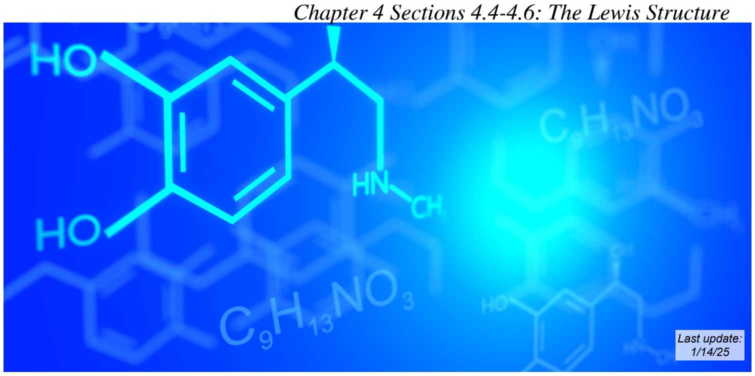
Chemistry 151: Basic Chemistry

Chapter 4 Sections 4.4-4.6: The Lewis Structure



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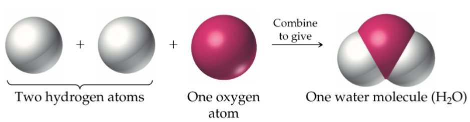
Covalent Bonds

A ***covalent bond*** is a bond formed by sharing electrons between atoms.

A ***molecule*** is a group of atoms held together by covalent bonds.

Nonmetals form covalent bonds with nonmetals. They reach the Noble Gas configuration by *sharing* an appropriate number of electrons.

A water molecule results when two hydrogen atoms and one oxygen atom are covalently bonded:



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Test Yourself

Are these compounds bonded through ionic or covalent bonding?

PCl_5
Na_2O
SO_3
$CaSO_3$
SbAs

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***Nomenclature** of covalent compounds different from ionic compounds; important to know the difference*
ex: Na_2O = sodium oxide, CO_2 = carbon dioxide

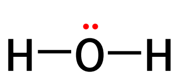
Covalent Bonds and the Periodic Table

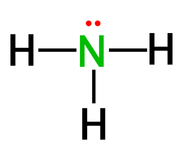
Covalent bonds can form between nonmetals, making possible a vast number of molecular compounds.

Water, H_2O, consists of two hydrogen atoms joined by covalent bonds to one oxygen atom with two lone pairs on oxygen

Ammonia, NH_3, consists of three hydrogen atoms joined by covalent bonds to one nitrogen atom with one lone pair on nitrogen

In most covalent molecules, each atom shares enough electrons to achieve a noble gas configuration.





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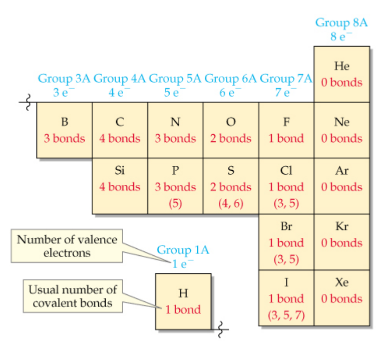

Note that **number of valence electrons equals the group number** for the element!

Note: For P, S, Cl, and other elements in the third period and below, the number of covalent bonds may vary, as indicated by the numbers shown in parentheses.

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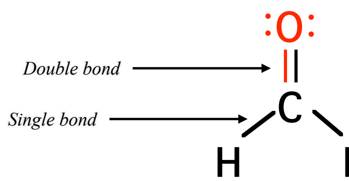
Multiple Covalent Bonds

Single bond: A bond formed by sharing two electrons or one pair – represented by a single line between the atoms.

Double bond: A bond formed by sharing four electrons or two pairs – represented by two lines (=) between the atoms.



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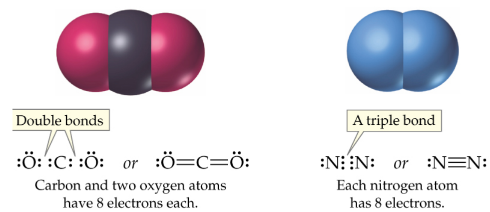
Triple bond: A bond formed by sharing six electrons or three pairs– represented by three lines (≡) between the atoms.



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Molecular Formulas and Lewis Structures

Molecular Formula: A formula that shows the number and kind of atoms in a molecule

Structural formula: Molecular representation that shows the connections among atoms by using lines to represent covalent bonds

Example for water:

H_2O = molecular formula

H-O-H = structural formula

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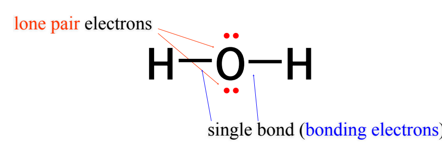
Lewis structure: Molecular representation showing both the connections among atoms *and* the locations of lone pair valence electrons.

A ***lone pair*** is a pair of electrons not used for bonding.

Lewis structure example for water:



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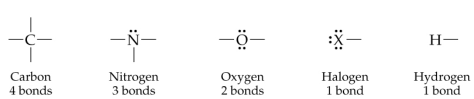
Drawing Lewis Structure

To draw a Lewis structure, you need to know the connections among atoms.

Knowing *common bonding patterns* simplifies writing Lewis structure.



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Building a Lewis Dot Structure

Ammonia, NH_3

1. **Count valence electrons**

   H = 1 and N = 5

   Total = (3 x 1) + 5

   = 8 electrons *or*

   4 pairs of electrons

2. **Decide on the central atom; never H.**

   Central atom is atom of lowest affinity for electrons.

   Therefore, N is central

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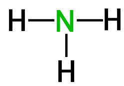
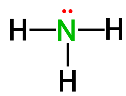
Building a Lewis Dot Structure

3. Form a sigma bond *(single bond)* between the central atom and surrounding atoms.



4. Remaining electrons form LONE PAIRS to complete octet as needed.



Unshared electron pairs ("lone pairs") take up more volume than shared electron pairs ("bonding pairs")

3 BOND PAIRS and 1 LONE PAIR. *Note that N has a share in 4 pairs (8 electrons), while H shares 1 pair.*

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Lewis structure rules:

1: Decide on a central atom (usually listed first in formula, never H) and find the total number of valence electrons in molecule or ion

2: Draw a line between each pair of connected atoms to represent a covalent bond

3: Add lone pairs so that each *peripheral* atom (except H) gets an octet

4: Place all remaining electrons on the central atom

5: If central atom does not have an octet, take lone pair(s) from neighboring atom(s) and form multiple bond(s) to the central atom

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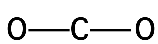
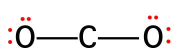
Carbon Dioxide, CO_2

1. Central atom = ______
2. Valence electrons = __ or __ pairs
3. Form sigma bonds.



This leaves 6 pairs.

4. Place lone pairs on outer atoms.



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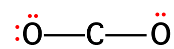
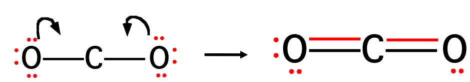
Carbon Dioxide, CO_2

4. Place lone pairs on outer atoms.



5. So that C has an octet, we shall form DOUBLE BONDS between C and O.



The second bonding pair forms a pi (π) bond.

This is the Lewis structure for CO_2

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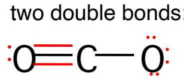
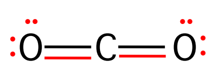
Resonance Structures of CO_2

Could have written CO_2 with a triple bond instead of two double bonds:

 *or* 

Energetically similar, they are called resonance structures of CO_2.

Provides additional stability to the molecule.

Practice, practice, practice!

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Shape of Molecules

Molecular shapes can be predicted by noting how many bonds and electron pairs surround individual atoms and applying ***valence-shell electron-pair repulsion*** *(VSEPR)* theory.

Basic idea of VSEPR: negatively charged electron clouds in bonds and lone pairs repel each other, keeping them as far apart as possible

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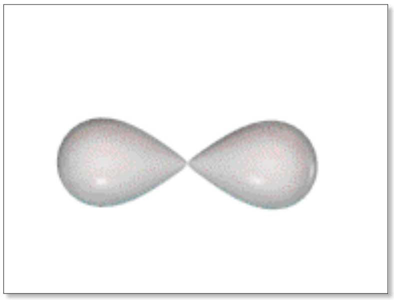
VSEPR



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VSEPR Rules

To apply VSEPR theory:

1: Draw the Lewis structure of the molecule and identify the central atom
2: Count the number of electron charge clouds (lone *and* bonding pairs) surrounding the central atom.
3: Predict molecular shape by assuming that clouds orient so they are as far away from one another as possible.

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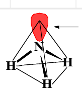
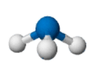
VSEPR Shape Predictor Table

Clouds	Bonds	Lone Pairs	Electron Pair Geometry (EPG)	Molecular Geometry (MG)	Angles	Notes
2	2	0	linear	linear	180	
3	3	0	trigonal planar	trigonal planar	120	
3	2	1	trigonal planar	bent	120	*always polar*
4	4	0	tetrahedral	tetrahedral	109	
4	3	1	tetrahedral	trigonal pyramid	109	*always polar*
4	2	2	tetrahedral	bent	109	*always polar*



tetrahedral EPG, trigonal pyramid Molecular shape



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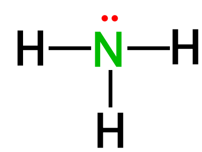
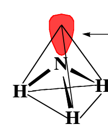
Structure Determination by VSEPR

Ammonia, NH_3

1. **Draw electron dot structure**
2. **Count BPs and LPs = 4**
3. **The 4 electron pairs are at the corners of a tetrahedron.**





lone pair of electrons in tetrahedral position

The ELECTRON PAIR GEOMETRY is tetrahedral.

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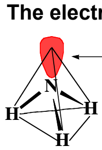
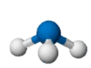
Structure Determination by VSEPR

Ammonia, NH_3

The electron pair geometry is tetrahedral.



lone pair of electrons in tetrahedral position



The MOLECULAR GEOMETRY - the positions of the atoms - is TRIGONAL PYRAMID

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Test Yourself

Describe the Lewis structure, electron pair geometry and molecular shape of methane, CH_4.

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Polar Covalent Bonds and Electronegativity

Electrons in a covalent bond occupy the region between the bonded atoms.

If atoms in bond identical (H_2, Cl_2, etc.) electrons are attracted equally to both atoms and are shared equally (**nonpolar**)

If atoms in bond different (HCl, HF, etc.) electrons may be attracted more strongly by one atom than by the other and are shared unequally.

Such bonds are known as **polar *covalent bonds.*** *Most bonds are polar!*

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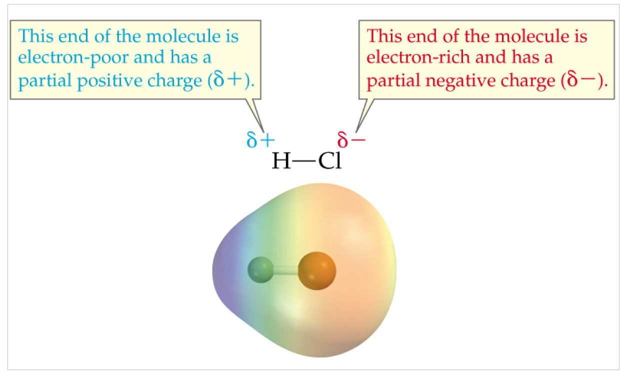

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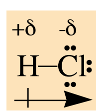
In HCl, electrons spend more time near Cl than H. Although molecule is overall neutral, the chlorine is more negative than the hydrogen, resulting in partial charges on the atoms.

Partial charges represented by placing δ- on the more negative atom and δ+ on the more positive atom.

Ability of an atom to attract electrons is called the atom's ***electronegativity***.

Fluorine, the most electronegative element, assigned a value of 4, and less electronegative atoms assigned lower values



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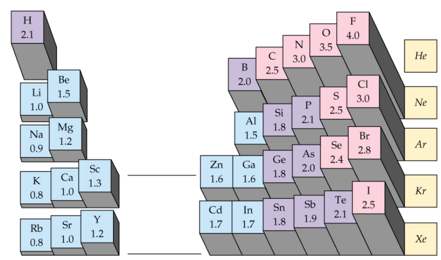

Electronegativities and the periodic table

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Polar Molecules

Entire molecule can be polar *if* electrons are attracted more strongly to one part of the molecule than another.

Molecule's polarity is due to the sum of all individual bond polarities *and* lone-pair contributions in the molecule.

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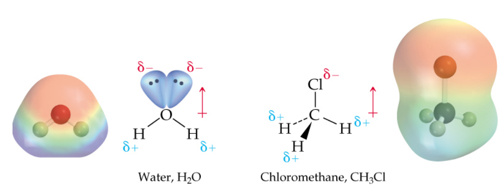
Molecular polarity is represented by an arrow pointing at the negative end and is crossed at the positive end to resemble a positive sign.



*Asymmetric molecules are **polar***

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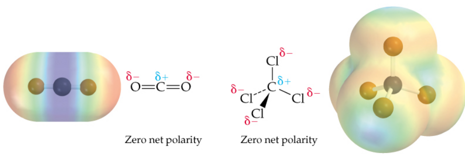
Molecular polarity depends on the shape of the molecule as well as the presence of polar covalent bonds and lone-pairs



Symmetric molecules are ***nonpolar***

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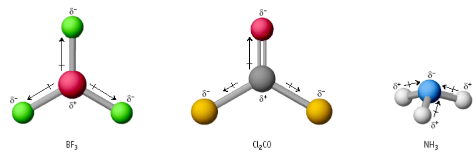
Test Yourself

Are BF_3, Cl_2CO, and NH_3 polar or nonpolar?



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End of Chapter 4 Sections 4.4 - 4.6