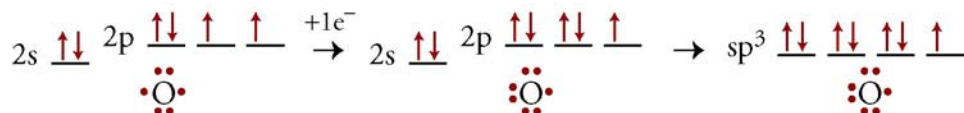


- c. O - 1 bond, 3 lone pairs, and a -1 formal charge.

Only the highest energy electrons participate in bonding. Oxygen is in group 6A, so it has six valence electrons per atom. If it gains one electron, it will have a total of seven. The 2s and the three 2p orbitals blend to form four equivalent sp^3 hybrid orbitals.

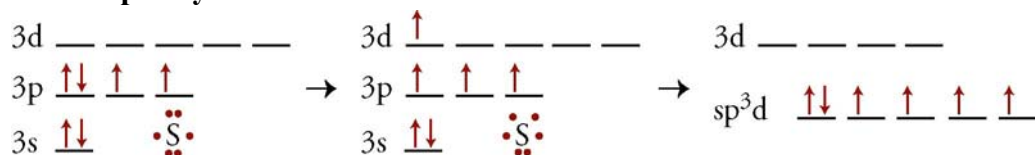


Covalent bonds form to in order to pair unpaired electrons. The one unpaired electron leads to one bond, and the three pairs of electrons give oxygen atoms with an extra electron three lone pairs. The gained electron leads to the -1 formal charge.



- d. S - 4 bonds, one lone pair, and no formal charge.

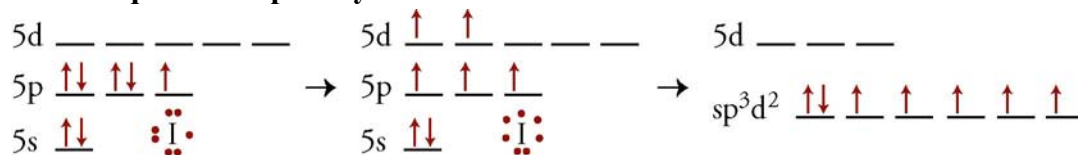
Only the highest energy electrons participate in bonding. Sulfur is in group 6A, so it has six valence electrons per atom. It is as if one electron is promoted from the 3p orbital to an empty 3d orbital. The 3s, the three 3p orbitals, and one 3d orbital blend to form 5 equivalent sp^3d hybrid orbitals.



Covalent bonds form to in order to pair unpaired electrons. The four unpaired electrons lead to four bonds, and the one pair of electrons is the lone pair.

- e. I - 5 bonds, one lone pair, and no formal charge.

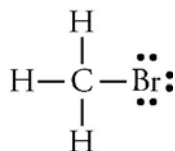
Only the highest energy electrons participate in bonding. Iodine is in group 7A, so it has seven valence electrons per atom. It is as if two electrons are promoted from two 5p orbitals to two empty 5d orbitals. The 5s, the three 5p orbitals, and two 5d orbitals blend to form 6 equivalent sp^3d^2 hybrid orbitals.



Covalent bonds form to in order to pair unpaired electrons. The five unpaired electrons lead to five bonds, and the one pair of electrons is the lone pair.

Exercise 8,9.2 – Drawing Lewis Structures: Draw a reasonable Lewis structure for each of the following. If the structure has resonance, draw all the reasonable resonance structures and the resonance hybrid.

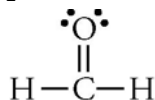
a. CH_3Br



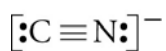
b. ClF_3



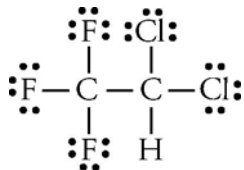
c. CH_2O



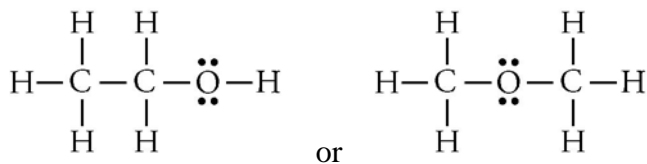
d. CN^-



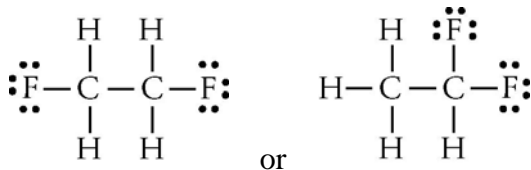
e. CF_3CHCl_2



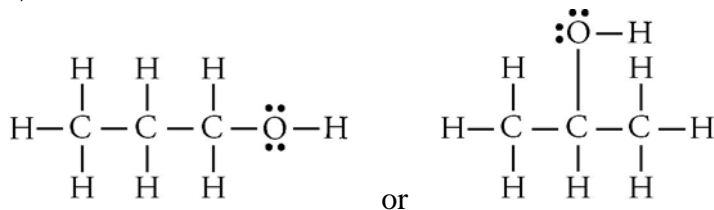
f. $\text{C}_2\text{H}_6\text{O}$



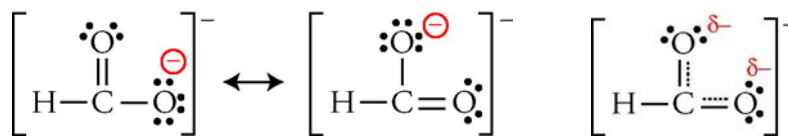
g. $\text{C}_2\text{H}_4\text{F}_2$



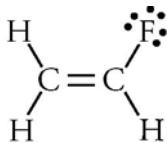
h. $\text{C}_3\text{H}_7\text{OH}$



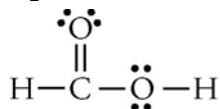
i. HCO_2^-



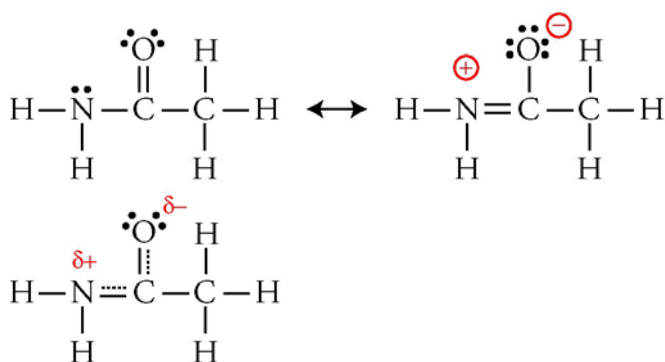
j. CH_2CHF



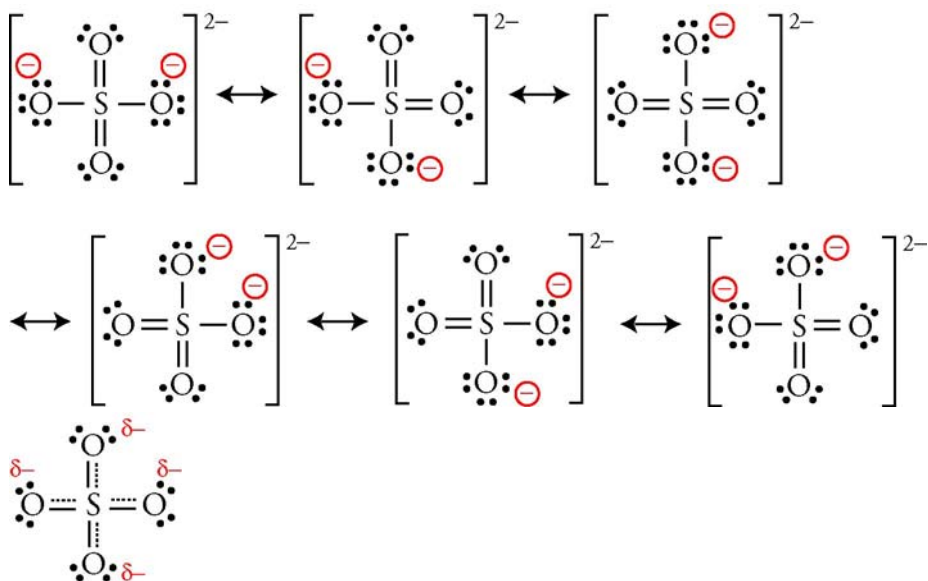
k. HCO_2H

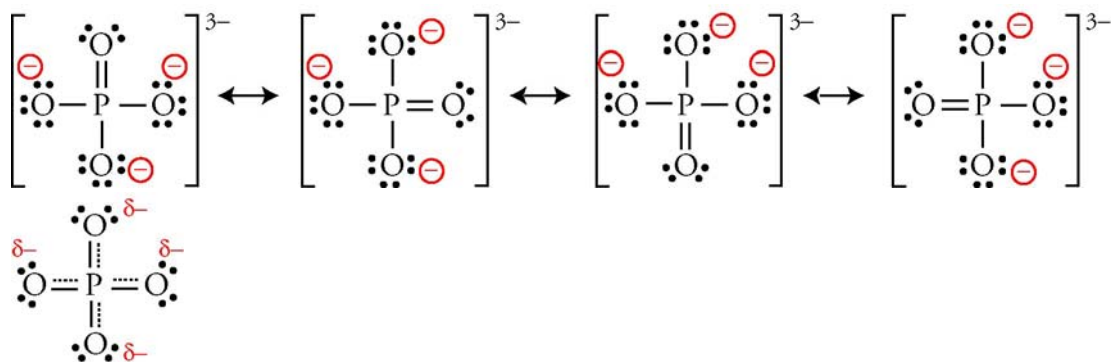


l. NH_2COCH_3

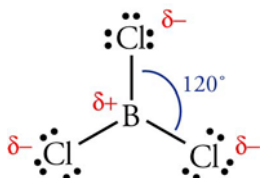


m. SO_4^{2-}



n. PO_4^{3-} 

Exercise 8,9.3 – Predicting Molecular Polarity: Decide whether the molecules represented by the following formulas are polar or nonpolar. (You may need to draw Lewis structures and geometric sketches in order to do this.)

a. BCl_3 

The 1.12 difference in electronegativity between B (2.04) and Cl (3.16) tells us that the bonds are polar covalent bonds, but there is a symmetrical distribution of these polar bonds. BCl_3 molecules are nonpolar.

b. C_3H_8

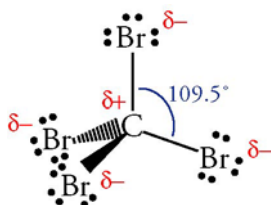
All hydrocarbons are composed of nonpolar molecules.

c. CS_2

The 0.03 difference in electronegativity between C (2.55) and S (2.58) tells us that the bonds are nonpolar covalent bonds. If there are no polar bonds, the molecules are nonpolar.

d. NH_3 

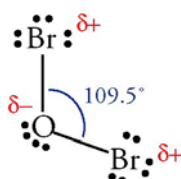
The 0.84 difference in electronegativity between N (3.04) and H (2.20) tells us that the bonds are polar covalent bonds. The asymmetrical distribution of these polar bonds leads to polar NH_3 molecules. All molecular compounds with the N-H bond are polar.

e. CBr_4 

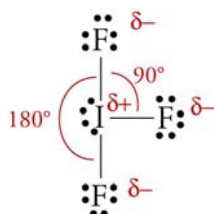
The 0.44 difference in electronegativity between C (2.55) and Br (2.96) tells us that the bonds are polar covalent bonds, but there is a symmetrical distribution of these polar bonds. CBr_4 molecules are nonpolar.

f. $\text{CH}_3\text{CO}_2\text{H}$

This is a carboxylic acid (acetic acid), and all carboxylic acids are composed of polar molecules.

g. Br_2O 

The 0.48 difference in electronegativity between O (3.44) and Br (2.96) tells us that the bonds are polar covalent bonds. The asymmetrical distribution of these polar bonds leads to polar Br_2O molecules.

h. IF_3 

The 1.32 difference in electronegativity between I (2.66) and F (3.98) tells us that the bonds are polar covalent bonds. The asymmetrical distribution of these polar bonds leads to polar IF_3 molecules.

Exercise 8,9.4 – Predicting Type of Attraction: What is the primary type of attraction that holds the particles of each of these substances in the liquid or solid form.

- a. Ir **metallic bonds**
- b. AlF_3 **ionic bonds**
- c. C_2H_6 **London forces**
- d. HCN **dipole-dipole attractions**
- e. C(dia) **covalent bonds**
- f. CH_3NH_2 **hydrogen bonds**
- g. NH_4OH **ionic bonds**
- h. BrF_3 **dipole-dipole attractions**
- i. CH_2O **dipole-dipole attractions**
- j. $\text{C}_3\text{H}_7\text{OH}$ **hydrogen bonds**

Exercise 8,9.5 - Predicting Types and Strengths of Attractions Between Particles:

For each of the following substances, write the name for the type of particle that forms its basic structure and the name of the primary type of attraction that can form between these particles. From each *pair* of substances, choose the one that you expect would have stronger interparticle attractions.

- a. ethanol, $\text{C}_2\text{H}_5\text{OH}$ (in alcoholic beverages)

Polar molecules and Hydrogen bonds (and London forces)

or dimethyl ether, CH_3OCH_3 (refrigerant and solvent for propellant sprays)

Polar molecules and dipole-dipole attractions (and London forces)

For molecules of about the same size, hydrogen bonds are stronger than dipole-dipole attractions, so we expect $\text{C}_2\text{H}_5\text{OH}$ to have the stronger attractions.

- b. manganese, Mn (improves corrosion resistance and hardness when alloyed with other metals)

Cations in a sea of electrons and metallic bonds

or carbon tetrachloride, CCl_4 (used for metal degreasing and to add chlorine atoms to organic compounds)

Nonpolar molecules and London forces

Metallic bonds are generally stronger than London forces, so we expect Mn to have the stronger attractions.

- c. propane, C_3H_8 (fuel and aerosol propellant)
 or n-heptane, C_7H_{16} (standard for octane rating – 100% n-hexane is zero octane)

Both nonpolar molecules and London forces

The larger the molecules are, the stronger the London forces between them, so C_7H_{16} has stronger attractions.

- d. acetone, CH_3COCH_3 (solvent for paints and varnishes)

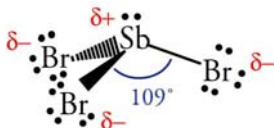
Polar molecules and dipole-dipole attractions (and London forces)

or lithium nitrate, $LiNO_3$ (used to make ceramics and in rocket propellants)

Cations and anions held together by ionic bonds

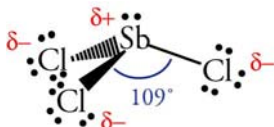
Ionic bonds are generally stronger than dipole-dipole attractions, so $LiNO_3$ has the stronger attractions.

- e. antimony tribromide, $SbBr_3$ (a molecular compound used as a mordant, which fixes dyes to textiles)



**Polar Molecules and dipole-dipole attractions
(and London forces)**

or antimony trichloride, $SbCl_3$ (a molecular compound used for fireproofing textiles)



**Polar Molecules and dipole-dipole attractions
(and London forces)**

London forces play a part in the attractions between polar molecules as well as nonpolar molecules. The larger the molecules are, the stronger the London forces between them are. Therefore, $SbBr_3$ has stronger attractions.