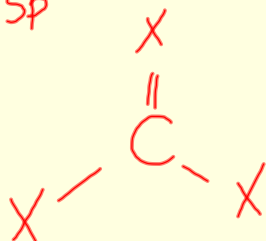
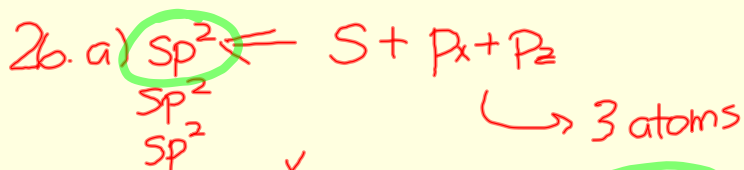
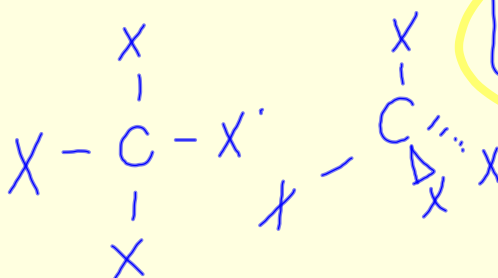
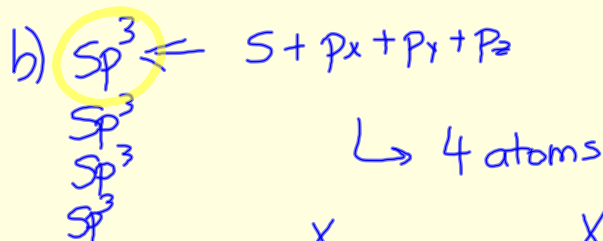


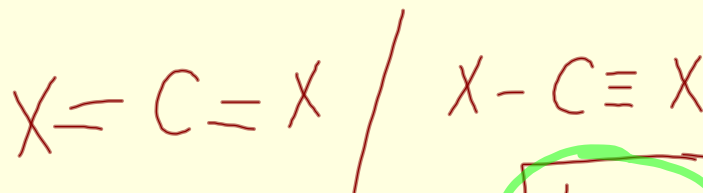
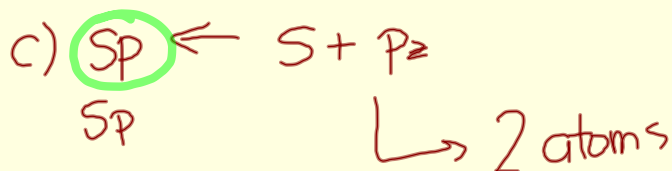
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TRIGONAL  
PLANAR



TETRAHEDRAL

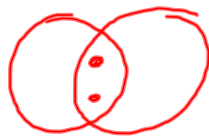


LINEAR

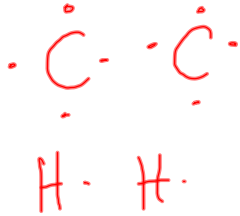


27

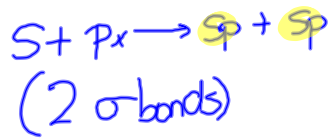
SIGMA BOND → head-on overlap



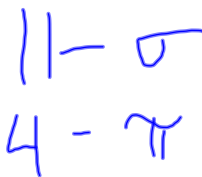
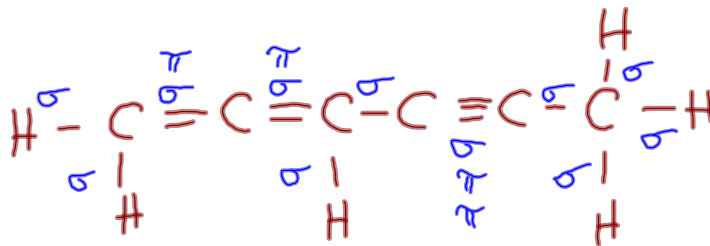
28 C<sub>2</sub>H<sub>2</sub>



C<sub>1</sub>



C<sub>2</sub>





### **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

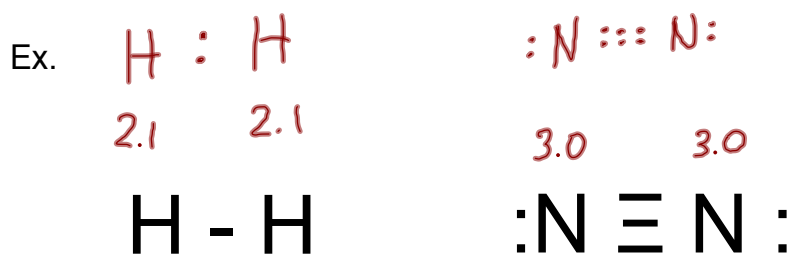
linear

# 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

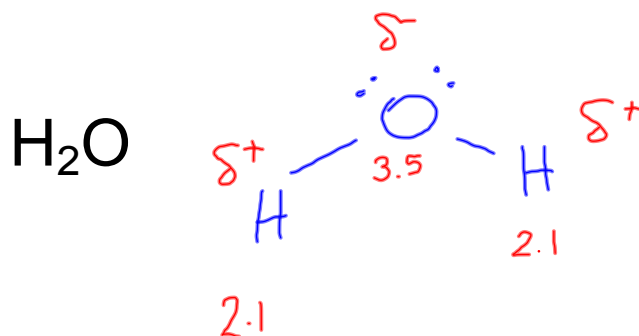
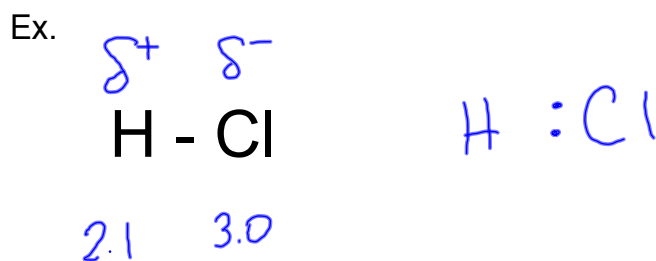
Bond that forms when the atoms in the bond pull equally, and the electrons are shared equally.



## 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.



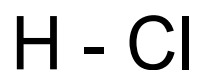
**Table 6.3 Electronegativity Differences and Bond Types**

<b>Electronegativity difference range</b>	<b>Most probable type of bond</b>	<b>Example</b>
<b>0.0-0.4</b>	<b>Nonpolar covalent</b>	<b>H - H (0.0)</b>
<b>0.4-1.0</b>	<b>Moderately polar covalent</b>	<b>H - Cl (0.9)</b>
<b>1.0-2.0</b>	<b>Very polar covalent</b>	<b>H - F (1.9)</b>
<b>≥ 2.0</b>	<b>Ionic</b>	<b>Na<sup>+</sup> Cl<sup>-</sup> (2.1)</b>

**\* No sharp boundary between ionic and covalent**

# 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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**Practice Problems #30, 31**