

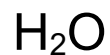
## Formal Charges of Lewis Structures

In order to determine the best Lewis Structure, the formal charges of each element must be calculated, an ideal structure has an overall formal charge of 0 with each element having a formal charge of 0.

## Formula for Formal Charges

Valence electrons - bonds - lone electrons

Draw the Lewis structure and determine the formal charges of each



Using the idea of formal charges, draw the resonance structures of  $\text{CO}_2$  and determine which is the best.

## Exceptions to the Octet Rule in Lewis structures

### 1) Species with odd number of electrons

-free radicals

example: NO (nitrogen monoxide)

### 2) Incomplete octets

-beryllium, aluminum, boron

Example:  $\text{BH}_3$  (Borane, boron trihydride)

Example:  $\text{BF}_3$  (boron trifluoride)

### 3) Expanded Octet

-only elements with a minimum of  $n=3$  can have expanded octets

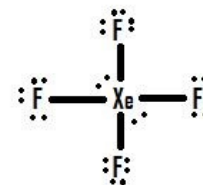
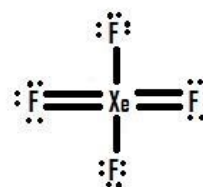
Example:  $\text{SF}_6$  (sulfur hexafluoride)

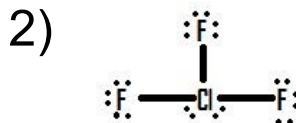
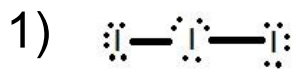
Example:  $\text{SO}_4^{2-}$  (sulfate)

You must look at the resonance structures and determine which has better formal charges.

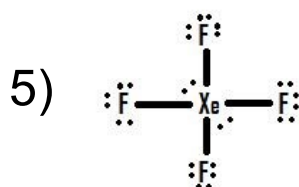
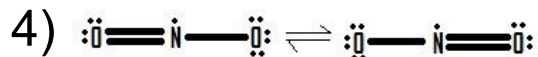
#### Practice

- 1) Draw the Lewis structure for the molecule  $\text{I}_3^-$
- 2) Draw the molecule  $\text{ClF}_3$
- 3) The central atom for an expanded octet must have an atomic number larger than what?
- 4) Draw the Lewis structure for  $\text{NO}_2$
- 5) Which Lewis structure is more likely?





3) 10 (sodium or higher)

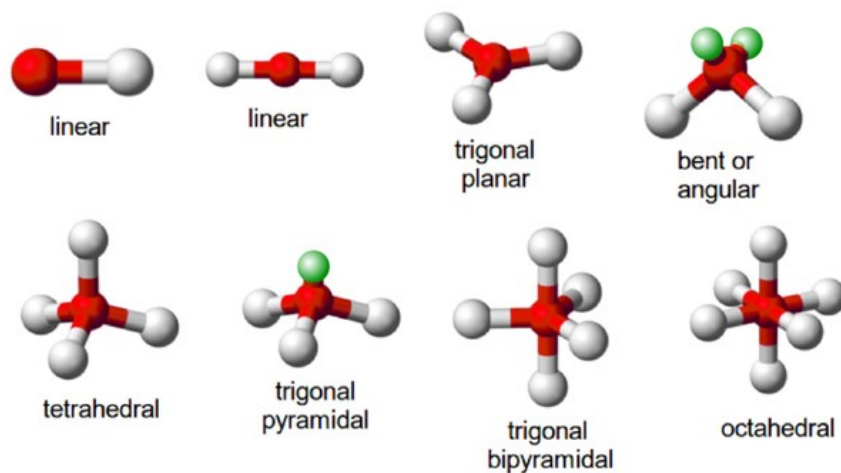


## VSEPR Theory

### Valence Shell Electron Pair Repulsion Theory

This is the theory that gives us molecular geometry. The general idea is atoms want to be as far from each other as they can, but lone electron pairs have a stronger repulsion than bonded atoms.

## VSEPR Shapes



Molecule Type	Shape <sup>[1][2]</sup>	Electron arrangement <sup>[1][2]</sup>	Geometry <sup>[1][2]</sup>	Examples
$AX_2E_0$	Linear			$BeCl_2$ <sup>[1]</sup> $HgCl_2$ <sup>[1]</sup> $CO_2$ <sup>[1]</sup>
$AX_2E_1$	Bent			$NO_2$ <sup>[1]</sup> $SO_2$ <sup>[1][2]</sup> $O_3$ <sup>[1]</sup> $CCl_2$
$AX_2E_2$	Bent			$H_2O$ <sup>[1][2]</sup> $OF_2$ <sup>[1]</sup>
$AX_2E_3$	Linear			$XeF_2$ <sup>[1][2]</sup> $I_3$ <sup>[1][8]</sup> $XeCl_2$
$AX_3E_0$	Trigonal planar			$BF_3$ <sup>[1][2]</sup> $CO_3^{2-}$ <sup>[1][8]</sup> $NO_3^-$ <sup>[1]</sup> $SO_3$ <sup>[1]</sup>
$AX_3E_1$	Trigonal pyramidal			$NH_3$ <sup>[1][2]</sup> $PCl_3$ <sup>[2]</sup>
$AX_3E_2$	T-shaped			$ClF_3$ <sup>[1][2]</sup> $BrF_3$ <sup>[2]</sup>
$AX_4E_0$	Tetrahedral			$CH_4$ <sup>[1][2]</sup> $PO_4^{3-}$ $SO_4^{2-}$ <sup>[1]</sup> $ClO_4^-$ <sup>[1]</sup> $XeO_4$ <sup>[2]</sup>
$AX_4E_1$	Seesaw (also called disphenoidal)			$SF_4$ <sup>[1][2][2]</sup>
$AX_4E_2$	Square planar			$XeF_4$ <sup>[1][2]</sup>
$AX_5E_0$	Trigonal bipyramidal			$PCl_5$ <sup>[1][2]</sup>
$AX_5E_1$	Square pyramidal			$ClF_5$ <sup>[2]</sup> $BrF_5$ <sup>[1][2]</sup> $XeOF_4$ <sup>[1]</sup>
$AX_6E_0$	Octahedral			$SF_6$ <sup>[1][2]</sup> $WCl_6$ <sup>[2]</sup>

## Bond angles

linear =  $180^\circ$

trigonal planar =  $120^\circ$

bent / angular =  $<120^\circ$

$\text{H}_2\text{O}$  is  $104.5^\circ$ ,  $\text{SO}_2$  is  $119^\circ$

tetrahedral =  $109.5^\circ$

trigonal pyramidal =  $107^\circ$

trigonal bipyramidal =  $90^\circ$ ,  $120^\circ$ ,  $180^\circ$

octahedral =  $90^\circ$ ,  $180^\circ$

## Polar Bonds

A type of covalent bond between two atoms in which electrons are shared unequally.

Because of this, one end of the molecule has a slightly negative charge and the other a slightly positive charge.

Polar Bonds depend on geometric structure of the atom and the difference in electronegativity of the elements in the compounds.

Water is a polar molecule.

Is Carbon Dioxide a polar molecule?



Is hydrogen sulfide a polar molecule?

How about methane? (CH<sub>4</sub>)