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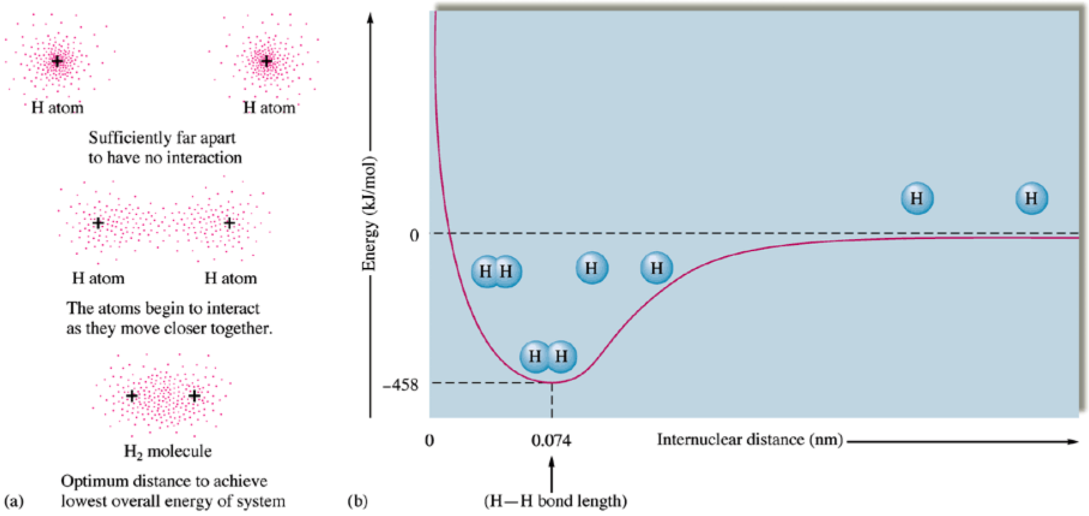
Chapter 8 Outline – Bonding: General Concepts

Types of Chemical Bonds

- Information about the strength of a bonding interaction is obtained by measuring the **bond energy**, which is the energy required to break the bond.
- Ionic bonding occurs between an element that loses electrons easily reacts with an atom that has a high affinity for electrons. Ionic compounds are formed between a metal and a non-metal.
- The energy of interaction between a pair of ions can be calculated using Coulomb's law in the form:
  $E = (2.31 \times 10^{-19}\ J * nm)(\frac{Q_1Q_2}{r})$; E has units of joules, r is the distance between the ion centers in nanometers, and Q_1 and Q_2 are the numerical ion charges.
- For example: in solid NaCl, the distance between Na^+ and Cl^- ions is 2.76 Å (0.276 nm), and the ionic energy per pair of ions is: $E = (2.31 \times 10^{-19}\ J * nm)(\frac{(+1)(-1)}{0.276}) = -8.37 \times 10^{-19}\ J$ where the negative sign equals an attractive force. The ion pair has lower energy than the separated ions.
- A bond will form if the system can lower its total energy in the process.
- In a bond between two hydrogen atoms, the system will act to minimize the sum of the positive (repulsive) energy terms and the negative (attractive) energy term. The distance where the energy is minimal is called the bond length.

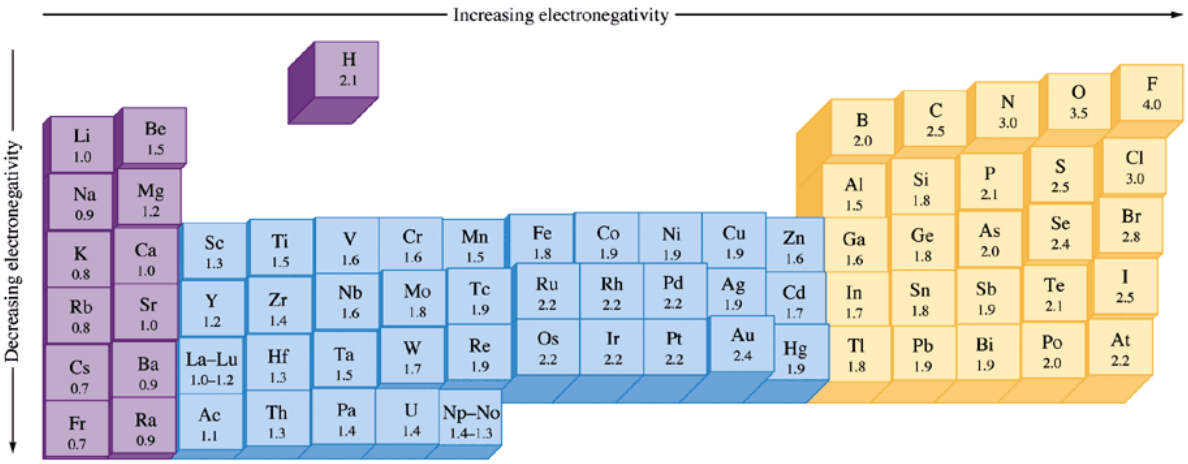


- The zero point of energy is defined with atoms at infinite separation.
- At very short distances the energy rises steeply because of the importance of the repulsive forces when atoms are very close together.
- The bond length is the distance at which the system has minimum energy.
- H_2 is a non-polar covalent bond where the two hydrogen nuclei share the valence electrons equally.
- HF exhibits a polar-covalent bond where the electrons are shared unequally, showing an attraction to the more electronegative element.

Electronegativity

- Electronegativity is the ability of an atom in a molecule to attract shared electron.
- The widely accepted method for determining electronegativity values is that of Linus Pauling (1901-1995), an American scientist who won the Nobel Prizes for both chemistry and peace.

- Electronegativity values of the elements are shown below. Note that electronegativity tends to increase across a period. Electronegativity tends to decrease down a group. Fluorine has the highest electronegativity (4.0) and cesium has the lowest electronegativity (0.7).



- The relationship between electronegativity difference and bond type is shown to the right.

- Practice Problem: Order the following bonds according to polarity: H-H, O-H, Cl-H, S-H and F-H.

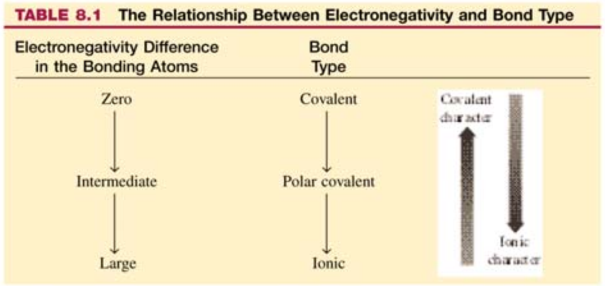
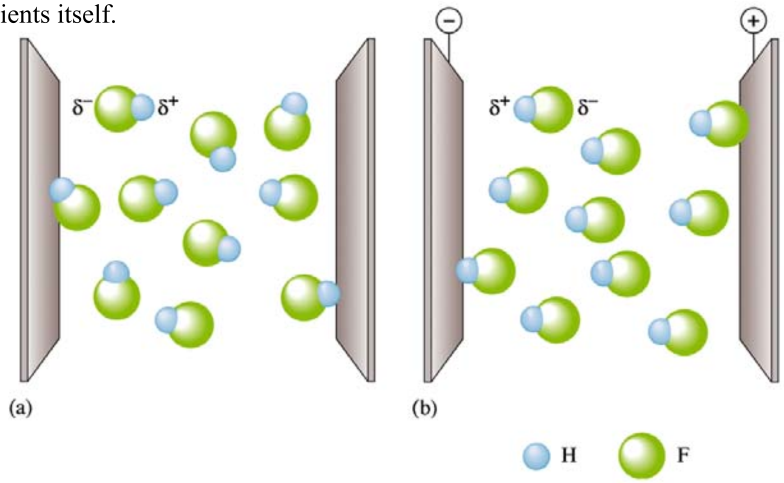


TABLE 8.1 The Relationship Between Electronegativity and Bond Type

Electronegativity Difference in the Bonding Atoms	Bond Type
Zero	Covalent
Intermediate	Polar covalent
Large	Ionic

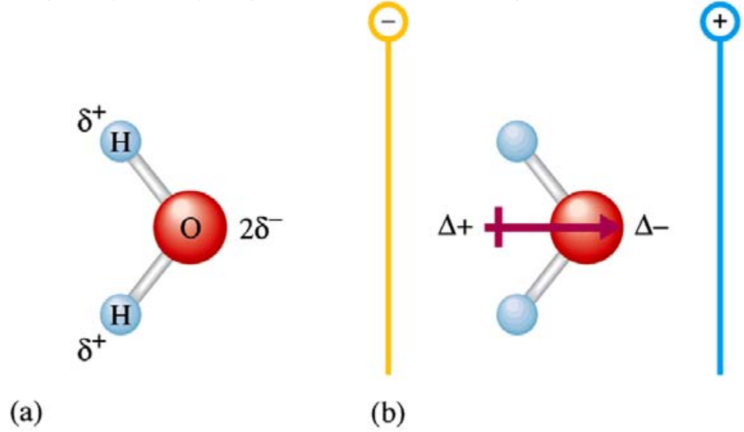
Bond Polarity and Dipole Moments

- When HF is placed in an electrical field, the molecule orients itself so that its electropositive end is facing the negative plate and its electronegative end is facing the positive plate.
- In diagram (a) below, the plates are not charged. In diagram (b) below, the plates are charged and HF orients itself.

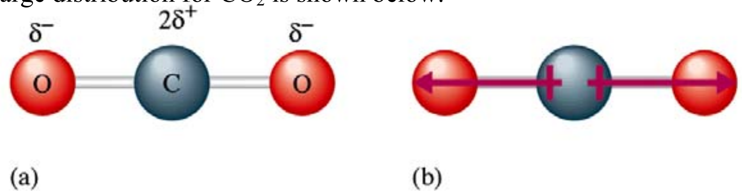


- A molecule such as HF that has a center of positive charge and a center of negative charge is said to be dipolar, or to have a dipolar moment.
- A dipole moment is shown using an arrow pointing to the more electronegative element. The electropositive end of the arrow has a cross through it. Copy here:

- Any diatomic (two-atom, but not necessarily the same atoms) molecule that has a polar bond also will show a dipole moment.
- Polyatomic molecules can also exhibit dipolar behavior. For example, in water, oxygen has a higher electronegativity than hydrogen. The molecular charge distribution is shown below.



- Some molecules have polar bonds but do not have a dipole moment. This occurs when the individual bond polarities are arranged in such a way that they cancel each other out. CO_2 is an example of this behavior. The charge distribution for CO_2 is shown below.



- SO_3 is also a molecule with polar bonds but overall no dipole moment. This is due to its molecular shape. More will be covered on molecular shapes later this chapter.

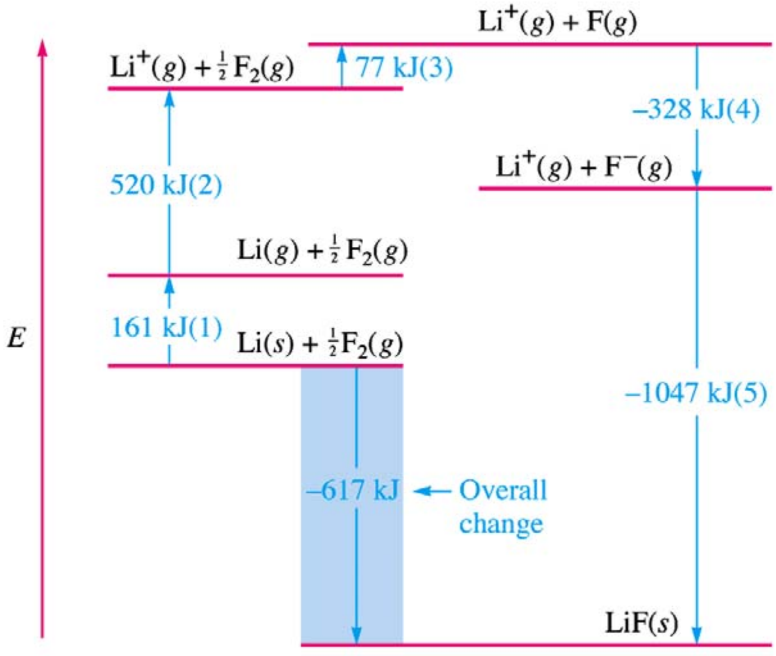
Ions: Electron Configuration and Sizes

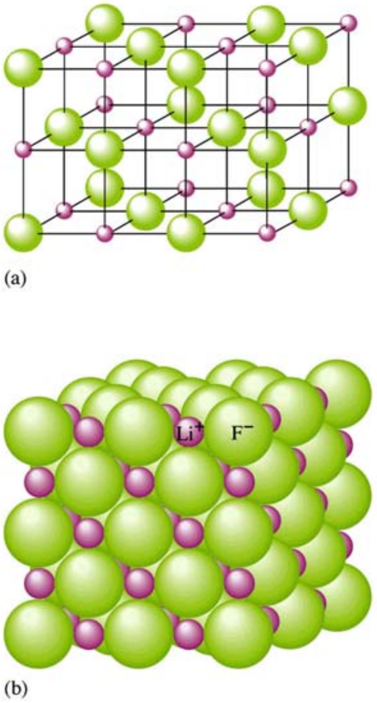
- In virtually every case, the atoms in a stable compound have a noble gas arrangement of electrons.
- When two nonmetals react to form a covalent bond, they share electrons in a way that completes the valence electron configurations of both atoms. Both nonmetals attain noble gas electron configurations.
- When a nonmetal and a representative-group metal react to form a binary ionic compound, the ions form so that the valence electron configuration of the nonmetal achieves the electron configuration of the next noble gas atom and the valence orbitals of the metal are emptied.
- There are some important exceptions to the rules above.
    - Sn forms Sn^{2+} and Sn^{4+}
    - Pb forms Pb^{2+} and Pb^{4+}
    - Bi forms Bi^{3+} and Bi^{5+}
    - Tl forms Tl^{+} and Tl^{3+}
- Ion size plays an important role in determining the structure and stability of ionic solids, the properties of ions in aqueous solution, and the biological effects of ions.
- Since a positive ion is formed by removing one or more electrons from a neutral atom, the resulting cation is smaller than its parent atom. The opposite is true for negative ions; the addition of electrons to a neutral atom produces an anion significantly larger than its parent atom.
- Ion size increases going down a group. Horizontally, the trend is complicated because metals on the left side of the periodic table form cations, and nonmetals on the right side of the periodic table form anions.
- One trend worth noting involves relative sizes of a set of isoelectronic ions – ions containing the same number of electrons.
- Consider the ions O^{2-}, F^{-}, Na^{+} , Mg^{2+} and Al^{3+}.

  - Since all species have the same number of electrons, it can be assumed that as the number of protons increases, the attraction between protons and electrons will increase and make the ionic size smaller.
- **In summary, for isoelectronic ions, the greater the atomic number, the smaller the ionic radius.**
- You try: Arrange the ions Se^{2-}, Br^-, Rb^+ and Sr^{2+} in order of decreasing size.

Formation of Binary Ionic Compounds

- Ionic compounds form because the aggregated oppositely charged ions have a lower energy that the original elements.
- Just how strongly the ions attract each other in the solid state is indicated by the lattice energy – *the change in energy that takes place when separated gaseous ions are packet together to form an ionic solid:*
  - $M^+(g) + X^-(g) \rightarrow MX(s)$
- Lattice energy is often defined as the amount of energy released when an ionic solid forms from its ions, so lattice energy will have a negative value.
- Consider the energy change associated with the following reaction: $Li(s) + \frac{1}{2} F_2(g) \rightarrow LiF(s)$
  - Step 1: Sublimation of solid lithium. $Li(s) \rightarrow Li(g)$ $\Delta E = 161$ kJ/mol
  - Step 2: Ionization of lithium atom. $Li(g) \rightarrow Li^+(g) + e^-$ $\Delta E = 520$ kJ/mol
  - Step 3: Dissociation of fluorine molecules. $\frac{1}{2} F_2(g) \rightarrow F(g)$ $\Delta E = 77$ kJ/mol
  - Step 4: Formation of fluoride ions. $F(g) + e^- \rightarrow F^-(g)$ $\Delta E = -328$ kJ/mol
  - Step 5: Formation of LiF from $Li^+(g)$ and $F^-(g)$ ions. $Li^+(g) + F^-(g) \rightarrow LiF(s)$ $\Delta E = -1047$ kJ/mol
  - By adding the ΔE values, the $\Delta E_{overall} = -617$ kJ/mol of LiF. This is an exothermic reaction.





- The lattice structure of lithium fluoride is shown above. All the binary ionic compounds formed by an alkali metal(except cesium) and a halogen have the same lattice structure as lithium fluoride. It is sometimes called the sodium chloride structure.

Partial Ionic Character of Covalent Bonds

- When atoms with different electronegativities react to form molecules, the electrons are not shared equally. The possible result is a polar covalent bond or, if the electronegativity difference is great enough, a complete transfer of one or more electrons to form ions.
- Percent ionic character of a bond = (measured dipole moment of X—Y) ÷ (calculated dipole moment of X^+Y^-) x 100%
- Generally compounds with more than 50% ionic character are normally considered to be ionic solids.
- Since compounds such as ammonium chloride and sodium sulfate contain both ionic and covalent bonds, ionic compounds will be defined as any compound that conducts an electric current when melted.

Covalent Bond Energies and Chemical Reactions

- Consider the stepwise decomposition of methane:
  - $CH_4(g) \rightarrow CH_3(g) + H(g)$ energy required: 435 kJ/mol
  - $CH_3(g) \rightarrow CH_2(g) + H(g)$ energy required: 453 kJ/mol
  - $CH_2(g) \rightarrow CH(g) + H(g)$ energy required: 425 kJ/mol
  - $CH(g) \rightarrow C(g) + H(g)$ energy required: 339 kJ/mol
  - Total = 1652 ÷ 4 = 413 kJ/mol
  - Note that the C—H bond is somewhat sensitive to its environment
- Consider the following molecules and the measured C—H bond energy (kJ/mol)

Molecule	Measure C-H bond energy (kJ/mol)
$HCBr_3$	380
$HCCl_3$	380
HCF_3	430
C_2H_6	410

- Again, C—H bond strength varies significantly with its environment, but the concept of an average bond energy is helpful to chemists.
- The average bond energies for various types of bonds are listed below.

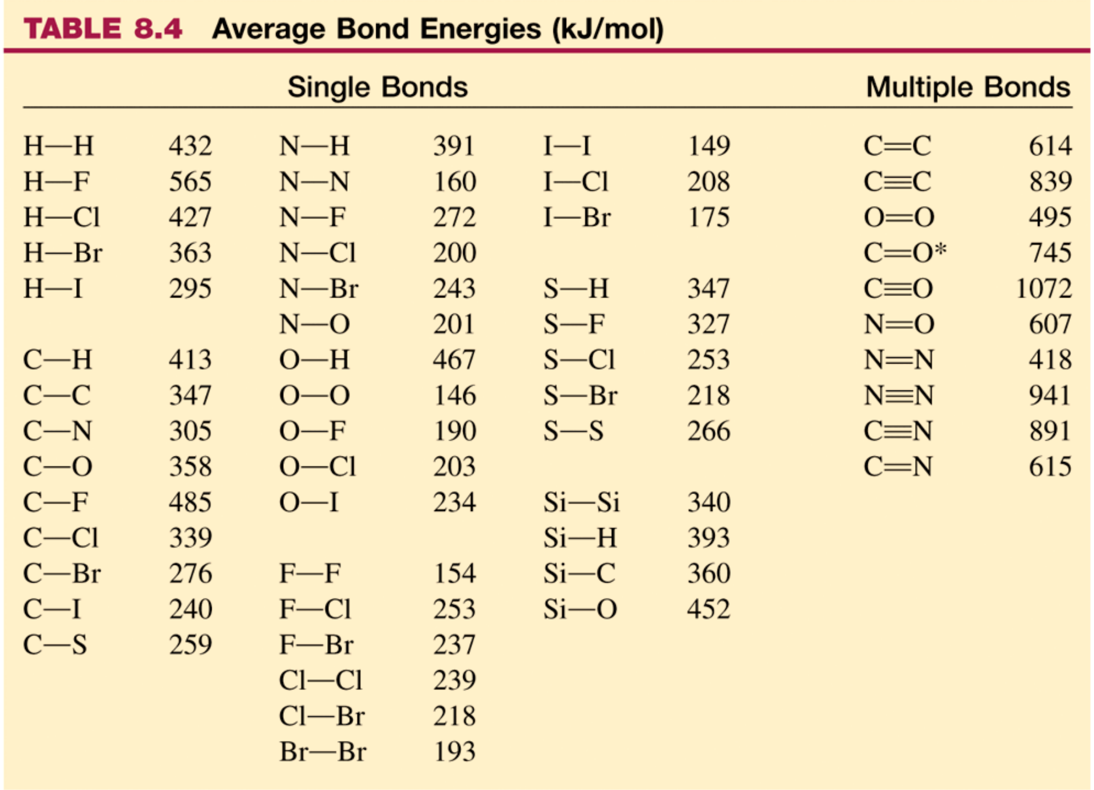


TABLE 8.4 Average Bond Energies (kJ/mol)

Single Bonds						Multiple Bonds	
H—H	432	N—H	391	I—I	149	C=C	614
H—F	565	N—N	160	I—Cl	208	C≡C	839
H—Cl	427	N—F	272	I—Br	175	O=O	495
H—Br	363	N—Cl	200			C=O*	745
H—I	295	N—Br	243	S—H	347	C≡O	1072
		N—O	201	S—F	327	N=O	607
C—H	413	O—H	467	S—Cl	253	N=N	418
C—C	347	O—O	146	S—Br	218	N≡N	941
C—N	305	O—F	190	S—S	266	C≡N	891
C—O	358	O—Cl	203			C=N	615
C—F	485	O—I	234	Si—Si	340		
C—Cl	339			Si—H	393		
C—Br	276	F—F	154	Si—C	360		
C—I	240	F—Cl	253	Si—O	452		
C—S	259	F—Br	237				
		Cl—Cl	239				
		Cl—Br	218				
		Br—Br	193				

*C=O(CO_2) = 799

- Note that single, double and triple bonds exist. In a single bond, one pair of electrons is shared, in a double bond, two pairs of electrons are shared and in a triple bond, three pairs of electrons are shared. Single bonds are the longest and weakest of the bonds. Triple bonds are the shortest and the strongest of the bonds.
- For bonds to be broken, energy must be added to a system – an endothermic process.
- **$\Delta H = \Sigma D$ (bonds broken) - ΣD (bonds formed),** Σ represents the sum of terms and D represents the bond energy per mole of bonds
- Example: Using the bond energies from above, calculate the ΔH for the reaction of methane with chlorine and fluorine to give Freon-12 (CF_2Cl_2)
  - $CH_4(g) + 2Cl_2(g) + 2F_2(g) \rightarrow CF_2Cl_2(g) + 2HF(g) + 2HCl(g)$
  - Bonds broken:
    - C—H: 4 x 413 = 1652 kJ
    - Cl—Cl : 2 x 239 = 478 kJ
    - F—F: 2 x 154 kJ = 308 kJ
    - Total = 2438 kJ
  - Bonds formed:
    - C—F: 2 x 485 = 970 kJ
    - C—Cl: 2 x 339 = 678 kJ
    - H—F: 2 x 565 = 1130 kJ
    - H—Cl: 2 x 427 = 854 kJ
    - Total energy released = 3632 kJ
  - $\Delta H = \Sigma D$ (bonds broken) - ΣD (bonds formed)
  - $\Delta H = 2438 \text{ kJ} - 3632 \text{ kJ} = -1194 \text{ kJ}$
- You try: Calculate the ΔH for the reaction below:

  $N{\equiv}N(g) + 3H_2(g) \rightarrow 2NH_3(g)$

The Localized Electron Bonding Model

- The localized electron (LE) model assumes that a molecule is composed of atoms that are bound together by sharing pairs of electrons using atomic orbitals of the bound atoms. Electron pairs in the molecule are assumed to be localized on a particular atom or in the space between two atoms. Those pairs of electrons localized on an atom are called lone pairs, and those found in the space between the atoms are called bonding pairs.

Lewis Structures

- The Lewis structure of a molecule shows how the valence electrons are arranged among the atoms in the molecule.
- Elements tend to follow the octet rule where there are surrounded by eight electrons.
- Hydrogen follows the duet rule and only has two electrons in a covalent bond.
- When writing Lewis structures, do not worry about which electrons come from which atoms in a molecule. The best way to look at a molecule is to regard it as a new entity that uses all available valence electrons on the atoms to achieve the lowest possible energy.

How to Write Lewis Structures

Example: Write the structural formula for diatomic nitrogen, N_2

Description of Action	Action
1. Determine how many electrons you have to work with; this is done by multiplying the number of valence electrons for that element by the number of those atoms. Label your total as "HAVE".	**1.** N_2 nitrogen's valence # of nitrogen **HAVE** 5 x 2 = **10** So, we have 10 electrons with which to work.
2. If the given molecule is an ion, add or subtract electrons from your total, depending on the charge of the molecule. If the molecule has a positive charge, subtract electrons from your total. If the molecule has a negative charge, add electrons to your total.	**2.** HAVE N_2 5 x 2 = 10 N_2 is not an ion so we do not have to add or subtract electrons to or from our total.
3. Write the symbols for the given formula with some space between them. If there are more than two atoms, put the one with the **lowest electronegativity** in the center and the others on each of the four sides. **Exception:** Hydrogen can NEVER be in the center. An element's electronegativity can be found on the front of your periodic table.	**3.** HAVE N_2 5 x 2 = 10 $\boxed{\mathrm{N \quad N}}$
4. Draw one dash between the symbols to represent a bond. Each dash drawn represents 2 electrons. Subtract 2 from your "HAVE" amount for each bond you draw.	**4.** HAVE N_2 5 x 2 = 10 $\boxed{\mathrm{N{-}N}}$ **- 2** **8**
5. Each element in a covalent compound wants to share 8 electrons. We have to check to see how many electrons each element needs to reach 8. Every bond that is drawn to a symbol should be counted as 2 electrons for that element. Add these numbers together and label this number as "NEED".	**5.** HAVE N_2 5 x 2 = 10 $\boxed{\mathrm{N{-}N}}$ **NEED** **- 2** **⑧** 6 6 = **⑫** Each nitrogen atom has 2 electrons because of the bond between them. That means each nitrogen still needs 6 more electrons. 6 + 6 = 12, so we NEED 12 electrons.
6. Compare the amount of electrons you HAVE with the number of electrons you NEED. If the values match, draw the correct amount of dots on each symbol. Dots must be drawn in pairs. If the number of electrons you HAVE and the number of electrons you NEED do not match, draw another bond between the symbols and subtract 2 more electrons from your total.	**6.** HAVE N_2 5 x 2 = 10 $\boxed{\mathrm{N{-}N}}$ **NEED** **- 2** **8** 6 6 = **12** **- 2** $\boxed{\mathrm{N{=}N}}$ **6** **-2** 4 4 = **8** **④** $\boxed{\mathrm{N{\equiv}N}}$ **2** **2** = **④**
7. Once the numbers match, draw in your dots. The number of dots to be drawn on each symbol is indicated by the number below it once the number of electrons you NEED and the number of electrons you HAVE match. All dots must be written in pairs.	**7.** Because we have a 2 under each nitrogen, each nitrogen gets 2 dots. Write them as a pair. $\boxed{\mathrm{:N{\equiv}N:}}$
8. If the molecule is an ion and has an overall charge, put brackets around the structural formula and write the charge outside the brackets in the upper right corner. (See the next example to see what I mean.)	**8.** Nitrogen is not an ion. Thus, our structural formula does not need brackets. We are done.

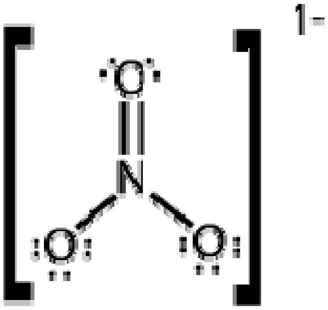
You Try: Write Lewis Structures for the following: HF, CH_4, NO^+, CO_2, CO_3^{2-}

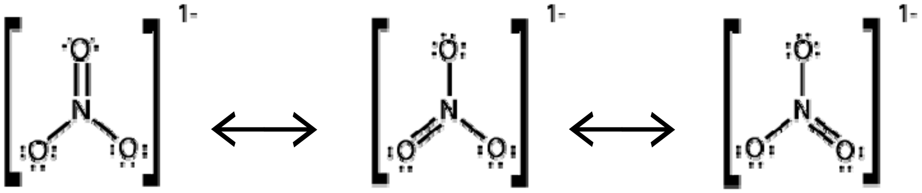
Exceptions to the Octet Rule

- The localized electron model is successful but there are some exceptions that must be noted.
- The second-row elements C,N,O and F should always be assumed to obey the octet rule.
- **The second row elements B and Be often have fewer than eight electrons around them in their compounds** (know BF_3). These electron deficient compounds are very reactive.
- **Second row elements never exceed the octet rule, since their valence orbitals (2s and 2p) can accommodate only 8 electrons.**
- **Third row and heavier elements often satisfy the octet rule but can exceed the octet rule by using their empty valence d orbitals.** (PCl_5, SF_6).
- When writing the Lewis structure for a molecule, satisfy the octet rule for all atoms first. If electrons remain after the octet rule has been satisfied, then place them on the elements having available d orbitals.
- When it is necessary to exceed the octet rule for one of several third row (or higher) elements, assume that the extra electrons should be placed on the central atom.
- You try: Write the structural formula for the following: SF_6, ClF_3, $RnCl_2$ and $BeCl_2$

Resonance

- Sometimes more than one valid Lewis structure is possible for a given molecule.
- Observe the Lewis structure for nitrate to the right. It shows one double bond and two single bonds. But, experiments clearly show that only one type of N – O bond occurs with length and strength between those expected for a single and double bond.
- The structure to the right is a valid Lewis Structure but it does not correctly represent the bonding in NO_3^-.
- Resonance occurs when more than one valid Lewis structure can be written for a particular molecule.
- The resulting electron structure of the molecule is the average of the resonance structures. The three resonance structures for nitrate are shown below.





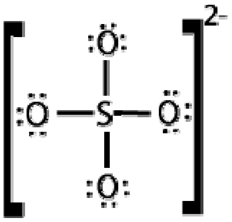
- Resonance shows that electrons are not localized to one atom but instead travel throughout the molecule.

Odd-Electron Molecules

- Few molecular compounds contain odd numbers of electrons.
- Two exceptions are nitric oxide (NO) and nitrogen dioxide (NO_2).
- The localized electron model is based on pairs of electrons; it does not handle odd-electron cases in a natural way. A more sophisticated model is needed.

Formal Charge

- The formal charge of an atom in a molecule is the difference between the number of valence electrons on the free atom and the number of valence electrons assigned to the atom in the molecule.
- Formal Charge = (valence e^- on free atom) – (valence e^- assigned to atom in molecule)
- Valence e^- assigned = (number of lone pair e^-) + ½ (number of shared e^-)
- For Example: Calculate the formal charge for each element in the structural formula shown to the right.
  - Oxygen:
    - Valence e^- assigned = (6) + ½ (2) = 7
    - Formal Charge = (6) – (7) = -1
  - Sulfur:
    - Valence e^- assigned = (0) + ½ (8) = 4
    - Formal Charge = (6) – (4) = +2



- Atoms in molecules try to achieve formal charges as close to zero as possible.
- Any negative formal charges are expected to reside on the most electronegative atoms.
- You try: Above it was mentioned that NO and NO_2 have an odd number of electrons. Use your knowledge of formal charge to determine the correct Lewis structures for each of the molecules. You must show a formal charge calculation for each atom in each possible Lewis structure.

Molecular Structure: The VSEPR Model

- The term trigonal can be used instead of triangular in shapes such as triangular planar.

Central Atoms with Three Electron Pairs				
# of lone pair	General Formula	shape	Example	Lewis Structure
0	AX_3E_0	triangular planar	BCl_3	
1	AX_2E_1	bent (angular)	NO_2^-	

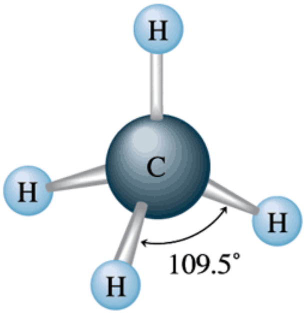
Central Atoms with Four Electron Pairs				
# of lone pair	General Formula	shape	Example	Lewis Structure
0	AX_4E_0	tetrahedral	CCl_4	
1	AX_3E_1	triangular pyramidal	NH_3	
2	AX_2E_2	bent (angular)	H_2O	

Central Atoms with Five Electron Pairs				
# of lone pair	General Formula	shape	Example	Lewis Structure
0	AX_5E_0	triangular bipyramidal	PF_5	
1	AX_4E_1	seesaw	SF_4	
2	AX_3E_2	T-shaped	ClF_3	
3	AX_2E_3	linear	XeF_2	

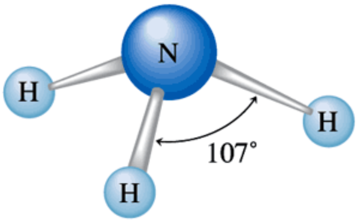
Central Atoms with Six Electron Pairs				
# of lone pair	General Formula	shape	Example	Lewis Structure
0	AX_6E_0	octahedral	SF_6	
1	AX_5E_1	square pyramidal	BrF_5	
2	AX_4E_2	square planar	XeF_4	

- Note in the chart below that even though CH_4, NH_3 and H_2O are composed of central atoms with 4 electron pair, the bond angles between atoms varies in the different molecules.

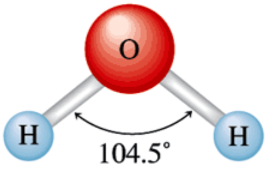
Methane



Ammonia



Water



- **These observations provide evidence that lone pairs of electrons need more room than bonding pairs and tend to compress the angles between the bonding pairs.**

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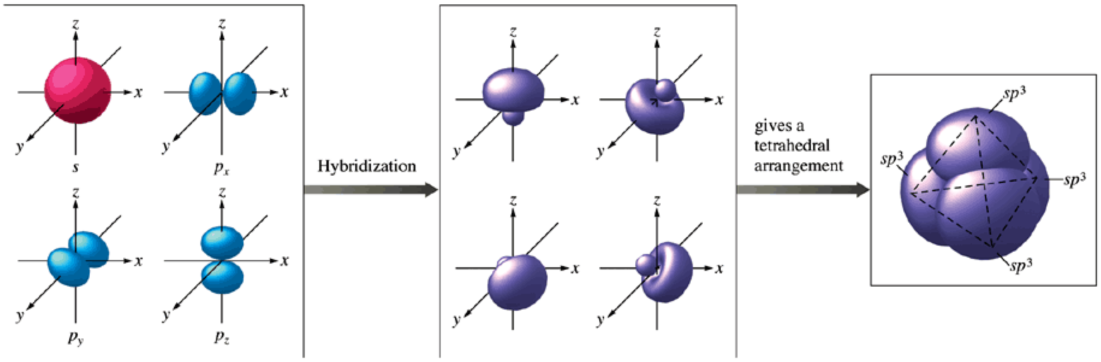
Chapter 9 Outline – Covalent Bonding: Orbitals

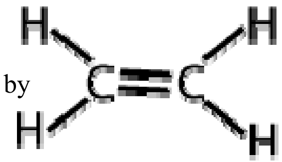
Hybridization and the Localized Electron Model

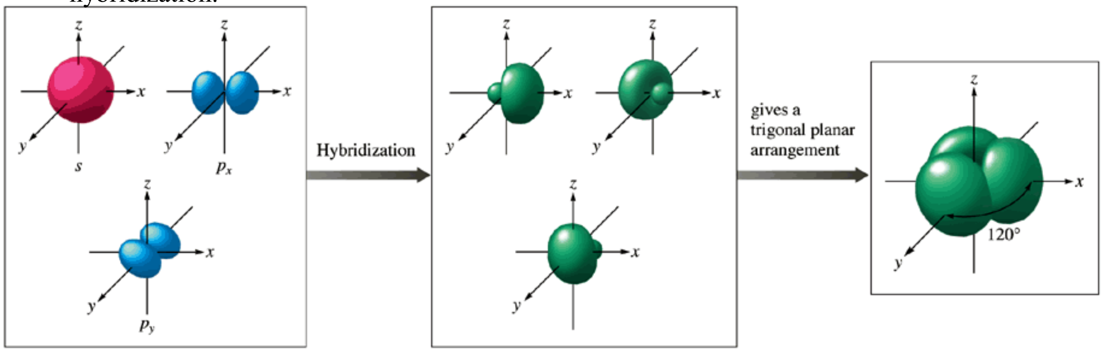
- **sp^3 Hybridization**
  - In general we assume that bonding only involves valence electrons.
  - Consider methane (CH_4).
    - The valence electrons of hydrogen use 1s orbitals.
    - The 2p and 2s atomic orbitals of carbon will lead to two different type of C—H bonds: those from the overlap of the carbon 2p orbital with the 1s orbital of hydrogen and those from the overlap of a carbon 2s orbital with the 1s orbital of hydrogen.
    - This is not the case. The methane molecule is tetrahedral with bond angles of 109.5° So either the localized electron model is wrong or carbon adopts a set of orbitals other than its "native" 2s and 2p orbitals to bond to the hydrogen atoms in forming the methane molecule.
    - The 2s and 2p orbitals present on an isolated carbon atom may not be the best set of orbitals for bonding. It makes sense to assume that the carbon atom has four equivalent atomic orbitals arranged tetrahedrally.
    - Such a set of orbitals can be obtained by combining the carbon 2s and 2p orbitals as shown below.

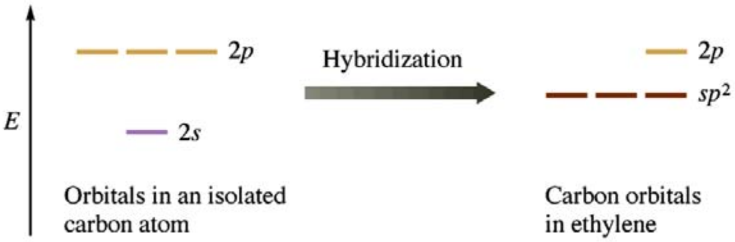


    - This mixing of the native atomic orbitals to form special orbitals for bonding is called hybridization.
    - The four new orbitals are called sp^3 orbitals because they are formed from one 2s orbital and three 2p orbitals(s^1p^3).
    - We say that the carbon atom undergoes sp^3 hybridization or is sp^3 hybridized.
    - Whenever a set of equivalent tetrahedral atomic orbitals is required by an atom, this model assumes that the atom adopts a set of sp^3 orbitals; the atom becomes sp^3 hybridized.
    - What the atoms in a molecule were like before the molecule was formed is not as important as how the electrons are best arranged in the molecule. This model assumes that the individual atoms respond as needed to achieve the minimum energy for the molecule.

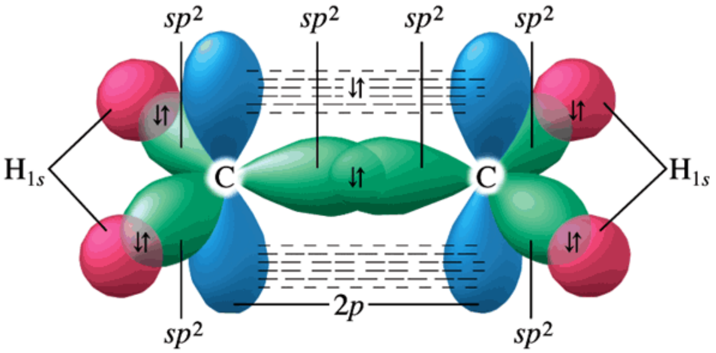
- **sp^2 hybridization**
  - Ethylene (C_2H_4) is an important material in the manufacture of plastics.
  - A double bond acts as one effective pair, so the ethylene is surrounded by three effective pair.
  - This requires a trigonal planar arrangement with bond angles of 120°.
  - Since one 2s and two 2p orbitals are used to form these hybrid orbitals, this is called sp^2 hybridization.





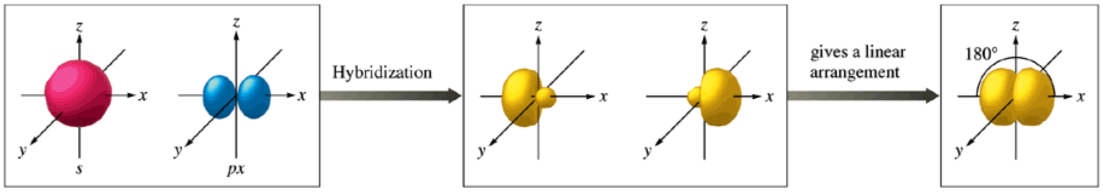


  - In forming the sp^2 orbitals, one 2p orbital on carbon has not been used. This remaining p orbital (p_z) is oriented perpendicular to the plane of the sp^2 orbitals.
  - The three sp^2 orbitals on each carbon can be used to share electrons, as shown below.

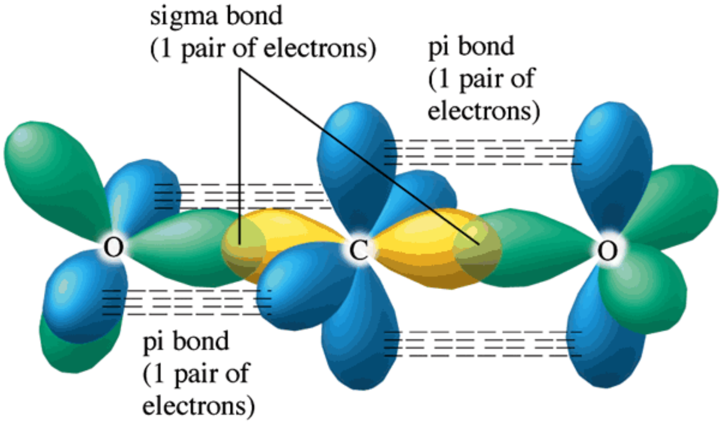


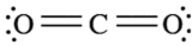
  - In each of these bonds, the electron pair is shared in an area centered on a line running between the atoms. This type of covalent bond is called a sigma(σ) bond. In the ethylene molecule, the σ bonds are formed using sp^2 orbitals on each carbon atom and the 1s orbital on each hydrogen atom.
  - The second bond must therefore result from sharing an electron pair in the space above and below the σ bond.
  - The parallel p orbitals can share an electron pair, to form a pi(π) bond.
  - σ bonds are formed from orbitals whose lobes point toward each other, but π bonds result from parallel orbitals. A double bond always consists of one σ bond where the electrons are located between the atoms and one π bond which occupies the space above and below the σ bond.

- **sp hybridization**
  - In the CO_2 molecule the carbon has two effective pair that will be arranged at an angle of 180°.
  - To obtain two hybrid orbitals arranged at 180° requires sp hybridization, involving one s orbital and one p orbital.



  - Two effective pairs around an atom will always require sp hybridization of that atom.
  - In CO_2 the sp orbitals on carbon form σ bonds with the sp^2 orbitals on the two oxygen atoms. The remaining sp^2 orbitals on the oxygen atoms hold lone pair. The π bonds between the carbon atom and each oxygen atom are formed by the overlap of parallel 2p orbitals.



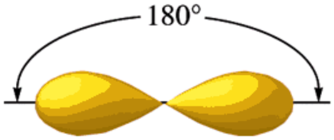
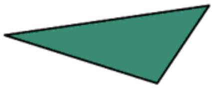
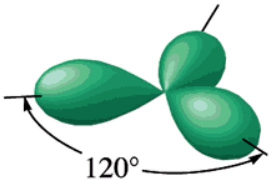
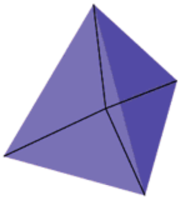
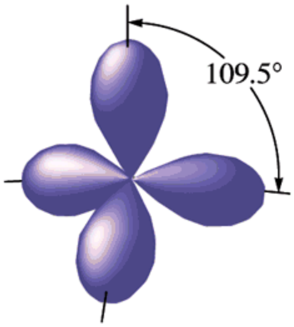
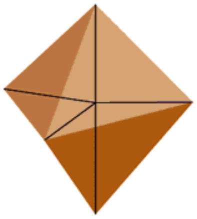
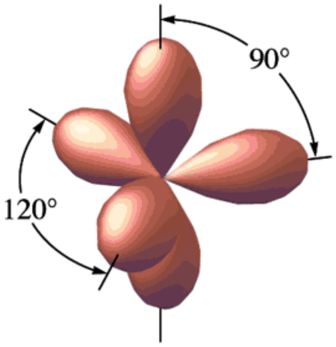
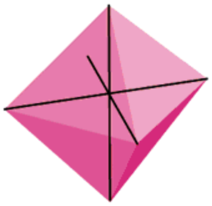


- **dsp^3 hybridization**
  - Since 5 pairs of electrons are needed for atoms such as phosphorus in PCl_5, a dsp^3 hybridization is necessary.
  - The dsp^3 hybridized phosphorus atom in PCl_5 molecule uses its 5 dsp^3 orbitals to share electrons with five chlorine atoms.
  - A set of five effective pairs around a given atom always requires a trigonal bipyramidal arrangement which in turn requires dsp^3 hybridization of that atom.

- **d^2sp^3 hybridization**
  - Since 6 pairs of electrons are needed for atoms such as sulfur in SF_6, a d^2sp^3 hybridization is necessary.
  - A set of six effective pairs around a given atom always requires an octahedral arrangement which in turn requires d^2sp^3 hybridization of that atom.

Paramagnetism

- Most materials have no magnetism until they are placed in a magnetic field. However, in the presence of such a field, magnetism of two types can be induced.
- Paramagnetism causes a substance to be attracted to the inducing magnetic field.
- Diamagnetism causes the substance to be repelled from inducing magnetic field.
- Studies have shown that paramagnetism is associated with unpaired electrons and diamagnetism is associated with paired electrons.
- The O_2 molecule is known to be paramagnetic.

Number of Effective Pairs	Arrangement of Pairs		Hybridization Required	
2		Linear	*sp*	
3		Trigonal planar	*sp*^{2}	
4		Tetrahedral	*sp*^{3}	
5		Trigonal bipyramidal	*dsp*^{3}	
6		Octahedral	*d*^{2}*sp*^{3}