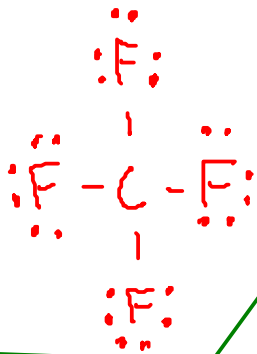


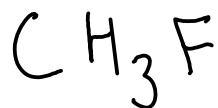
## Examples:



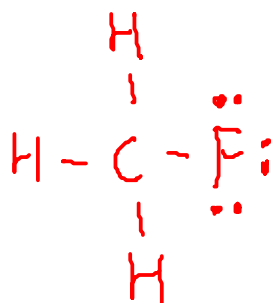
$$\begin{array}{l} C: 1 \times 4 \\ F: 4 \times 7 \\ \hline 32 e^- \end{array}$$



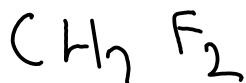
Polar? 1 - C-F bonds are polar. 2 - Geometry? This molecule is TETRAHEDRAL. Since the fluorine atoms are arranged SYMMETRICALLY around the center, they all pull against one another and the overall molecule is NONPOLAR.



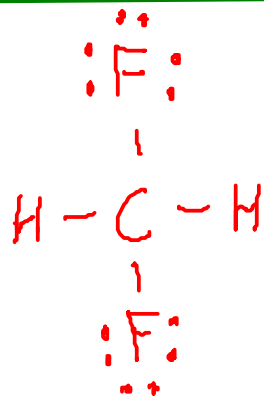
$$\begin{array}{l} C: 1 \times 4 \\ H: 3 \times 1 \\ F: 1 \times 7 \\ \hline 14 e^- \end{array}$$



Polar? 1 - The C-F bond is polar. 2 - Geometry? Tetrahedral again, like carbon tetrafluoride. But since there's only one fluorine in the structure, electrons are pulled towards the fluorine atom and away from the other side of the molecule (with the hydrogens). This is a POLAR molecule.



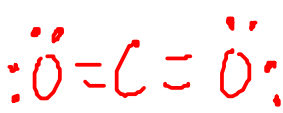
$$\begin{array}{l} C: 1 \times 4 \\ H: 2 \times 1 \\ F: 2 \times 7 \\ \hline 20 e^- \end{array}$$



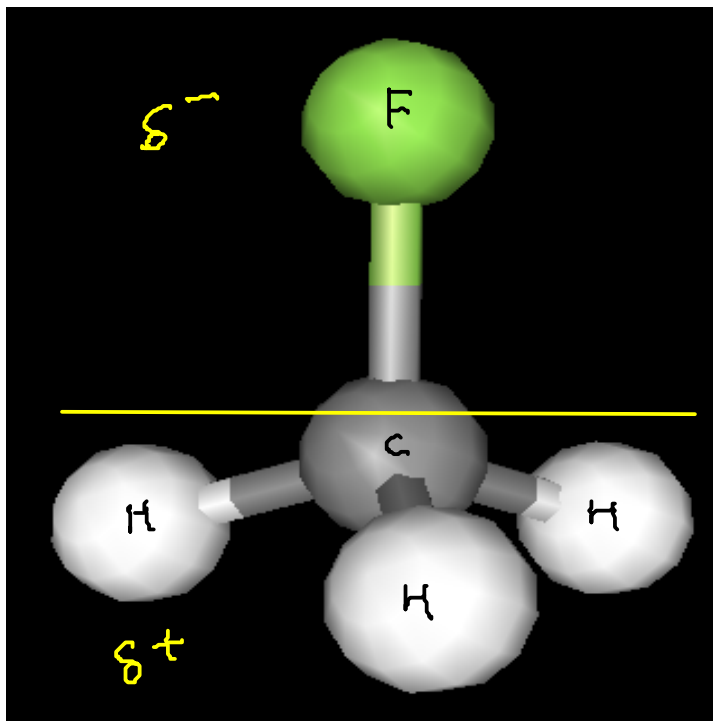
Polar? 1 - C-F bonds are polar. 2 - Geometry? Tetrahedral. But in the THREE-DIMENSIONAL tetrahedron, the fluorine atoms are on one side of the molecule and the hydrogen atoms are on the other. Fluorines pull electrons towards themselves, making a POLAR molecule.



$$\begin{array}{l} C: 1 \times 4 \\ O: 2 \times 6 \\ \hline 16 e^- \end{array} \quad \cdot\cdot \quad \cdot\cdot$$



Polar? 1 - C=O bonds are polar. 2 - Geometry? Linear. Both C=O bonds directly oppose each other, making the molecule as a whole NONPOLAR.

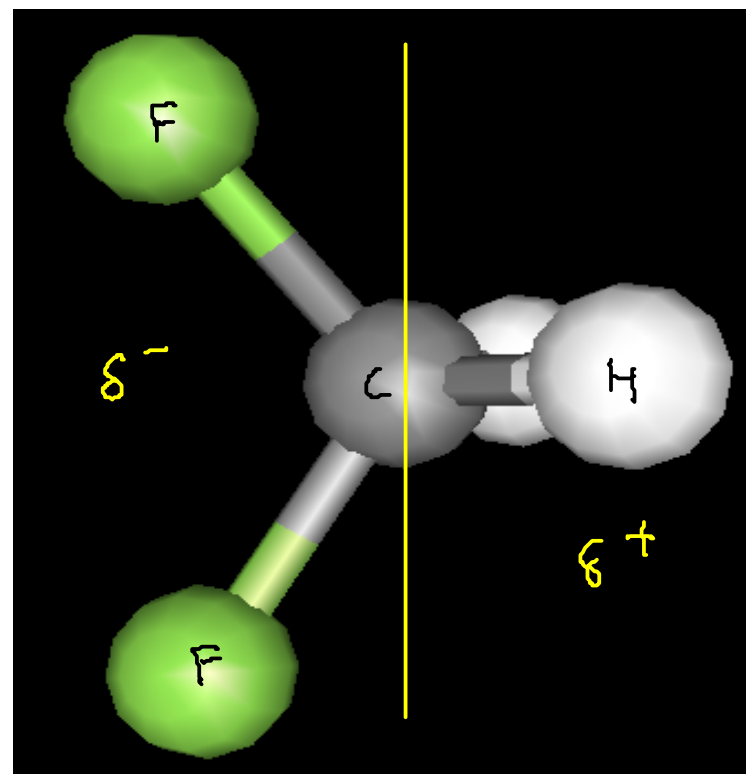
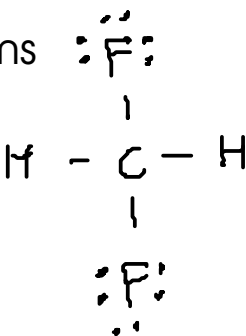


←  $\text{CH}_3\text{F}$  "fluoromethane"

Fluorine is able to pull electron density through the molecule, as it is being opposed by much less electronegative hydrogen atoms.

"difluoromethane"  $\text{CH}_2\text{F}_2$  →

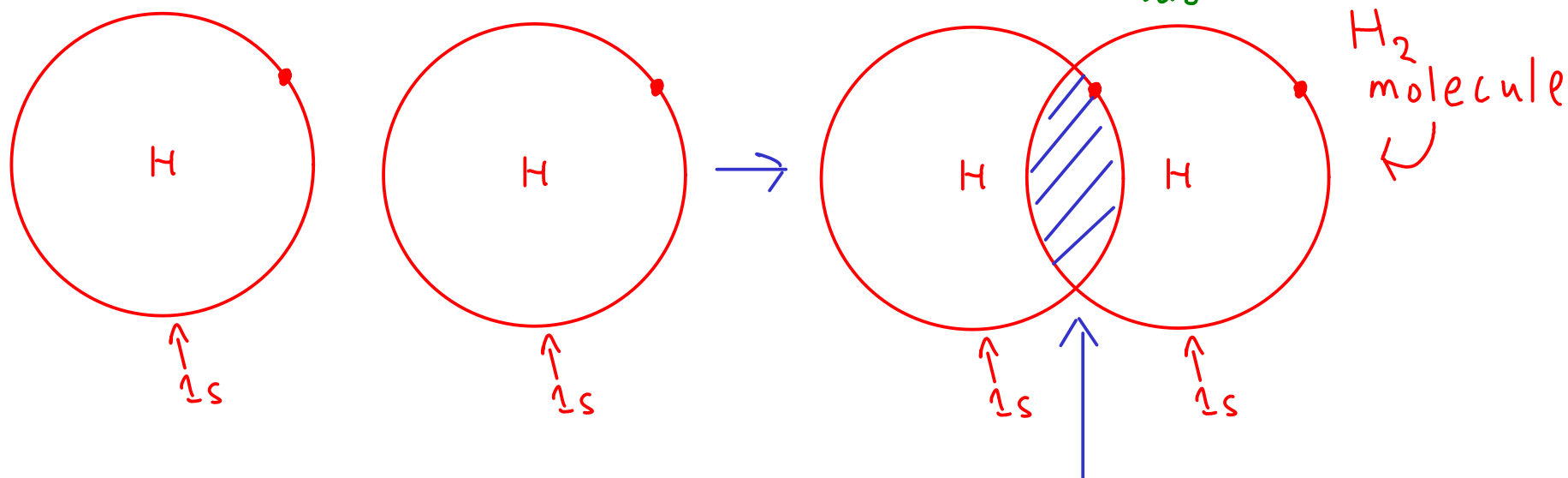
In 2D, the fluorine atoms appear to be on the opposite sides of the molecule, but in 3D they are on the same side.



## VALENCE BOND THEORY

- an attempt to explain why molecules behave in the way that the VSEPR model predicts.
- Describes the formation of bonds in terms of the OVERLAP of ORBITALS from the bonding atoms.

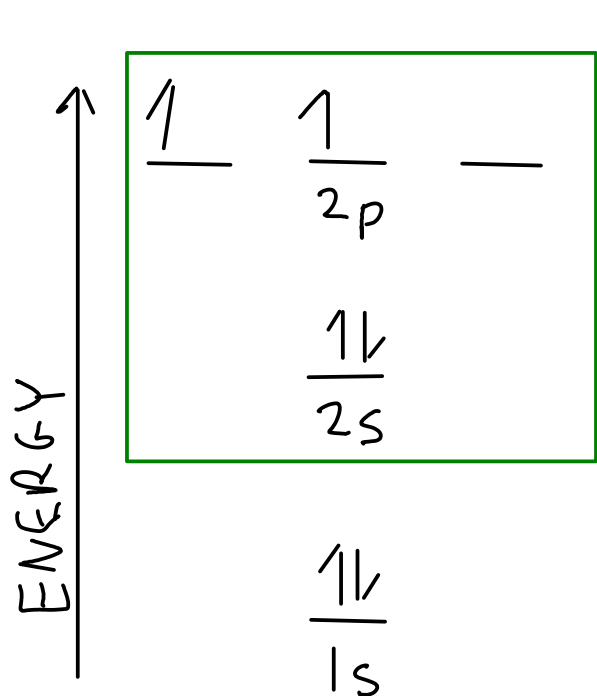
- 1 Bonds are formed when two atoms are close enough together so that their ORBITALS OVERLAP (share the same space).
- 2 Each SET of overlapping orbitals can contain at most a total of TWO electrons. So, two orbitals with one electron each may bond. An orbital with two electrons can only bond with an EMPTY orbital (This is called a COORDINATE COVALENT BOND.) *\*Ag<sup>+</sup> with :NH<sub>3</sub>... the cleanup in the AgCl lab*



These 1s orbitals overlap to form what we call a "sigma bond" with overlap BETWEEN the two atomic nuclei.

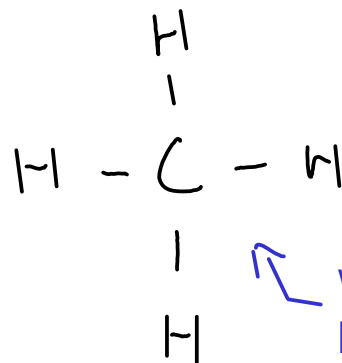
## Hybridization

- Look at carbon's electron configuration:



valence

You would expect that carbon would form several different kinds of bonds in a molecule like methane. But, methane's bonds are experimentally all identical. How does carbon form the four equivalent C-H bonds we see in methane?

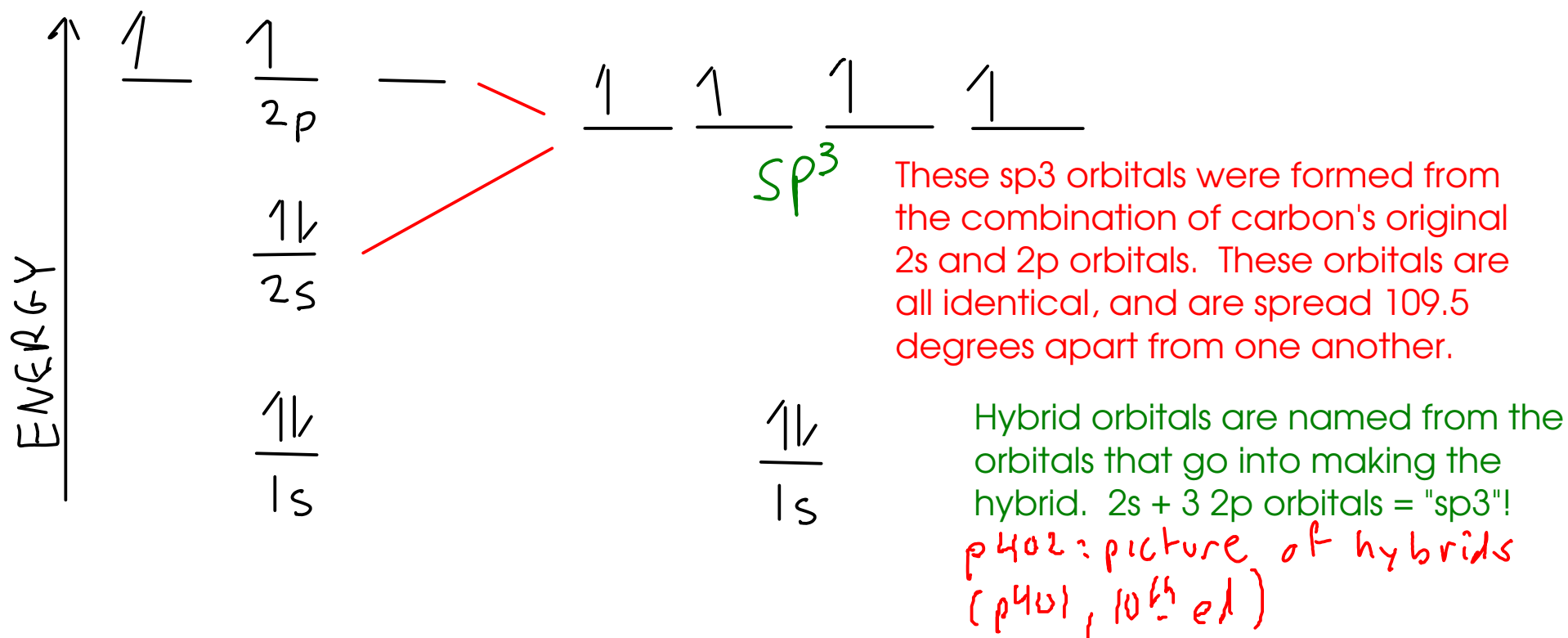


↙ We observe that these bonds are IDENTICAL! Same bond energy, distance, and angle.

- In valence bond theory, atomic orbitals can COMBINE to make new orbitals that can then go on to bond with other molecules.

- When orbitals combine to make HYBRID ORBITALS, ...

- ① The overall NUMBER OF ORBITALS does not change.
- ② The overall NUMBER OF ELECTRONS around the atom does not change
- ③ The energy of the orbitals is between the energies of the orbitals that combine.



Types of hybrid orbitals:

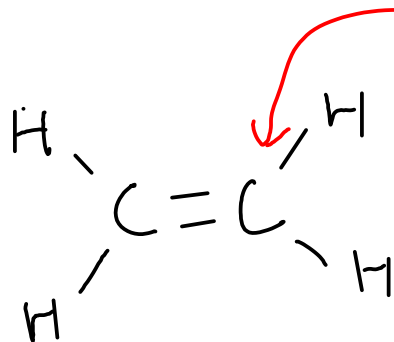
Hybrid type	Number of orbitals	Molecular shape
sp	2	linear
sp <sup>2</sup>	3	trigonal planar
sp <sup>3</sup>	4	tetrahedral (or derivatives)
sp <sup>3</sup> d	5	trigonal bipyramidal (or derivatives)
sp <sup>3</sup> d <sup>2</sup>	6	octahedral (or derivatives)

p402 : picture of hybrids  
(p401, 104b)

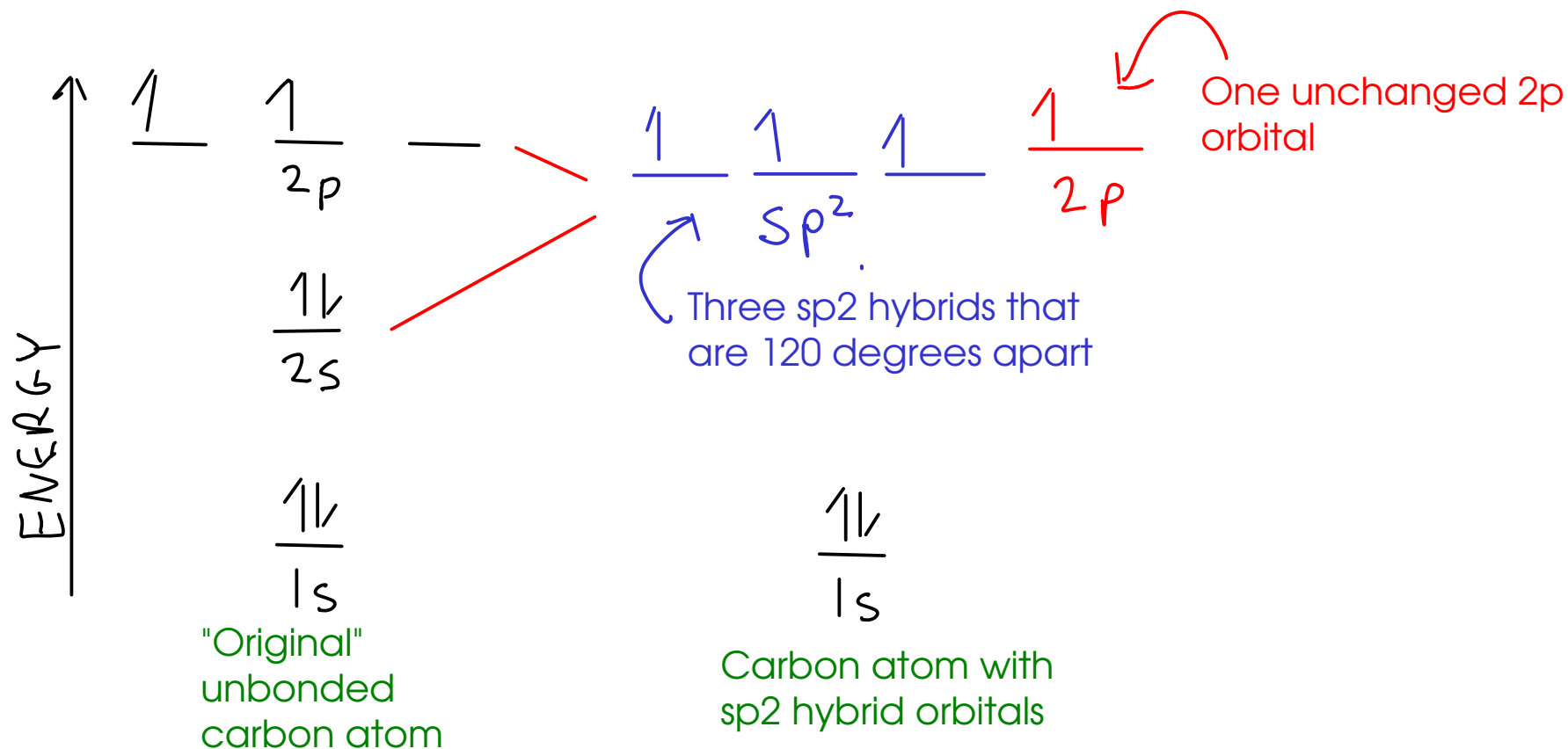
## MULTIPLE BONDS and VALENCE BOND THEORY

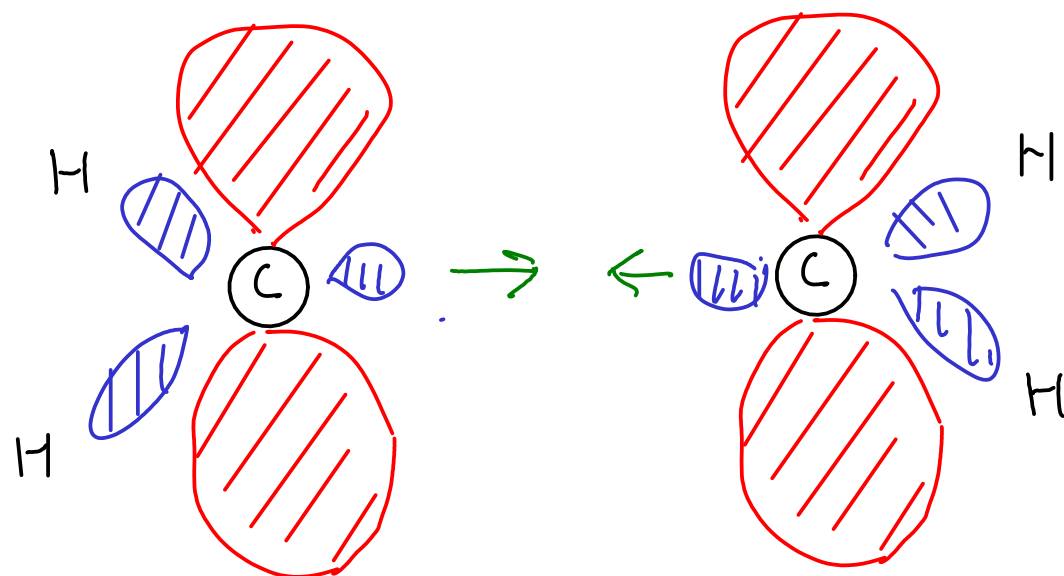
- Valence bond theory provides an explanation of multiple (double and triple) bonding that explains some interesting observations about these kinds of bonds.

$C_2H_4$  :  
ethylene



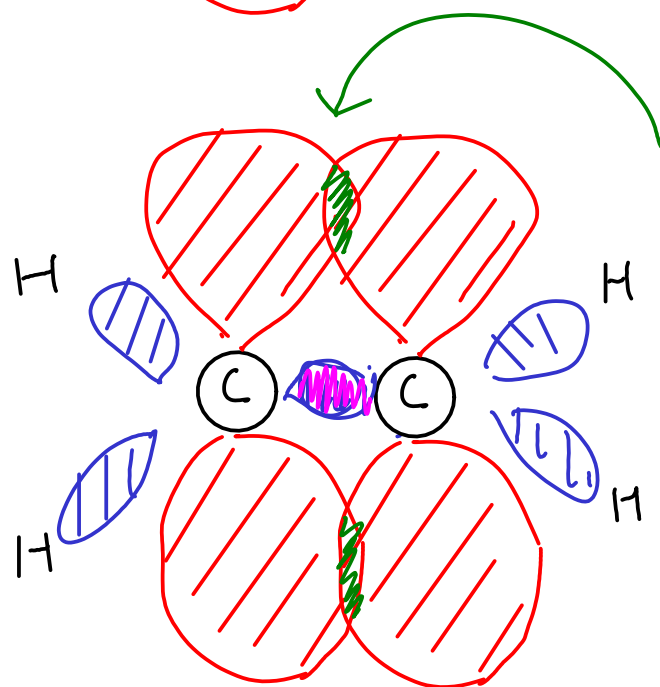
Each carbon has a TRIGONAL PLANAR geometry. This suggests that the carbons are "sp<sup>2</sup> hybridized".





$sp^2$  hybrid orbitals in BLUE

$2p$  orbital in RED



The  $2p$  orbitals overlap above and below the axis between the two carbon atoms. This OFF-AXIS overlap is called a PI BOND.

The  $sp^2$  hybrid orbitals overlap ON THE AXIS between the two carbon atoms. This bond is called a SIGMA BOND.

As you can see, the carbon-carbon double bond in ethylene is made up of TWO DIFFERENT KINDS OF BONDS!



## Some notes on sigma and pi bonds:

① SIGMA bonds are formed when orbitals overlap along the axis between two atoms. These bonds have good overlap between the bonding orbitals, meaning that they are strong. Single bonds are always sigma bonds. Double and triple bonds contain one sigma bond each.

② PI bonds are formed when off-axis orbitals (usually p orbitals) overlap. Since the overlapping orbitals do not face each other as in the sigma bond, the overlap in pi bonds tends to be poorer than in sigma bonds. As a result, pi bonds tend to be weaker than sigma bonds. Double bonds contain a single pi bond, and triple bonds contain two pi bonds.\*

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\*Experimentally, we observe that the bond energy of the C=C bond is less than the bond energy of two C-C bonds. This suggests that the second bond in a double bond is different from the first!

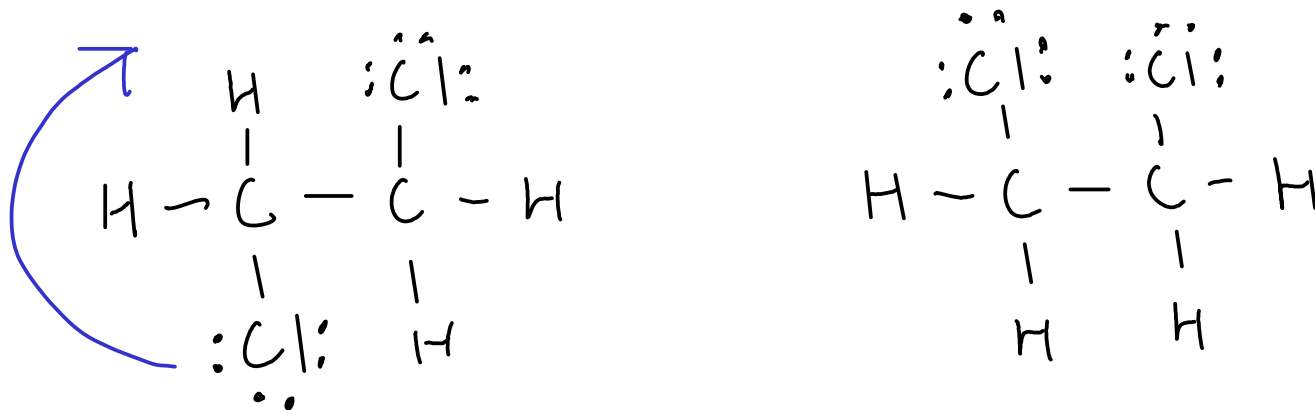
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③ Molecules may rotate around SIGMA bonds, since rotation around the axis between two atoms will not affect the overlap and break the bond. Off-axis PI BONDS prevent rotation because rotation would break the pi bond.

# ROTATION, ISOMERS, and VALENCE BOND THEORY

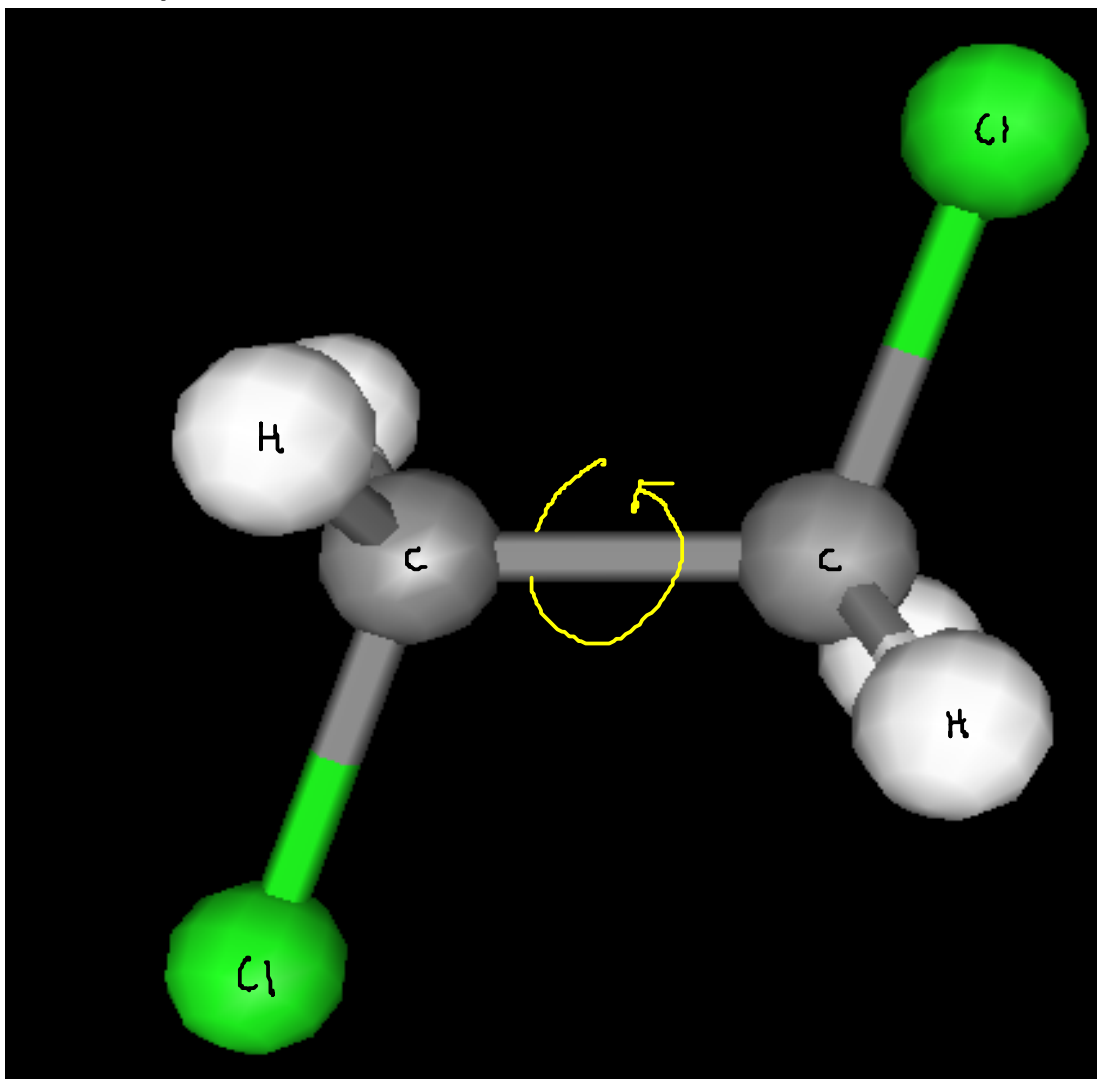
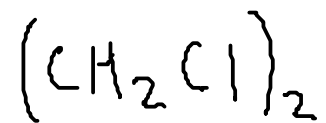
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- Consider this molecule:  $(\text{CH}_2\text{Cl})_2$   
 "1,2-dichloroethane"



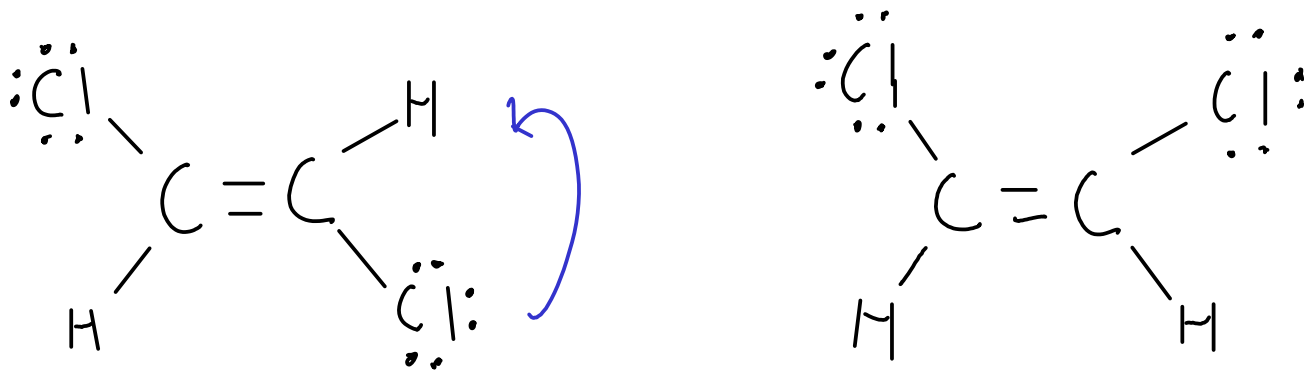
... are these two structures different?

No! The molecule is free to rotate around the C-C single (sigma) bond, and we do not observe two different versions of 1,2-dichloroethane. Both of the forms drawn above are equivalent.



The molecule is free to rotate about the carbon-carbon bond!

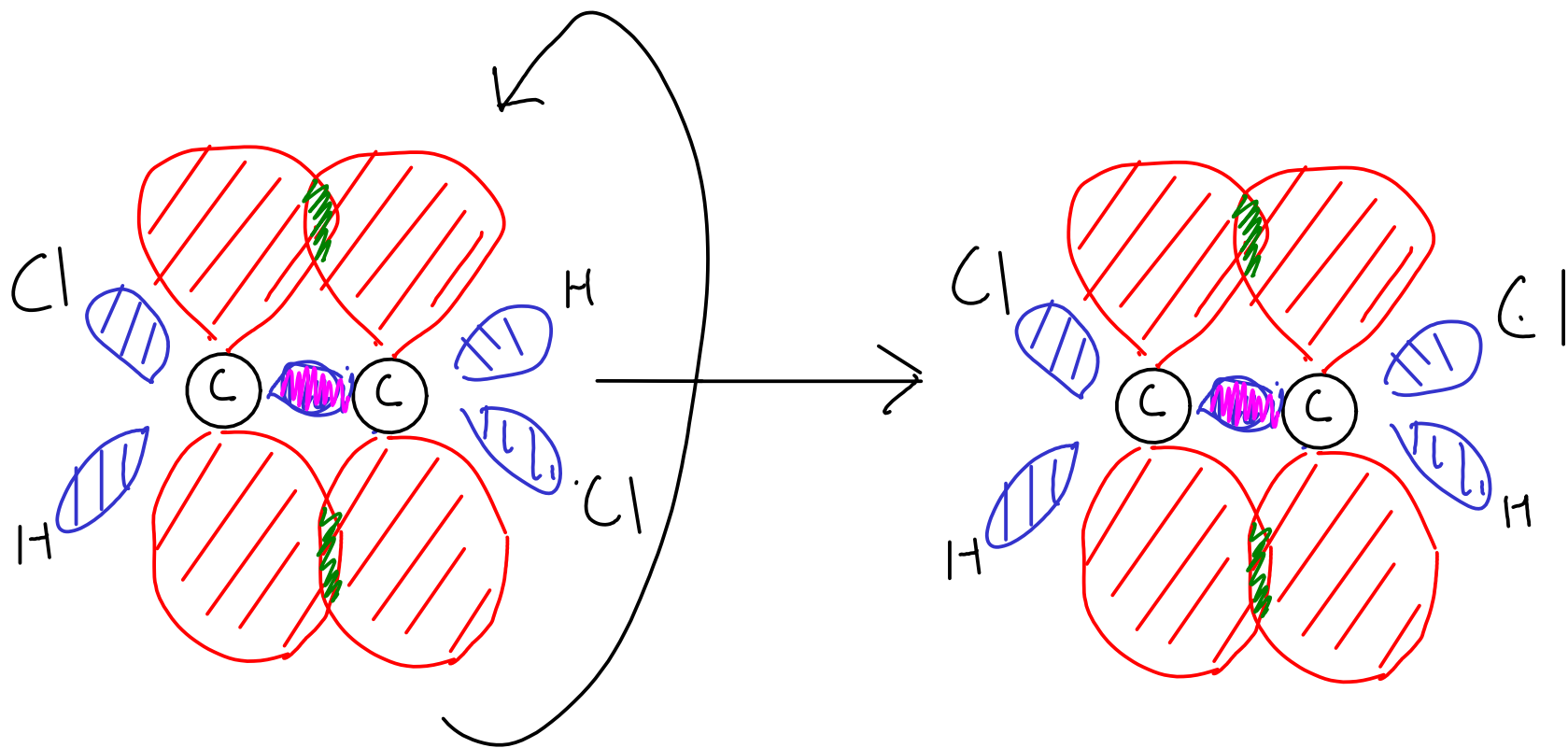
... now consider "1,2-dichloroethene":  $(\text{CHCl})_2$



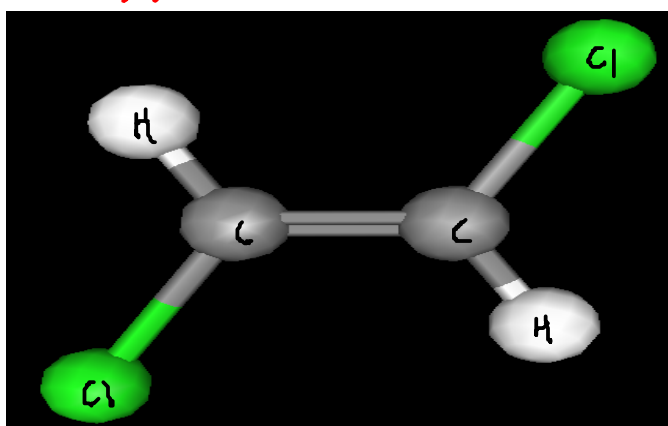
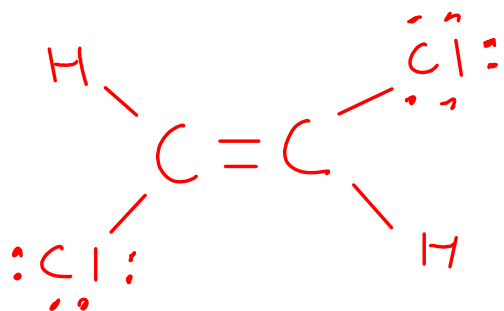
... are these two structures different?

YES! The two carbon atoms in these structures are held together by a DOUBLE BOND, which contains a pi bond. The molecule cannot rotate around the C=C double bond without breaking the pi bond, so the form with the two chlorine atoms on opposite sides cannot freely flip over to the form with the chlorine atoms on the same side.

These two Lewis structures actually represent DIFFERENT MOLECULES. They are called ISOMERS, since they have the same chemical formula but different arrangements of atoms.



For this rotation to take place, the PI BOND must break and then re-form!

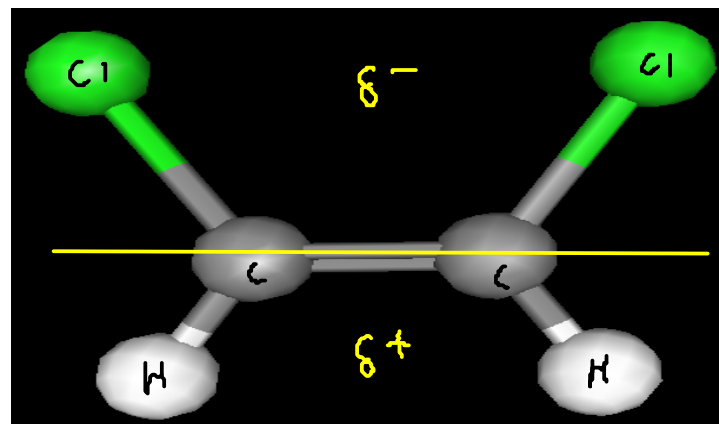
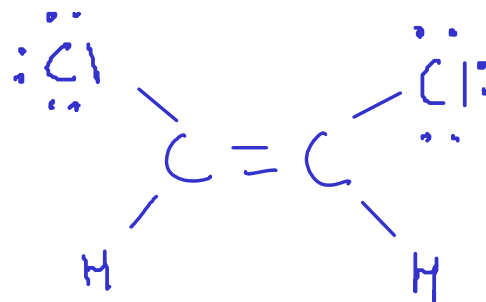


trans 1,2-dichloroethene

BOILING POINT: 47.5 C

POLARITY: NONPOLAR (0 D dipole moment)

DENSITY: 1.26 g/mL



cis 1,2-dichloroethene

BOILING POINT: 60.3 C

POLARITY: POLAR (1.9 D dipole moment)

DENSITY: 1.28 g/mL

\* As you can see, some of the properties of these two molecules are very different! The presence of the pi bond (part of the double bond) in each of these molecules means conversion from one form to the other requires a chemical reaction instead of a simple rotation.

\* Double (and triple) bonds prevent rotation and "fix" the structure of a molecule. This is easily explained by valence bond theory!