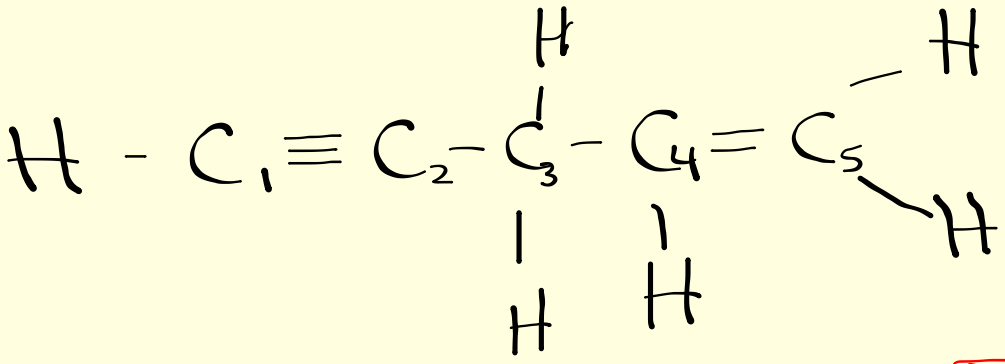
p. 236 #23-29



σ/π

$C_1 \rightarrow S + p_x \rightarrow sp + sp$ linear (2/2)

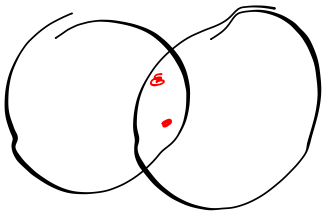
$C_2 \rightarrow S + p_x \rightarrow sp + sp$ linear (2/2)

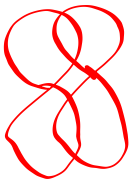
$C_3 \rightarrow S + p_x + p_y + p_z \rightarrow sp^3$ tetrahedral (4/0)

$C_4 \rightarrow S + p_x + p_y \rightarrow sp^2$ trig. planar (3/1)

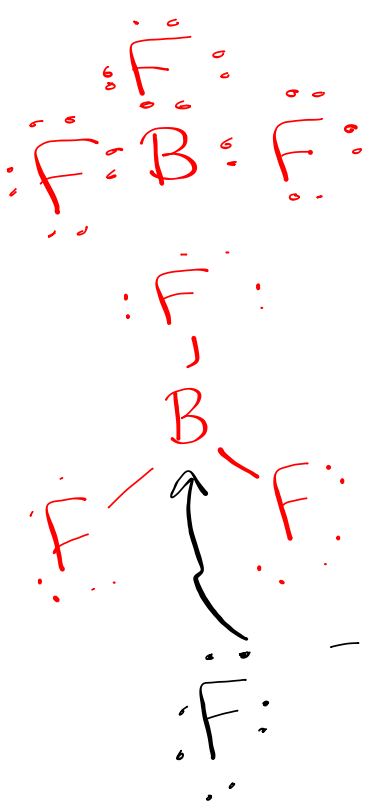
sp^2 sp^2

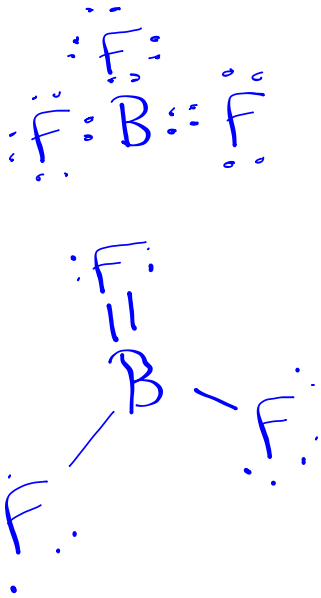
$C_5 \rightarrow S + p_x + p_y \rightarrow sp^2$ trig. planar (3/1)



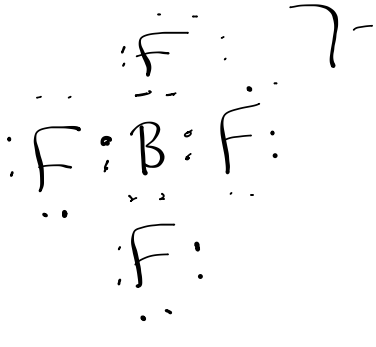


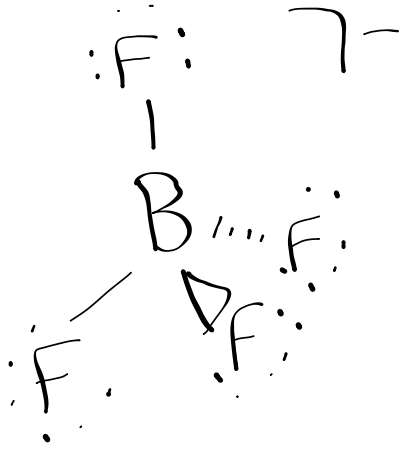
BF_3





tetrahedral





Single bonds

$4sp^3$ ($s + p_x + p_y + p_z$)

-sigma bonds

tetrahedral, pyramidal, bent

Double bonds

$3sp^2$ ($s + p_x + p_y$)

- 3 sigma bonds

p_z

- pi bond

trigonal planar

Triple bonds

$2sp$ ($s + p_x$)

- 2 sigma bonds

p_y, p_z

- 2 pi bonds

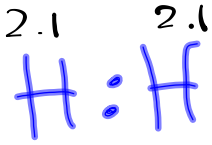
linear

Bond Polarity

- In covalent bonds, the bonding pairs of electrons are shared between atoms.
- Two nuclei 'pull' the electrons. Amount of 'pull' is dependent on the atoms' electronegativities.

Nonpolar covalent bond
Bond that forms when the atoms in the bond pull equally, and the electrons are shared equally.

Ex.

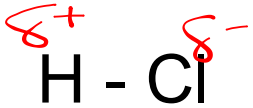


H - H

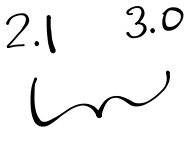
Polar covalent bond
Bond that forms when the electrons are shared unequally

- More electronegative atom attracts electrons more strongly and gains a slightly negative charge. Less electronegative atom has a slightly positive charge.

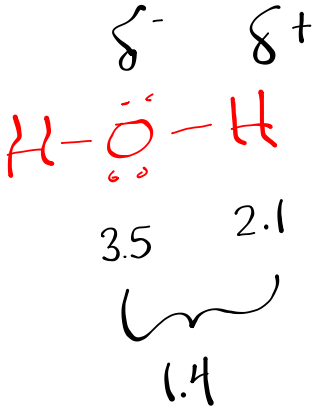
Ex.



H - Cl



H_2O



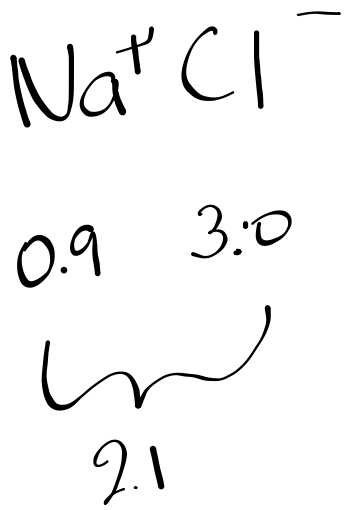
$$\overset{2.5}{C} : \overset{2.5}{S}$$

$$\overset{2.5}{C} : \overset{2.8}{Br}$$

Table 6.3 Electronegativity Differences and Bond Types

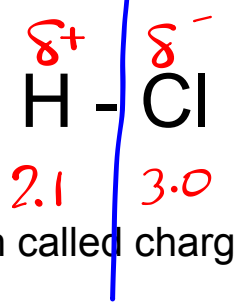
Electronegativty difference range	Most probable type of bond	Example
0.0-0.4	Nonpolar covalent	H - H (0.0)
0.4-1.0	Moderately polar covalent	H - Cl (0.9)
1.0-2.0	Very polar covalent	H - F (1.9)
≥ 2.0	Ionic	Na^+ Cl^- (2.1)

* No sharp boundary between ionic and covalent


$Na^+ Cl^-$
0.9
3.0
2.1

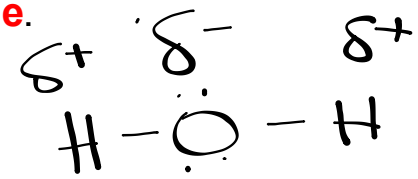
Polar Molecules

In a polar molecule, one end of the molecule is slightly negative, and the other end is slightly positive.



-Partial charges are often called charged regions or poles.

A molecule with two poles is called a **dipole**.



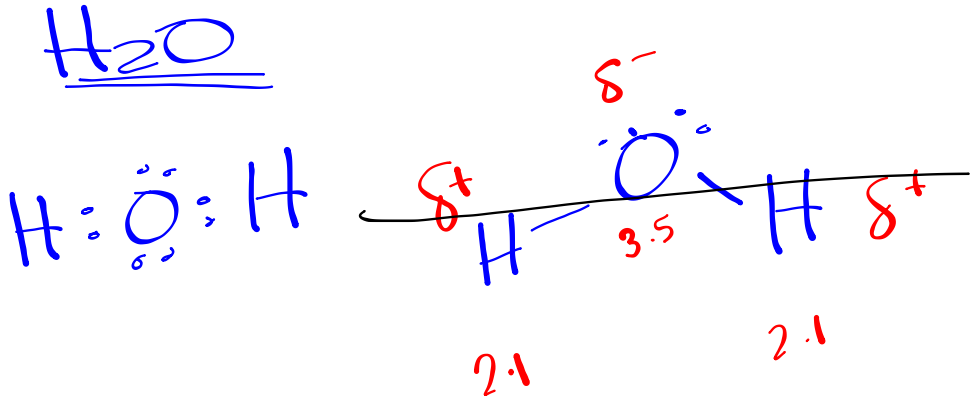


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Table 6.2

Electronegativity Values for Selected Elements

H 2.1						
Li 1.0	**Be** 1.5	**B** 2.0	**C** 2.5	**N** 3.0	**O** 3.5	**F** 4.0
Na 0.9	**Mg** 1.2	**Al** 1.5	**Si** 1.8	**P** 2.1	**S** 2.5	**Cl** 3.0
K 0.8	**Ca** 1.0	**Ga** 1.6	**Ge** 1.8	**As** 2.0	**Se** 2.4	**Br** 2.8
Rb 0.8	**Sr** 1.0	**In** 1.7	**Sn** 1.8	**Sb** 1.9	**Te** 2.1	**I** 2.5
Cs 0.7	**Ba** 0.9	**Tl** 1.8	**Pb** 1.9	**Bi** 1.9		