid 109.5°  NH <sub>3</sub>	 4 tetrahedral 109.5°  CH <sub>4</sub>
	2 bonding groups	3 bonding groups	4 bonding groups
	BONDING ELECTRON GROUPS <i>(Molecular Geometry)</i>		

# Molecular Polarity



## Examples of non-polar molecules

Type	Diagram	Cancellation of Polar Bonds	Example
Linear molecules with <u>two identical bonds</u>	$B-A-B$	$\leftarrow + \rightarrow$	$CO_2$
Trigonal planar molecules with <u>three identical bonds</u>			$SO_3$
Tetrahedral molecules with <u>four identical bonds</u>			$CH_4$

## Examples of polar molecules

Shape	polar bonds do not cancel	EXAMPLE
Linear molecules with different bonds	$\leftarrow + \rightarrow$	$O=C=S$
Trigonal planar molecules with different bonds		
Tetrahedral molecules with different bonds		
Trigonal pyramid molecules with identical bonds		$NH_3$
Bent molecules with identical bonds		$H_2O$