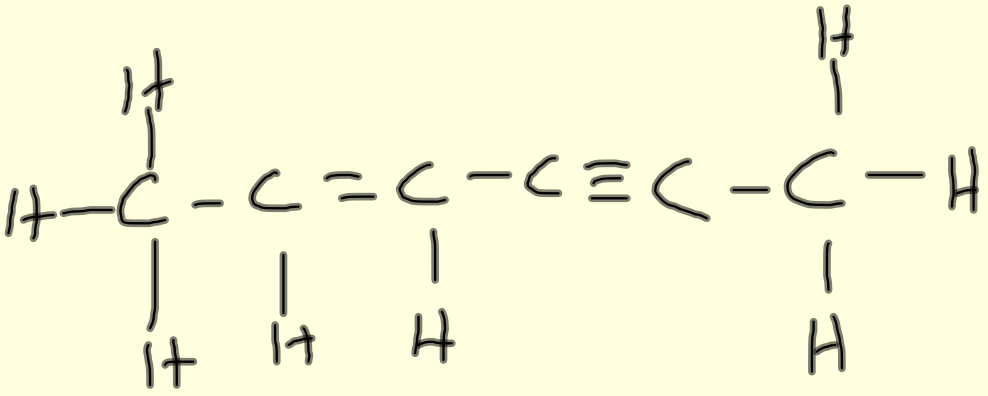
Part D



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

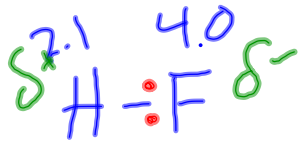
p_y, p_z

- 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.



Think-think about it:

a) What do you think happens in a diatomic molecule if the electronegativity is different?

b) What if the electronegativity is the same?

Nonpolar covalent bond

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

Ex. H - H :N ≡ N :

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_2O H - Cl

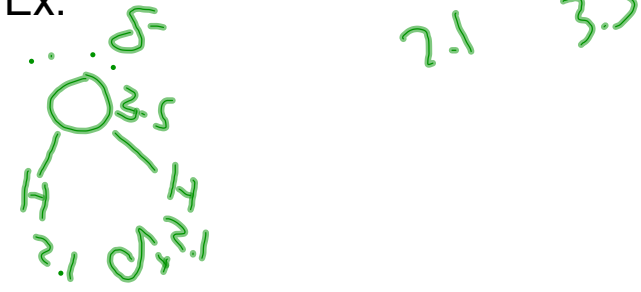


Table 6.3 Electronegativity Differences and Bond Types

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

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		