# 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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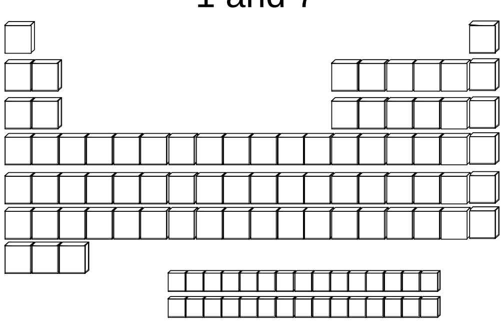
## 1 and 7



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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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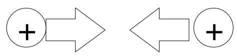
## How does $H_2$ form?

- The nuclei repel



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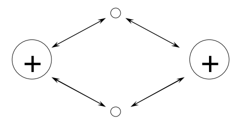
## How does $H_2$ form?

- 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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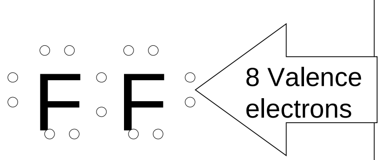


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## Covalent bonding

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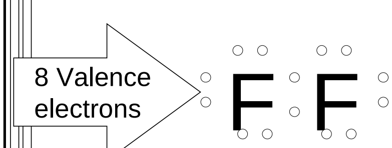


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## Single Covalent Bond

- 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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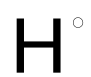
## Water



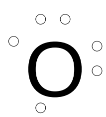

Each hydrogen has 1 valence electron

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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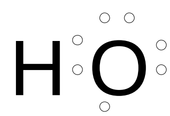
## Water

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

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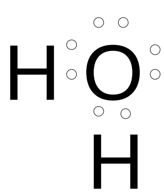
## Water

- 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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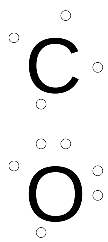
## Carbon dioxide




- $CO_2$ - 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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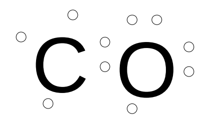
## Carbon dioxide

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

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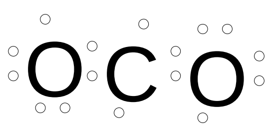
## Carbon dioxide

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

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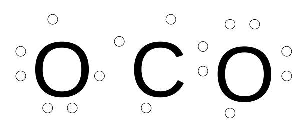
## Carbon dioxide

- The only solution is to share more

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## Carbon dioxide

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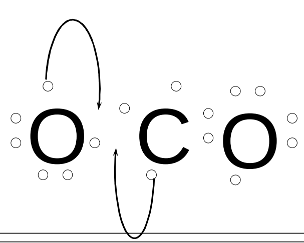


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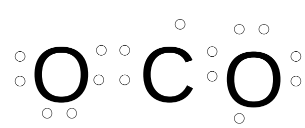


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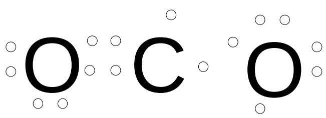

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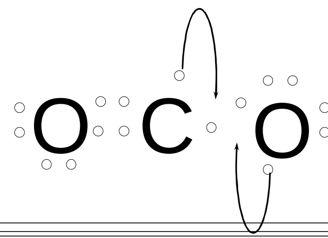

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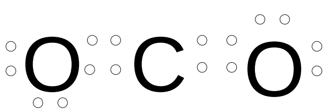
## Carbon dioxide

- The only solution is to share more



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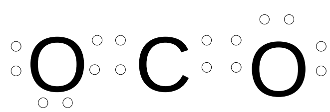
## Carbon dioxide

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



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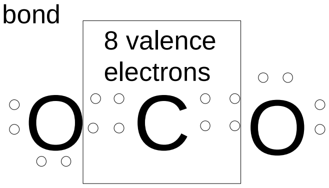


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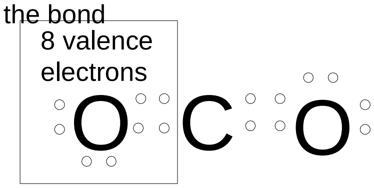


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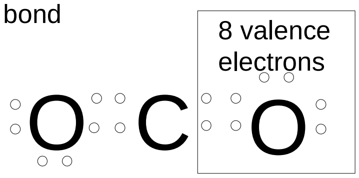


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## How to draw them

- 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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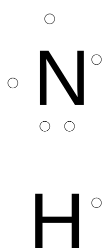
## Examples




- $NH_3$
- 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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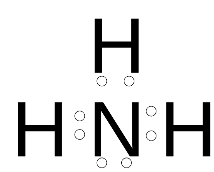
## Examples

- 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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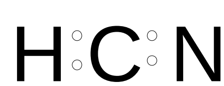
## HCN

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

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## HCN

- Put in single bonds
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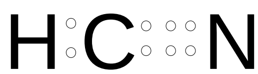


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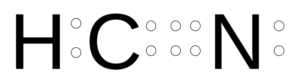
## HCN

- 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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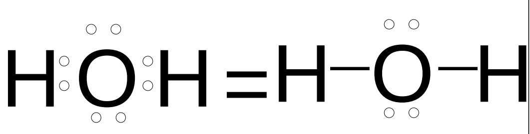
## Another way of indicating bonds

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

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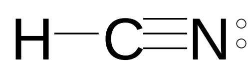
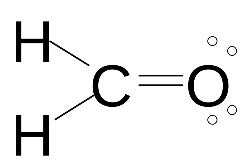
## Structural Examples




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




- Ditto for C here
- Ditto for O

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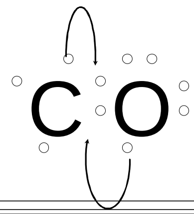
## Coordinate Covalent Bond

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

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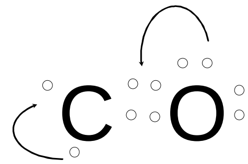
## Coordinate Covalent Bond

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

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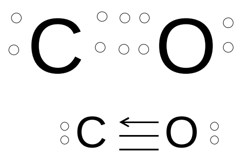


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## How do we know if

- 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 ⟶ 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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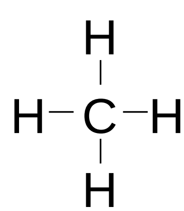
## VSEPR



- 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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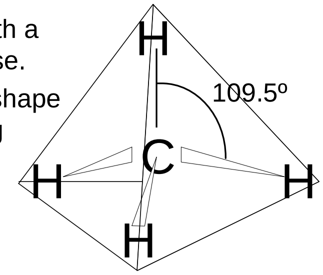
## 4 atoms bonded

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

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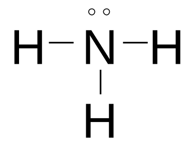
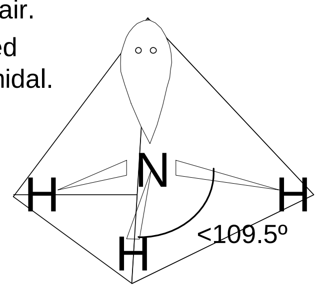
## 3 bonded - 1 lone pair

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

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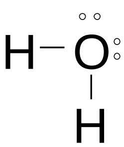
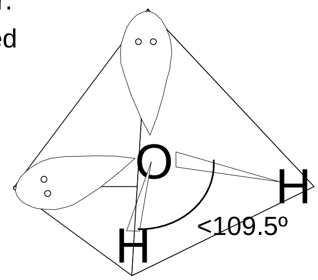
## 2 bonded - 2 lone pair

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

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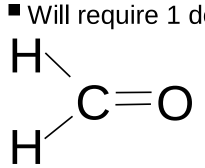
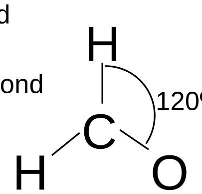
## 3 atoms no lone pair

- 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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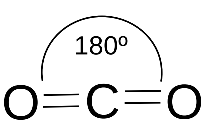
## 2 atoms no lone pair

- 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 ( $\sigma$ ) between atoms
  - Pi ( $\pi$ ) 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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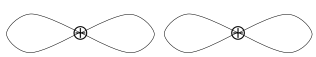
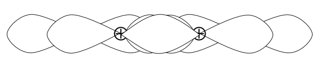
## Sigma bonding orbitals

- From p orbitals on separate atoms



p orbital p orbital



Sigma bonding molecular orbital

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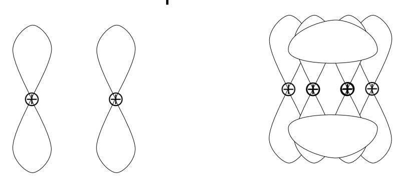
## Pi bonding orbitals

- 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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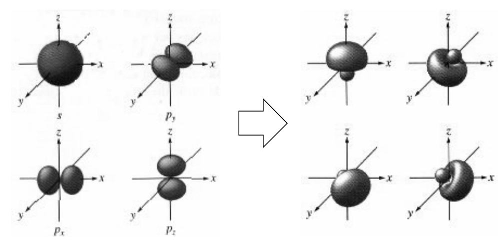

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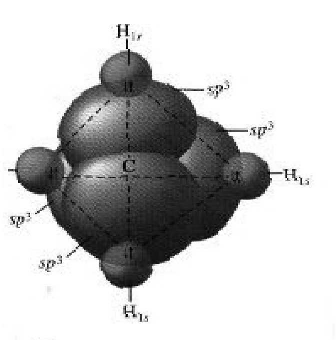


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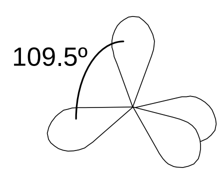
## $sp^3$ geometry

- 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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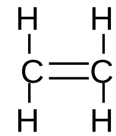
## $sp^2$ hybridization




- $C_2H_4$
- 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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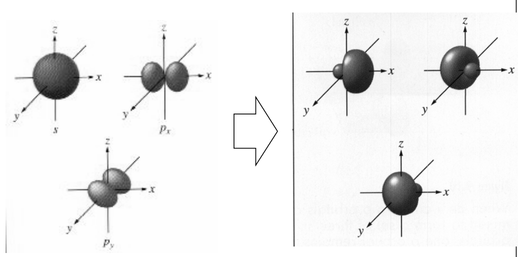

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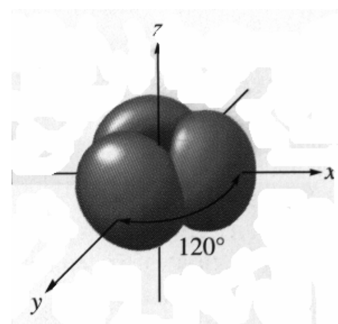


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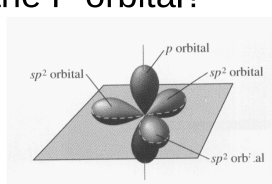
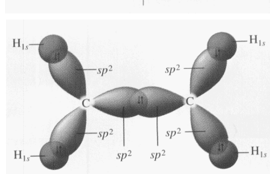
## Where is the P orbital?

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



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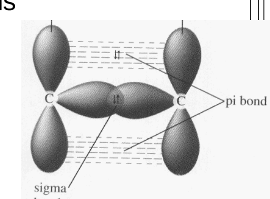
## Two types of Bonds

- 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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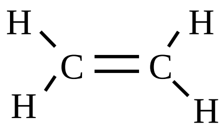

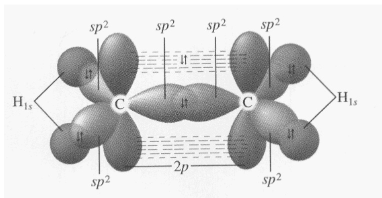


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## $sp^2$ hybridization

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

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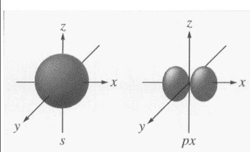
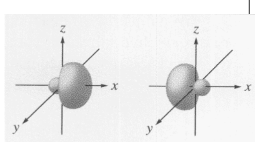
## What about two

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



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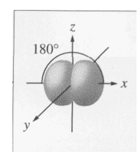
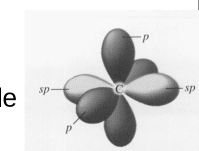
## sp hybridization

- 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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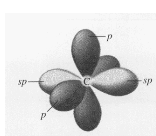
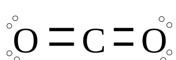
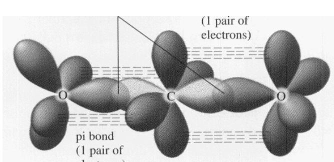
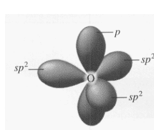
## $CO_2$

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







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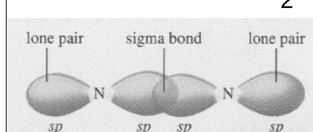
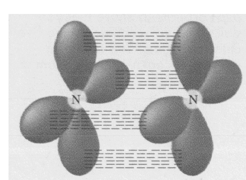
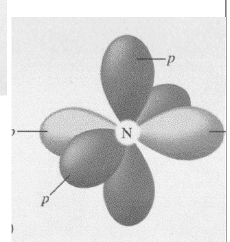
## $N_2$



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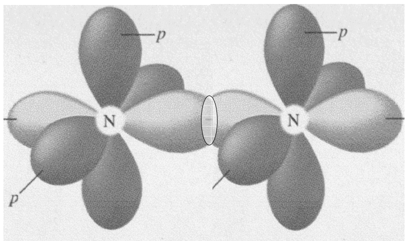
## $N_2$

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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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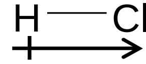
## How to show a bond is polar

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

δ+ δ–

H — Cl



- 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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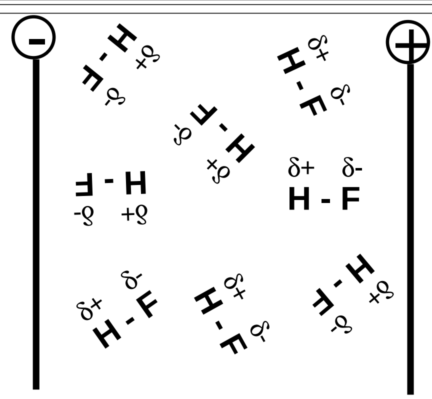


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## Intermolecular Forces

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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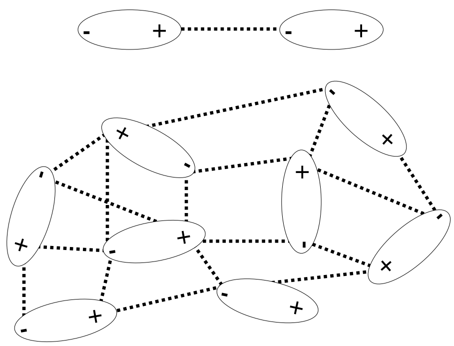

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## Dispersion Force

- 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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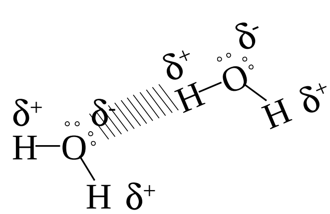
## Hydrogen Bonding

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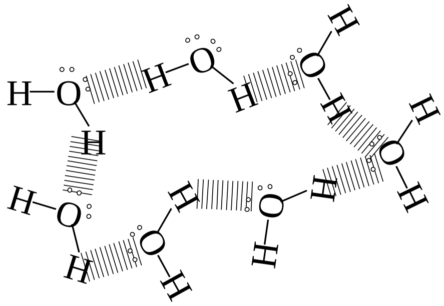
## Hydrogen bonding

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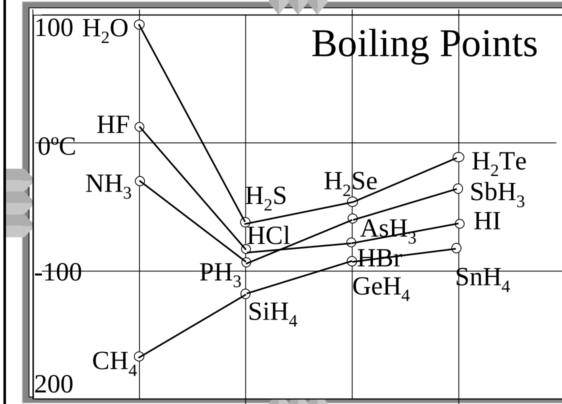
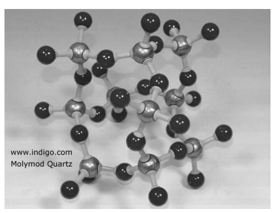



## Network solids

- 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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