Chapter 8

Covalent bonding

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Covalent Bonding

- A metal and a nonmetal transfer electrons
  – An ionic bond
- Two metals just mix and don't react
  – An alloy
- What do two nonmetals do?
  – Neither one will give away an electron
  – So they share their valence electrons
  – This is a covalent bond

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Covalent bonding

- Makes molecules
  – Specific atoms joined by sharing electrons
- Two kinds of molecules:
- Molecular compound
  – Sharing by different elements
- Diatomic molecules
  – Two of the same atom
  – O_2 N_2

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Diatomic elements

- There are 8 elements that always form molecules
- H_2 , N_2 , O_2 , F_2 , Cl_2 , Br_2 , I_2 , and At_2
- Oxygen by itself means O_2
- The –ogens and the –ines
- 1 + 7 pattern on the periodic table

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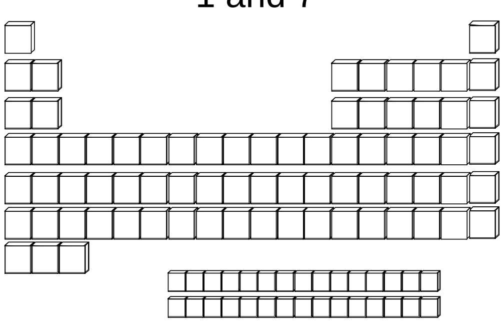
1 and 7



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Molecular compounds

- Tend to have low melting and boiling points
- Have a molecular formula which shows type and number of atoms in a molecule
- Not necessarily the lowest ratio
- $C_6H_{12}O_6$
- Formula doesn't tell you about how atoms are arranged

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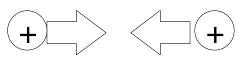
How does H_2 form?

- The nuclei repel



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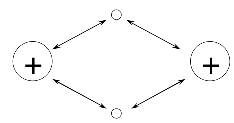
How does H_2 form?

- The nuclei repel
- But they are attracted to electrons
- They share the electrons



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Covalent bonds

- Nonmetals hold onto their valence electrons.
- They can't give away electrons to bond.
- Still need noble gas configuration.
- Get it by sharing valence electrons with each other.
- By sharing both atoms get to count the electrons toward noble gas configuration.

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Covalent bonding

- Fluorine has seven valence electrons
- A second atom also has seven
- By sharing electrons
- Both end with full orbitals

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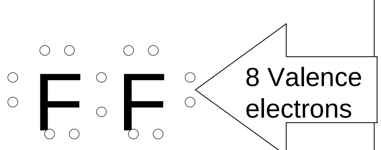

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Covalent bonding

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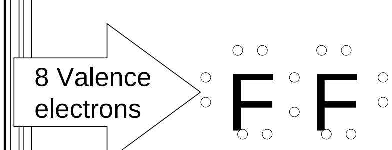

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Single Covalent Bond

- A sharing of two valence electrons.
- Only nonmetals and Hydrogen.
- Different from an ionic bond because they actually form molecules.
- Two specific atoms are joined.
- In an ionic solid you can’t tell which atom the electrons moved from or to.

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How to show how they formed

- It’s like a jigsaw puzzle.
- I have to tell you what the final formula is.
- You put the pieces together to end up with the right formula.
- For example- show how water is formed with covalent bonds.

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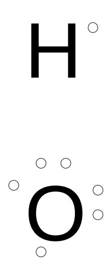
Water



Each hydrogen has 1 valence electron

and wants 1 more

The oxygen has 6 valence electrons

and wants 2 more

They share to make each other “happy”

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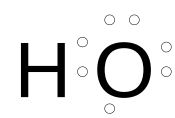
Water

- Put the pieces together
- The first hydrogen is happy
- The oxygen still wants one more



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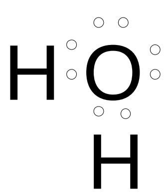
Water

- The second hydrogen attaches
- Every atom has full energy levels



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Multiple Bonds

- Sometimes atoms share more than one pair of valence electrons.
- A double bond is when atoms share two pair (4) of electrons.
- A triple bond is when atoms share three pair (6) of electrons.

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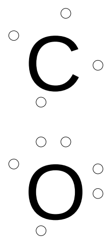
Carbon dioxide



- CO_2 - Carbon is central atom (I have to tell you)
- Carbon has 4 valence electrons
- Wants 4 more
- Oxygen has 6 valence electrons
- Wants 2 more

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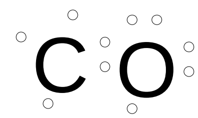
Carbon dioxide

- Attaching 1 oxygen leaves the oxygen 1 short and the carbon 3 short



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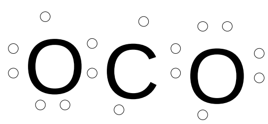
Carbon dioxide

- Attaching the second oxygen leaves both oxygen 1 short and the carbon 2 short



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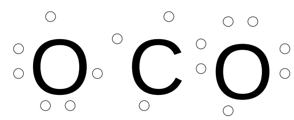
Carbon dioxide

- The only solution is to share more



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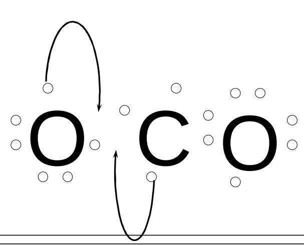

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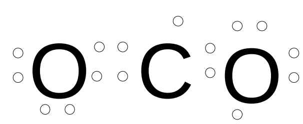

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Carbon dioxide

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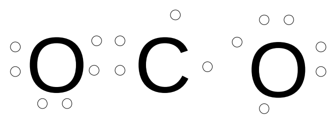

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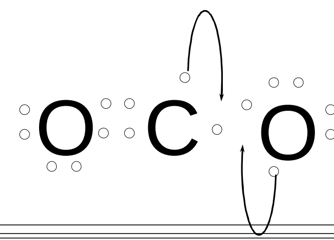
Carbon dioxide

- The only solution is to share more



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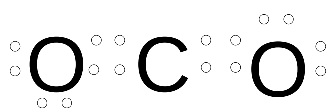
Carbon dioxide

- The only solution is to share more



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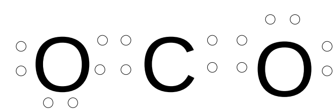
Carbon dioxide

- The only solution is to share more
- Requires two double bonds
- Each atom gets to count all the atoms in the bond



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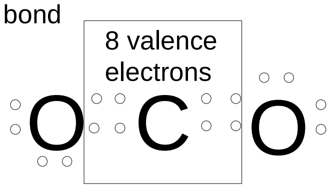

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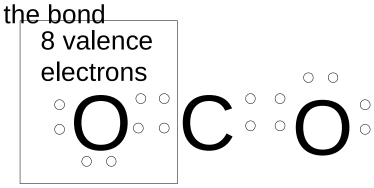

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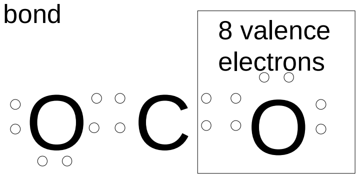

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How to draw them

- To figure out if you need multiple bonds
- Add up all the valence electrons.
- Count up the total number of electrons to make all atoms happy.
- Subtract.
- Divide by 2
- Tells you how many bonds - draw them.
- Fill in the rest of the valence electrons to fill atoms up.

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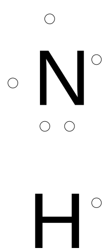
Examples



- NH_3
- N - has 5 valence electrons wants 8
- H - has 1 valence electrons wants 2
- NH_3 has 5+3(1) = 8
- NH_3 wants 8+3(2) = 14
- (14-8)/2= 3 bonds
- 4 atoms with 3 bonds

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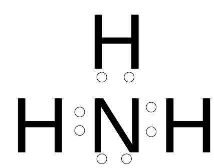
Examples

- Draw in the bonds
- All 8 electrons are accounted for
- Everything is full



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Examples

- HCN C is central atom
- N - has 5 valence electrons wants 8
- C - has 4 valence electrons wants 8
- H - has 1 valence electrons wants 2
- HCN has 5+4+1 = 10
- HCN wants 8+8+2 = 18
- (18-10)/2= 4 bonds
- 3 atoms with 4 bonds -will require multiple bonds - not to H

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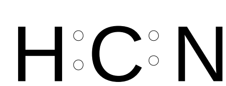
HCN

- Put in single bonds
- Need 2 more bonds
- Must go between C and N



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HCN

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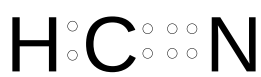
- Uses 8 electrons - 2 more to add



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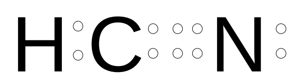
HCN

- Put in single bonds
- Need 2 more bonds
- Must go between C and N
- Uses 8 electrons - 2 more to add
- Must go on N to fill octet



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Where do bonds go?

- Think of how many electrons they are away from noble gas.
- H should form 1 bond- always
- O should form 2 bonds – if possible
- N should form 3 bonds – if possible
- C should form 4 bonds– if possible

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Practice

- Draw electron dot diagrams for the following.
- PCl_3
- H_2O_2
- CH_2O
- C_3H_6

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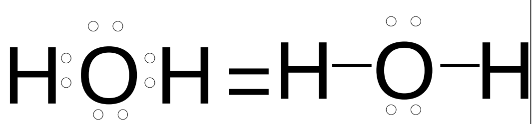
Another way of indicating bonds

- Often use a line to indicate a bond
- Called a structural formula
- Each line is 2 valence electrons



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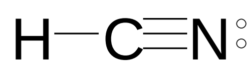
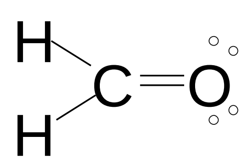
Structural Examples



- C has 8 electrons because each line is 2 electrons
- Ditto for N



- Ditto for C here
- Ditto for O

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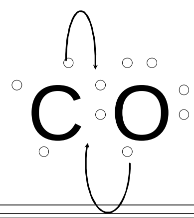
Coordinate Covalent Bond

- When one atom donates both electrons in a covalent bond.
- Carbon monoxide
- CO



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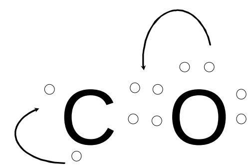

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Coordinate Covalent Bond

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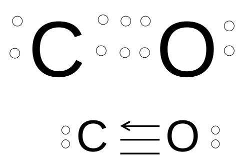

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How do we know if

- Have to draw the diagram and see what happens.
- Often happens with polyatomic ions
- If an element has the wrong number of bonds

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Polyatomic ions

- Groups of atoms held by covalent bonds, with a charge
- Can’t build directly, use (happy-have)/2
- Have number will be different
- Surround with [ ], and write charge
- NH_4^+
- SO_3^{2-}

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Resonance

- When more than one dot diagram with the same connections is possible.
- Choice for double bond
- NO_2^{1-}
- Which one is it?
- Does it go back and forth?
- Double bonds are shorter than single
- In NO_2^{1-} all the bonds are the same length

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Resonance

- It is a mixture of both, like a mule.
- CO_3^{2-}

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Bond Dissociation Energy

- The energy required to break a bond
- C - H + 393 kJ $\longrightarrow$ C + H
- Double bonds have larger bond dissociation energies than single
- Triple even larger
  - C-C 347 kJ
  - C=C 657 kJ
  - C≡C 908 kJ

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Bond Dissociation Energy

- The larger the bond energy, the harder it is to break
- Large bond energies make chemicals less reactive.

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VSEPR

- Valence Shell Electron Pair Repulsion.
- Predicts three dimensional geometry of molecules.
- Name tells you the theory.
- Valence shell - outside electrons.
- Electron Pair repulsion - electron pairs try to get as far away as possible.
- Can determine the angles of bonds.
- And the shape of molecules

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VSEPR

- Based on the number of pairs of valence electrons both bonded and unbonded.
- Unbonded pair are called lone pair.
- CH_4 - draw the structural formula

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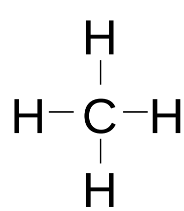
VSEPR



- Single bonds fill all atoms.
- There are 4 pairs of electrons pushing away.
- The furthest they can get away is 109.5º.

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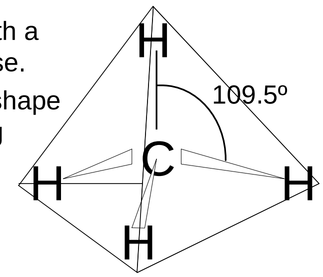
4 atoms bonded

- Basic shape is tetrahedral.
- A pyramid with a triangular base.
- Same basic shape for everything with 4 pairs.



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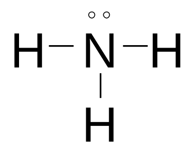
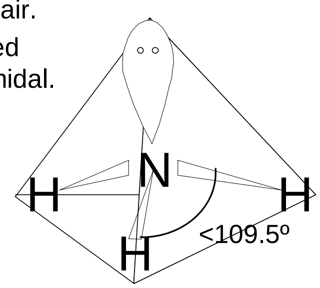
3 bonded - 1 lone pair

- Still basic tetrahedral but you can't see the electron pair.
- Shape is called trigonal pyramidal.





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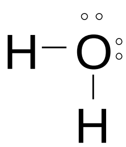
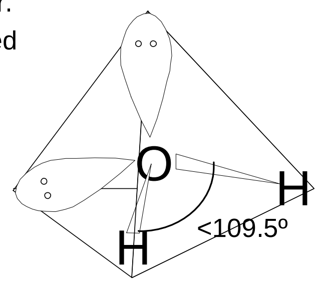
2 bonded - 2 lone pair

- Still basic tetrahedral but you can't see the 2 lone pair.
- Shape is called bent.





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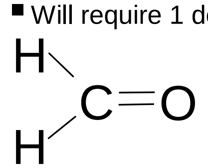
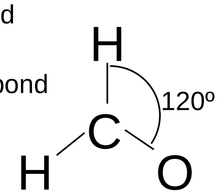
3 atoms no lone pair

- The farthest you can the electron pair apart is 120º.
- Shape is flat and called trigonal planar.
- Will require 1 double bond





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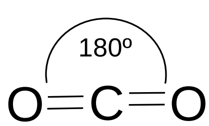
2 atoms no lone pair

- With three atoms the farthest they can get apart is 180º.
- Shape called linear.
- Will require 2 double bonds or one triple bond



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Number of things coming off the central atom	Atoms attached to central atom	Lone Pair on central atom	Shape	Bond angle
4				
4				
4				
3				
3				
2				

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Molecular Orbitals

- The overlap of atomic orbitals from separate atoms makes molecular orbitals
- Each molecular orbital has room for two electrons
- Two types of MO
  - Sigma (σ) between atoms
  - Pi (π) above and below atoms

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Sigma bonding orbitals

- From s orbitals on separate atoms

s orbital s orbital

Sigma bonding molecular orbital

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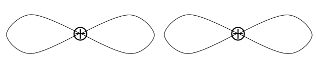
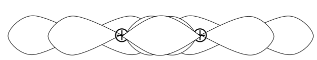
Sigma bonding orbitals

- From p orbitals on separate atoms



p orbital p orbital



Sigma bonding molecular orbital

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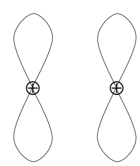
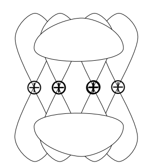
Pi bonding orbitals

- p orbitals on separate atoms





Pi bonding molecular orbital

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Sigma and pi bonds

- All single bonds are sigma bonds
- A double bond is one sigma and one pi bond
- A triple bond is one sigma and two pi bonds.

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Hybrid Orbitals

Combines bonding with geometry

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Hybridization

- The mixing of several atomic orbitals to form the same number of hybrid orbitals.
- All the hybrid orbitals that form are the same.
- sp^3 -1 s and 3 p orbitals mix to form 4 sp^3 orbitals.
- sp^2 -1 s and 2 p orbitals mix to form 3 sp^2 orbitals leaving 1 p orbital.
- sp -1 s and 1 p orbitals mix to form 2 sp orbitals leaving 2 p orbitals.

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Hybridization

- 109.5º with s and p
- Need 4 orbitals.
- We combine one s orbital and 3 p orbitals.
- Make sp^3 hybrid
- sp^3 hybridization has tetrahedral geometry.

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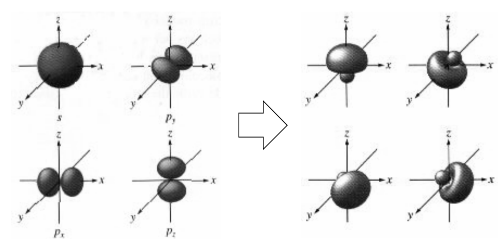

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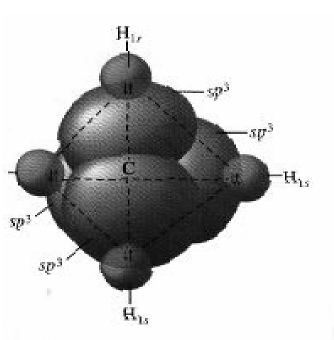

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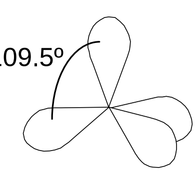
sp^3 geometry

- This leads to tetrahedral shape.
- Every molecule with a total of 4 atoms and lone pair is sp^3 hybridized.
- Gives us trigonal pyramidal and bent shapes also.



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How we get to hybridization

- We know the geometry from experiment.
- We know the orbitals of the atom
- hybridizing atomic orbitals can explain the geometry.
- So if the geometry requires a 109.5º bond angle, it is sp^3 hybridized.

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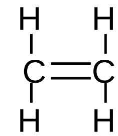
sp^2 hybridization



- C_2H_4
- double bond counts as one pair
- Two trigonal planar sections
- Have to end up with three blended orbitals
- use one s and two p orbitals to make sp^2 orbitals.
- leaves one p orbital perpendicular

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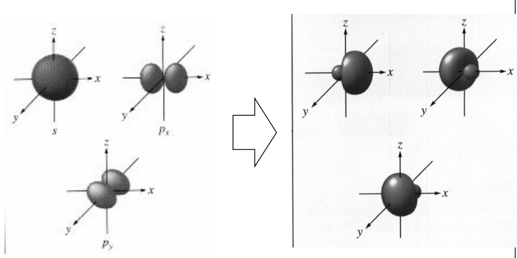

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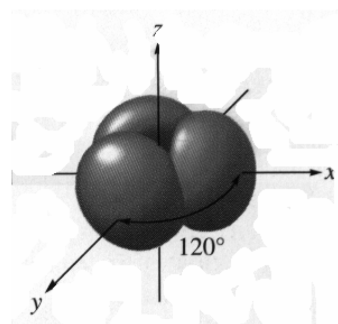

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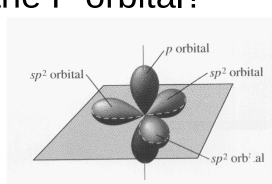
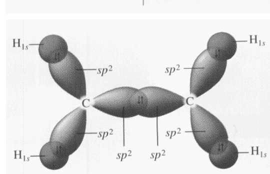
Where is the P orbital?

- Perpendicular
- The overlap of orbitals makes a sigma bond (σ bond)





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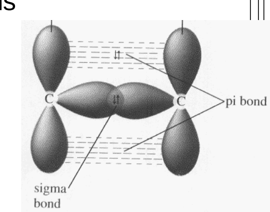
Two types of Bonds

- Sigma bonds (σ) from overlap of orbitals
- between the atoms
- Pi bond (π bond) between p orbitals.
- above and below atoms
- All single bonds are σ bonds
- Double bond is 1 σ and 1 π bond
- Triple bond is 1 σ and 2 π bonds



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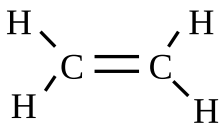

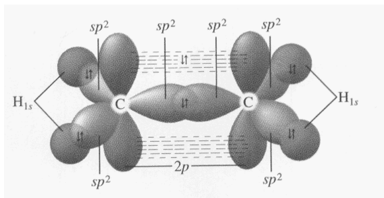

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sp^2 hybridization

- when three things come off atom
- trigonal planar
- 120º
- one π bond

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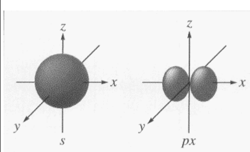
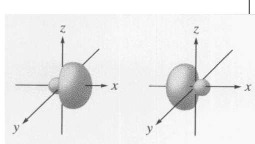
What about two

- when two things come off
- one s and one p hybridize
- linear





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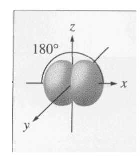
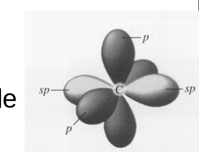
sp hybridization

- end up with two lobes 180º apart.
- p orbitals are at right angles
- makes room for two π bonds and two sigma bonds.
- a triple bond or two double bonds





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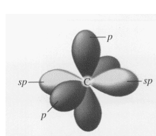
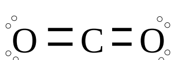
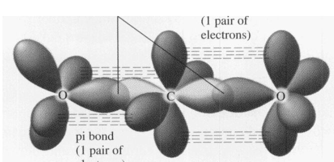
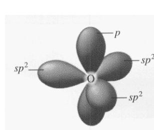
CO_2

- C can make two σ and two π
- O can make one σ and one π









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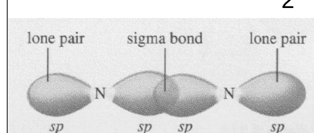
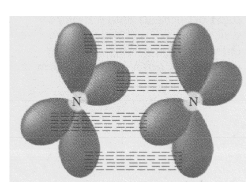
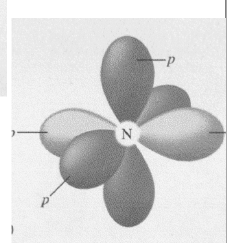
N_2







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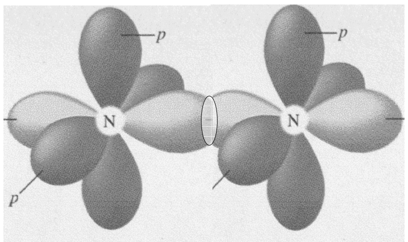
N_2



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Polar Bonds

- When the atoms in a bond are the same, the electrons are shared equally.
- This is a nonpolar covalent bond.
- When two different atoms are connected, the electrons may not be shared equally.
- This is a polar covalent bond.
- How do we measure how strong the atoms pull on electrons?

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Electronegativity

- A measure of how strongly the atoms attract electrons in a bond.
- The bigger the electronegativity difference the more polar the bond.
- Use table 6.2 Pg. 177
- 0.0 - 0.4 Covalent nonpolar
- 0.5 - 1.0 Covalent moderately polar
- 1.0 -2.0 Covalent polar
- >2.0 Ionic

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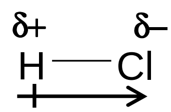
How to show a bond is polar

- Isn't a whole charge just a partial charge
- δ+ means a partially positive
- δ– means a partially negative



- The Cl pulls harder on the electrons
- The electrons spend more time near the Cl

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Polar Molecules

Molecules with ends

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Polar Molecules

- Molecules with a partially positive end and a partially negative end
- Requires two things to be true

① The molecule must contain polar bonds

This can be determined from differences in electronegativity.

②Symmetry can not cancel out the effects of the polar bonds.

Must determine geometry first.

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Polar Molecules

- Symmetrical shapes are those without lone pair on central atom
  - Tetrahedral
  - Trigonal planar
  - Linear
- Will be nonpolar if all the atoms are the same
- Shapes with lone pair on central atom are not symmetrical
- Can be polar even with the same atom

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Is it polar?

- HF
- H_2O
- NH_3
- CCl_4
- CO_2
- CH_3Cl

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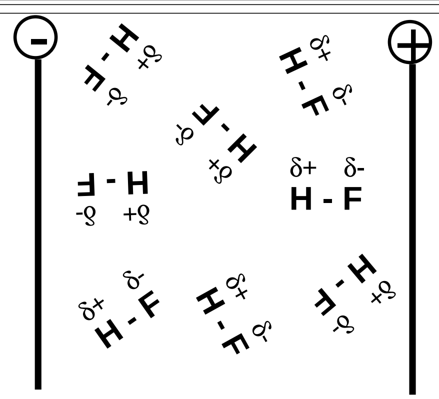

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Intermolecular Forces

What holds molecules to each other

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Intermolecular Forces

- They are what make solid and liquid molecular compounds possible.
- The weakest are called van der Waal's forces - there are two kinds
  - Dispersion forces
  - Dipole Interactions

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Dipole interactions

- Occur when polar molecules are attracted to each other.
- Partial positive on one molecule attracted to partial negative on another
- Opposites attract but not completely hooked like in ionic solids.

$$\overset{\delta^+}{H}—\overset{\delta^-}{F} \longleftrightarrow \overset{\delta^+}{H}—\overset{\delta^-}{F}$$

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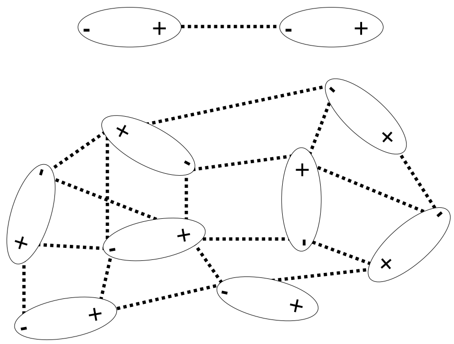

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Dispersion Force

- Electrons are not evenly distributed at every instant in time.
- Temporary partial charges
- An momentary dipole.
- Affects the electrons ion the molecule next to it.
- Called induced dipole
- Momentarily attracted

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Dispersion force

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Dispersion Force

- Depends only on the number of electrons in the molecule
- Bigger molecules more electrons
- More electrons stronger forces
  - F_2 is a gas
  - Br_2 is a liquid
  - I_2 is a solid

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Hydrogen bonding

- Are the attractive force caused by hydrogen bonded to F, O, or N.
- F, O, and N are very electronegative so it is a very strong dipole.
- They are small, so molecules can get close together
- The hydrogen partially share with the lone pair in the molecule next to it.
- The strongest of the intermolecular forces.

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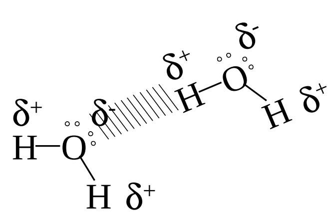
Hydrogen Bonding



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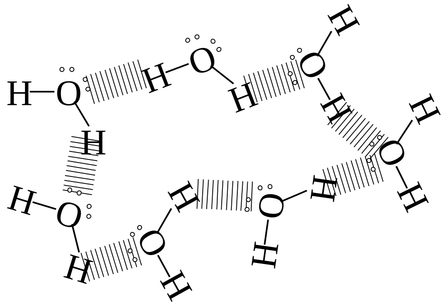
Hydrogen bonding



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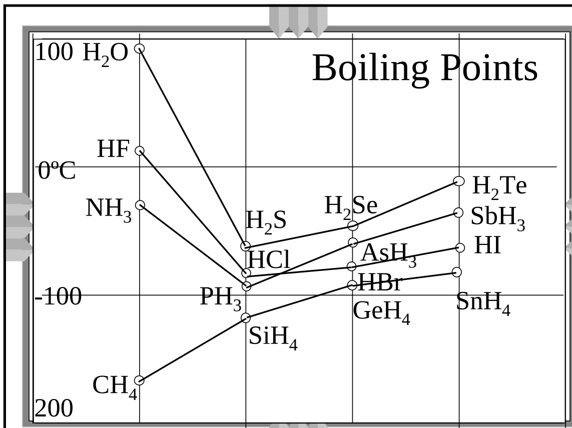
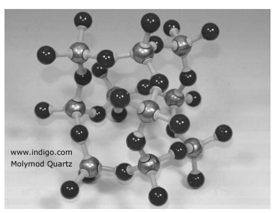


Network solids

- Compounds an elements where every atom is covalently bonded to each other.
- One big molecule.
- SiO_2 Quartz
- Diamond
- Graphite
- Glass



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Network solids

- Covalent bonds are much stronger than intermolecular forces
- So network solids are
  - Hard
  - Have high melting points

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Properties of Molecular Compounds

- Made of nonmetals
- Poor or nonconducting as solid, liquid or aqueous solution
- Low melting point
- Two kinds of crystals
  - Molecular solids held together by IMF
  - Network solids- held together by bonds

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