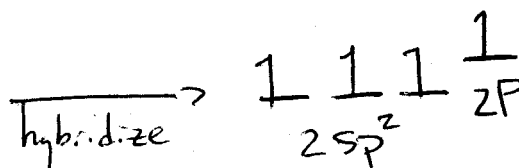
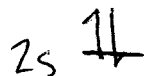
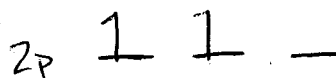
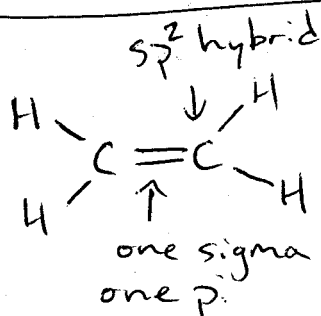
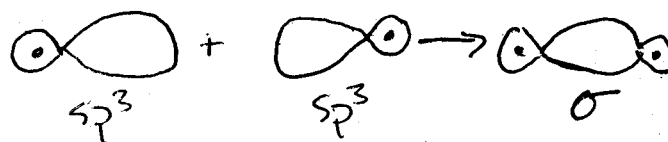
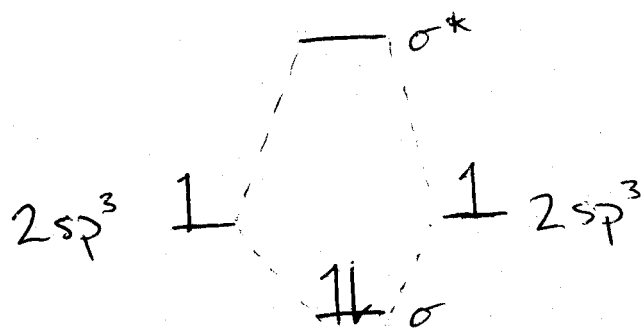
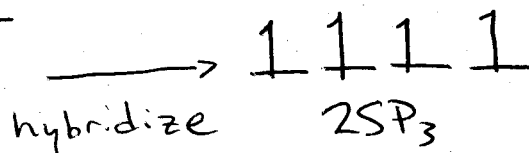
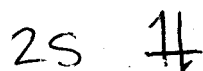
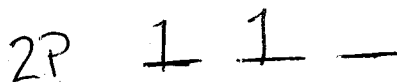
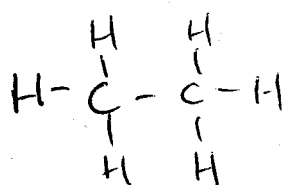
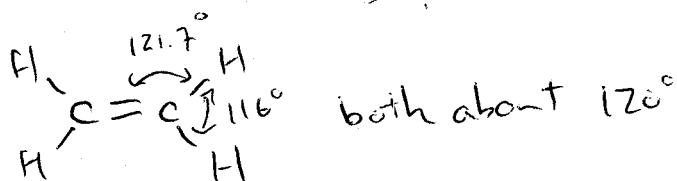


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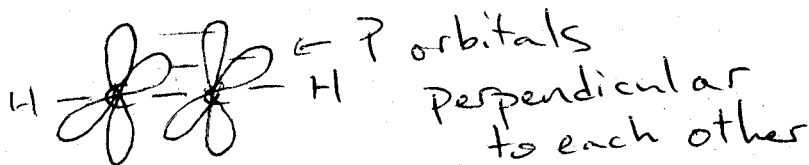
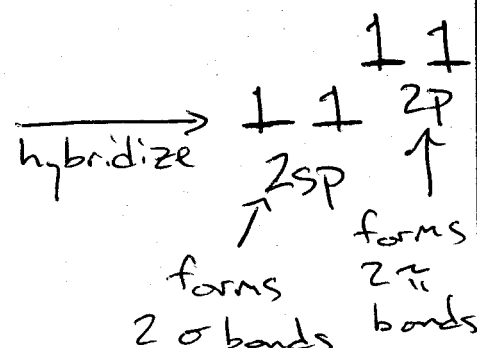
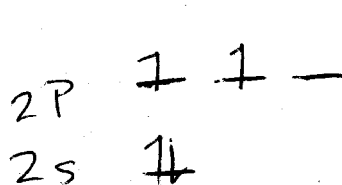
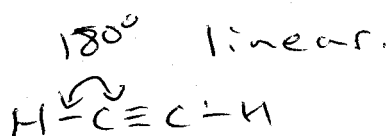
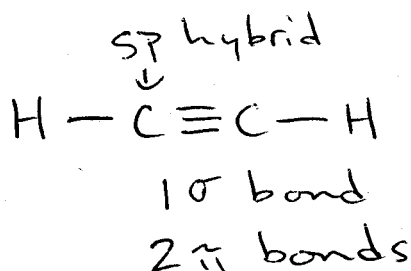
↑
forms 3 σ bonds
from sp^2 orbitals
one π bond from
P orbital

C=C B.O. E. diagram



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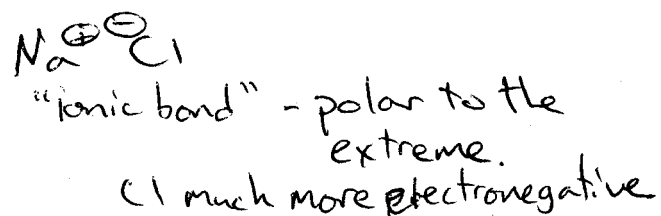
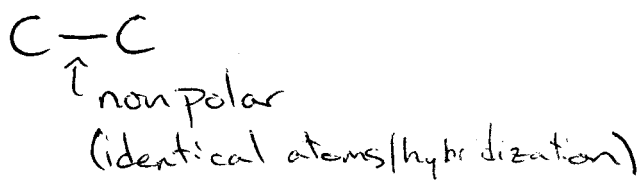


triple bond stronger than double bond & single
 Bond strength $3x > 2x > 1$

Electronegativity: increases as you go up & to the right on the periodic table.

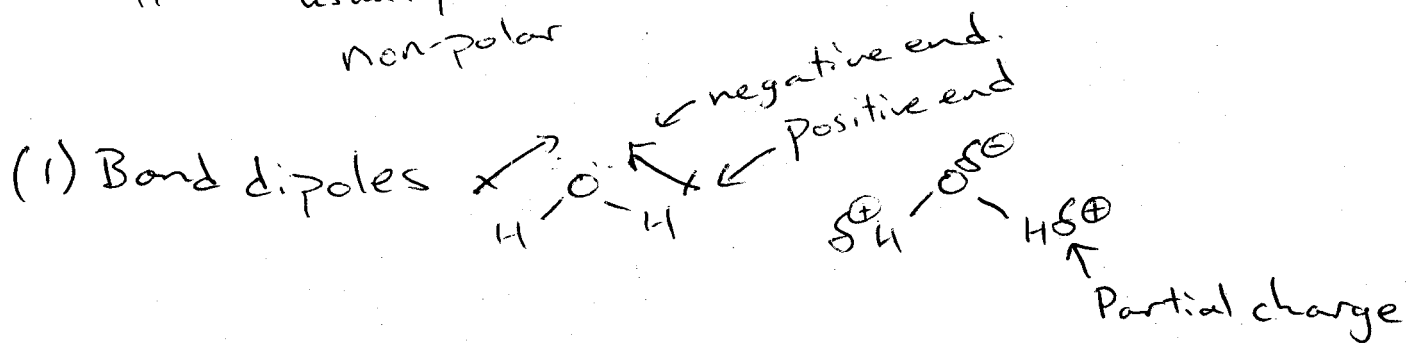
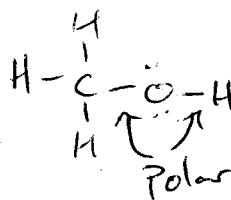
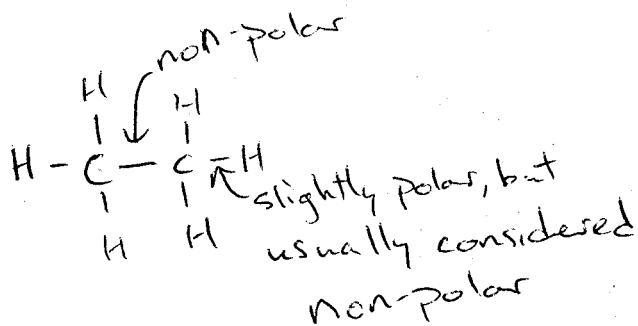
Polarity in bonds & molecules

the way in which e^- are shared between atoms is determined by the electronegativities of the partners

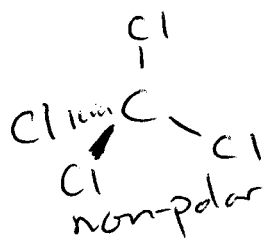
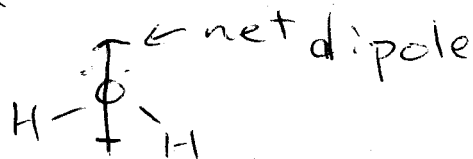


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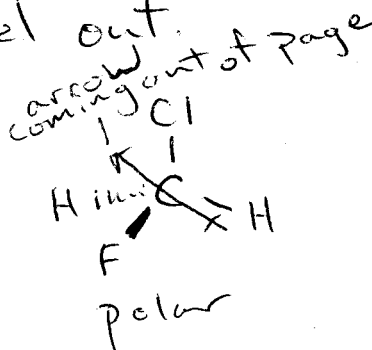
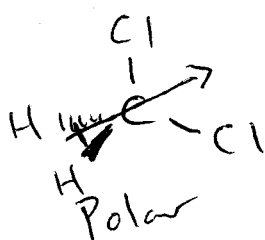
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(2) Molecular dipole moment
(sum of the bond dipoles)



each bond is polar, but no net dipole moment - bond dipoles cancel out.



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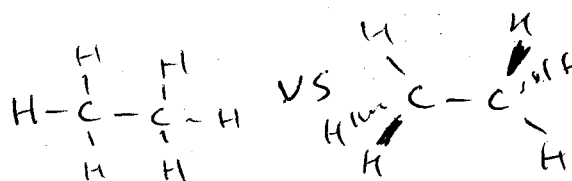
Ethane

 C_2H_6

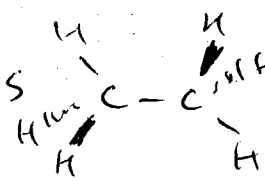
vs.

 CH_3CH_3

vs.



vs.



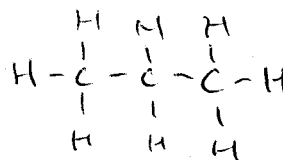
Propane

 C_3H_8

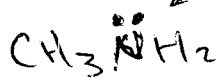
vs.

 $CH_3CH_2CH_3$

vs.

Line drawings!~~Ethane~~ Propane:↑
line ends
are carbon atoms.Hydrogens not drawn. Just
assumed to be there.Heteroatoms (anything other than C & H)
always drawn.Methylamine: CH_3NH_2

vs.

← when starting, it's a
good idea to draw
lone pairs

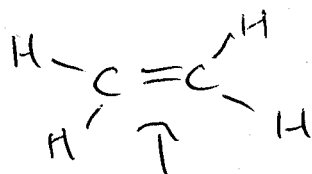
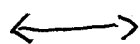
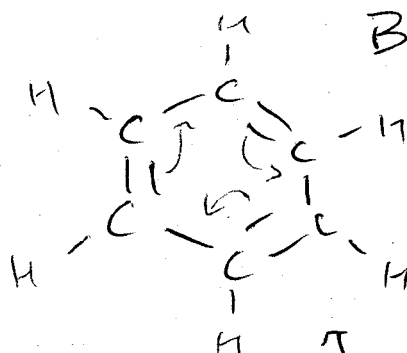
↑

lone pair on N assumed.

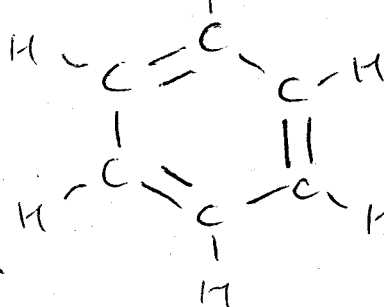
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Localized Bonding

 π bond
doesn't
move.double headed
arrow indicates
~~resonance~~ resonance
structures.

Benzene:

all bond
lengths are
the sameresonance
structuresDelocalized
Bonding