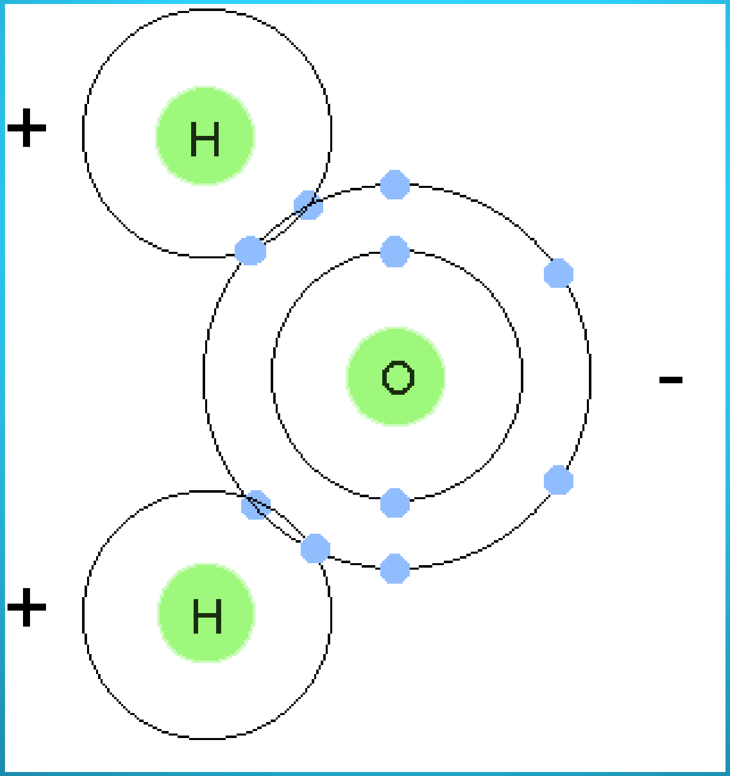
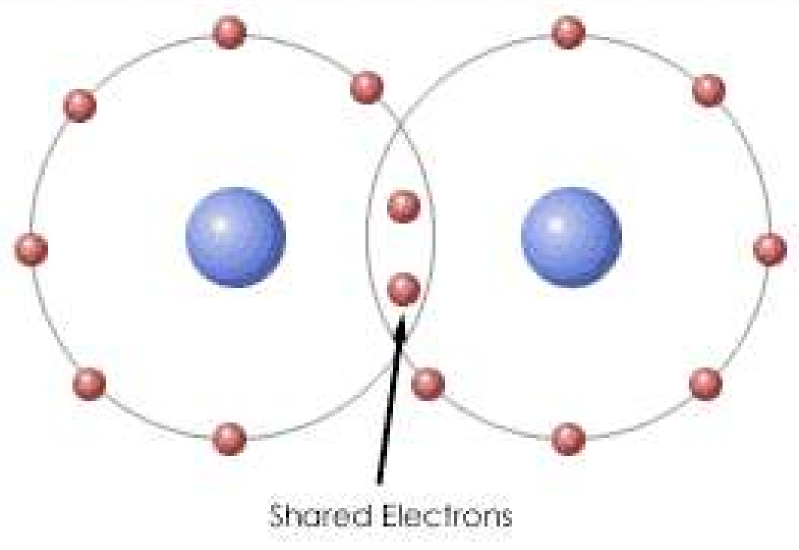
Chapter 8 – Covalent Bonding



Jennie L. Borders

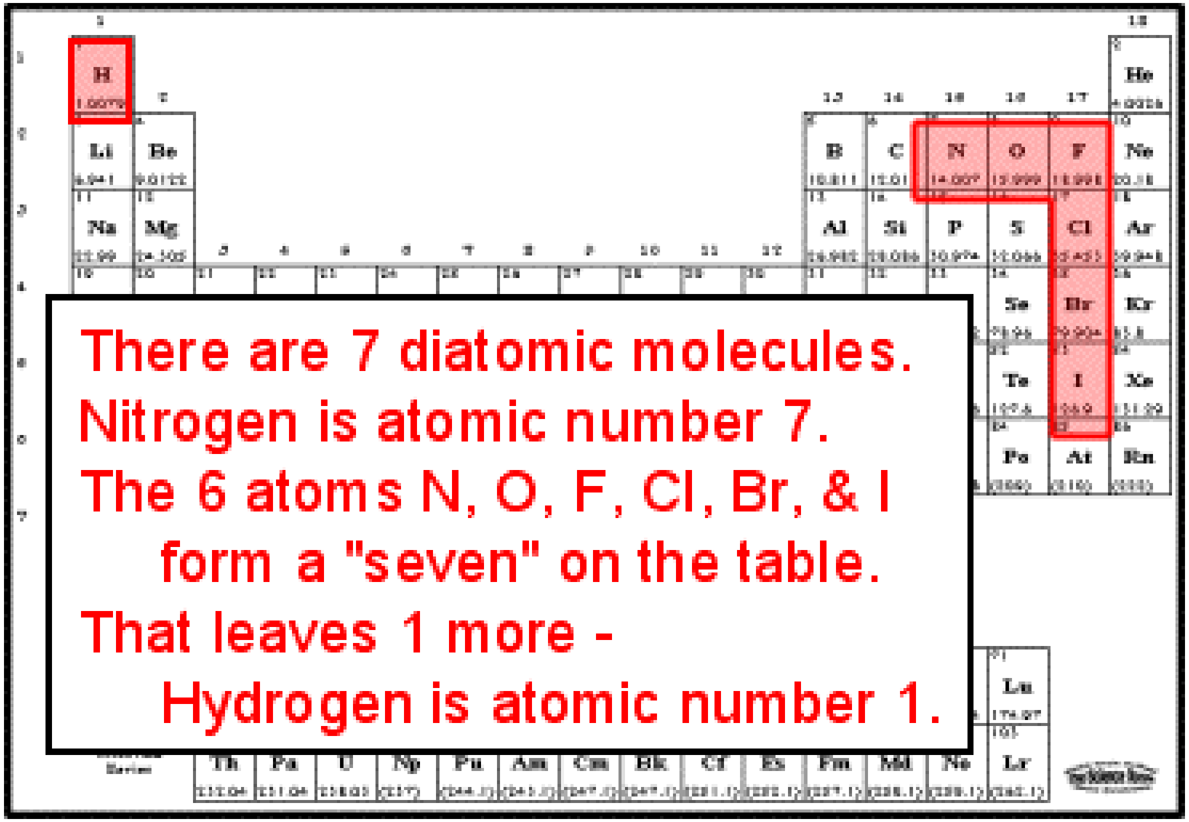
Section 8.1 – Molecular Compounds

- A <u>covalent bond</u> is formed between atoms held together by <u>sharing</u> electrons.
- A <u>molecule</u> is a group of atoms joined by <u>covalent bonds</u>.
- A <u>diatomic molecule</u> is 2 atoms bonded together.



Diatomic Elements

- There are 7 naturally existing diatomic <u>elements</u>. They are <u>N_2, O_2, F_2, Cl_2, Br_2, I_2, and H_2</u>.

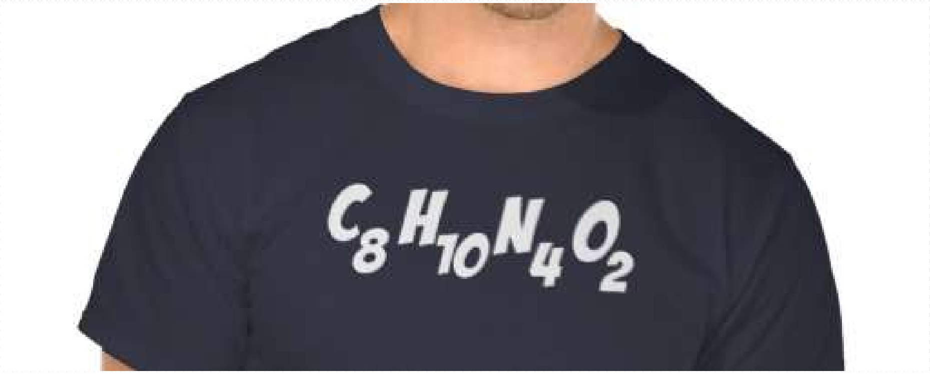


Properties of Molecular Compounds

Molecular Compounds	Ionic Compounds
Low melting points	high melting points
Tend to be gases or liquids	crystalline solids
Made of nonmetals	metal and nonmetal
Made of covalent bonds	ionic bonds
Poor conductors	conductor when molten or aqueous

Molecular Formulas

- A molecular formula is the chemical formula of a molecular compound.
- A molecular formula shows how many atoms of each element a molecule contains.
- A molecular formula shows the actual number of atoms while a formula unit shows the lowest whole-number ratio of ions.



Section 8.1 Assessment

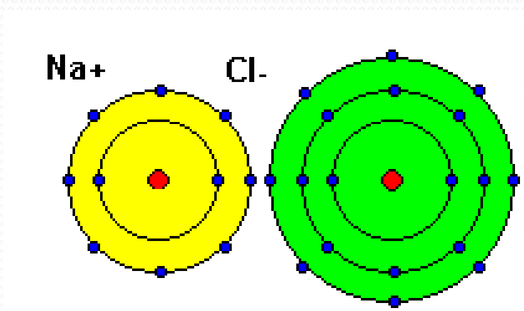
1. How are the melting points and boiling points of molecular compounds usually different from those of ionic compounds?
2. What information does a molecular formula provide?
3. What are the only elements that exist in nature as uncombined atoms? What term is used to describe such elements?
4. Describe how the molecule whose formula is NO is different from the molecule whose formula is N_2O.

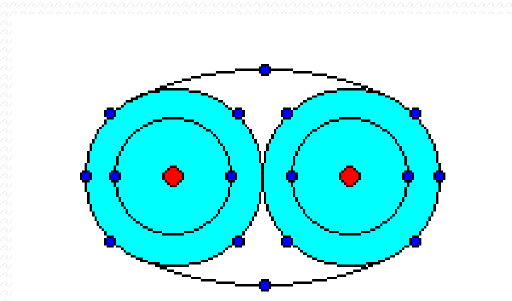
Section 8.1 Assessment

5. Give an example of a diatomic molecule found in Earth's atmosphere.

Section 8.2 – The Nature of Covalent Bonding

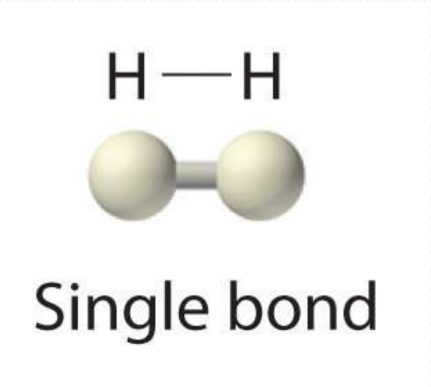
- In ionic bonding, atoms transfer electrons to achieve noble gas configuration.
- In covalent bonding, atoms share electrons to achieve noble gas configuration.
- Most atoms share electrons until they have a total of 8 valence electrons (octet rule). However, hydrogen only needs 2 electrons to be stable.

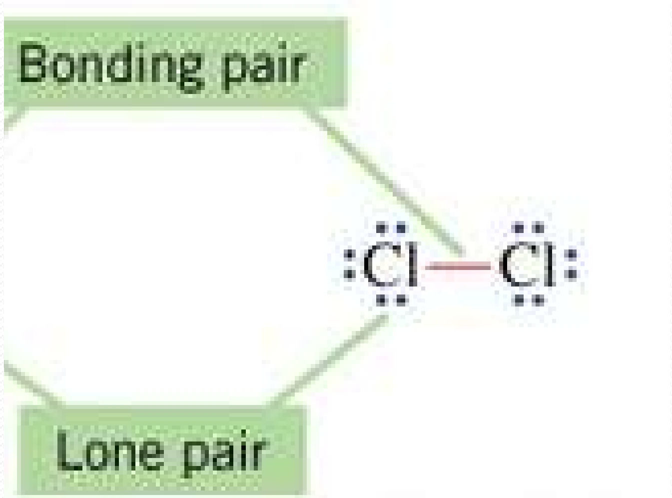




Covalent Bonds

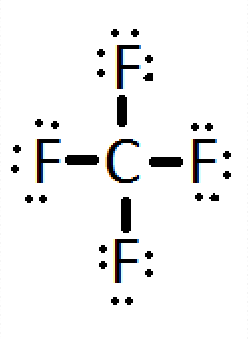
- A single covalent bond is created when 2 atoms share 1 pair of electrons.
- A dash represents a bond which is made of 2 electrons.
- An unshared pair of electrons, or a lone pair, is represented as dots.





Rules for Writing Lewis Dot Structures

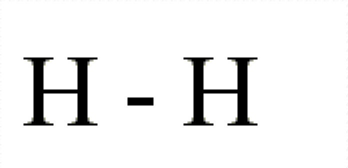
1. Add up the total number of valence electrons.
2. Bond the atoms with single bonds. (Single atoms go in the middle.)
3. Add electrons until each atom has a full octet and each hydrogen has a duet (2 electrons).
4. Add up total valence electrons in Lewis dot structure and compare to the total from step 1.



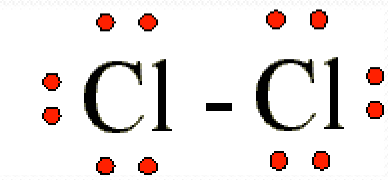
Sample Problems

- Write the Lewis dot structure for the following:

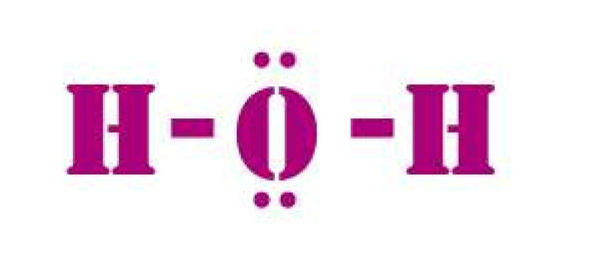
- H_2



- Cl_2

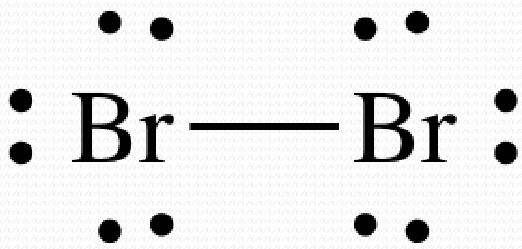


- H_2O

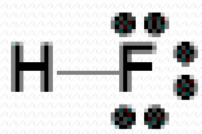


Practice Problems

- Write the Lewis dot structure for the following:
- Br_2



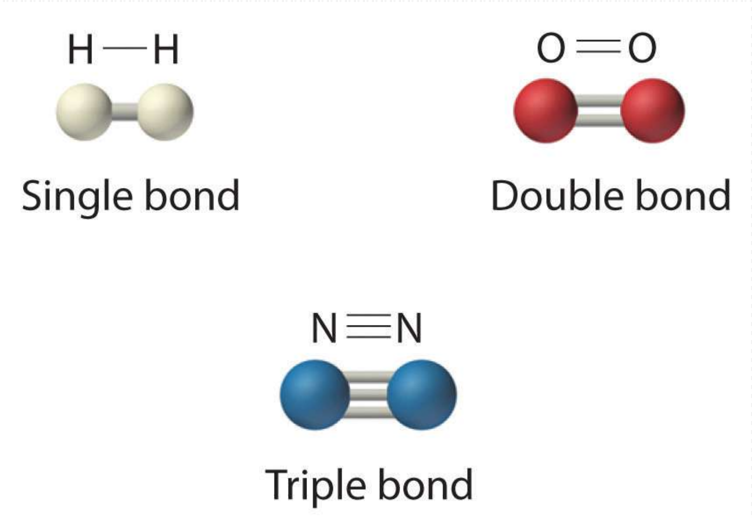
- HF



- NH_3

Multiple Bonds

- A <u>double bond</u> occurs when 2 atoms share <u>two</u> pairs of electrons. It is represented by <u>2</u> dashes which equal <u>4</u> electrons.
- A <u>triple bond</u> occurs when 2 atoms share <u>three</u> pairs of electrons. It is represented by <u>3</u> dashes which equal <u>6</u> electrons.

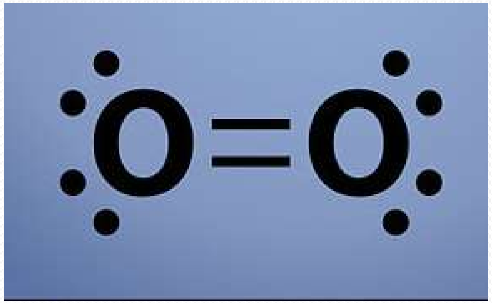


Lewis Dot Structures with Multiple Bonds

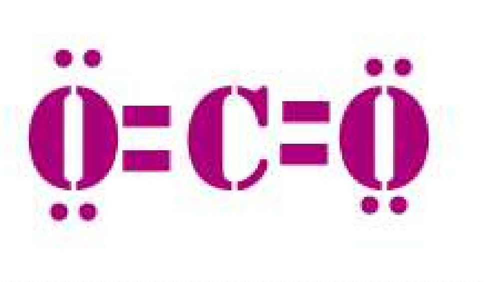
- When writing the Lewis dot structures, following the 4 steps we learned.
- When you add up the total number of electrons in your Lewis dot structure, sometimes it will not equal the total from step 1.
- For every extra electron pair you have, you need to add 1 more bond in your structure.

Sample Problem

- Write the Lewis dot structure for the following:
- O_2

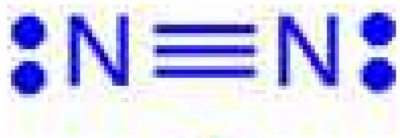


- CO_2

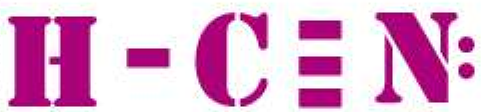


Practice Problems

- Write the Lewis dot structures for the following:
- N_2

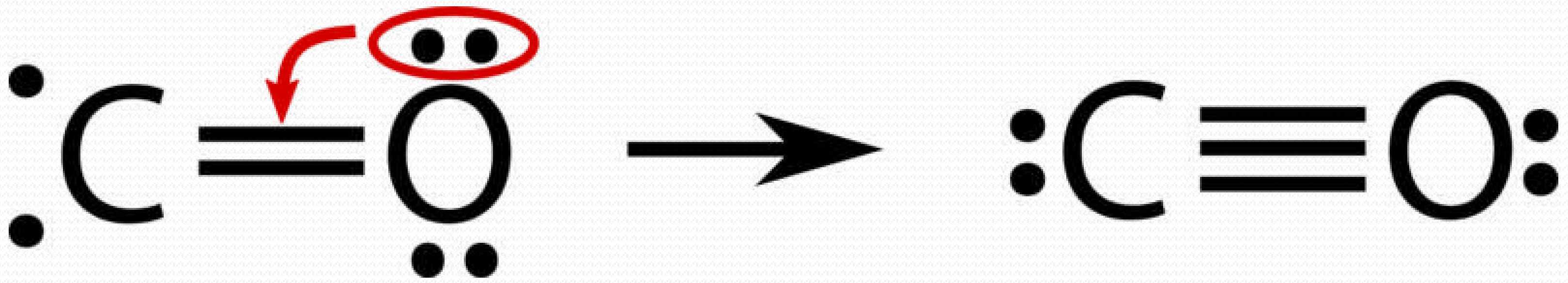


- HCN



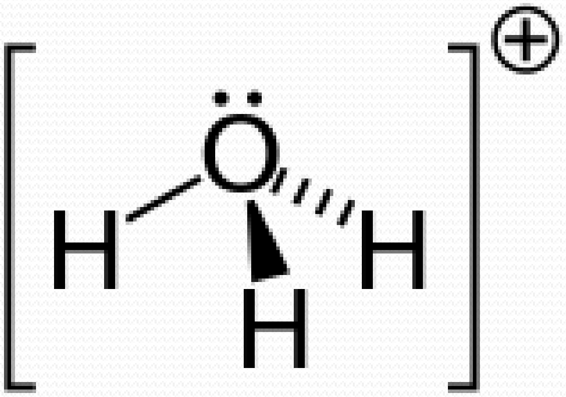
Coordinate Covalent Bond

- A coordinate covalent bond is a bond formed when one atom donates both of the shared electrons.
- In a regular bond, each atom donates 1 electron to form the bond.



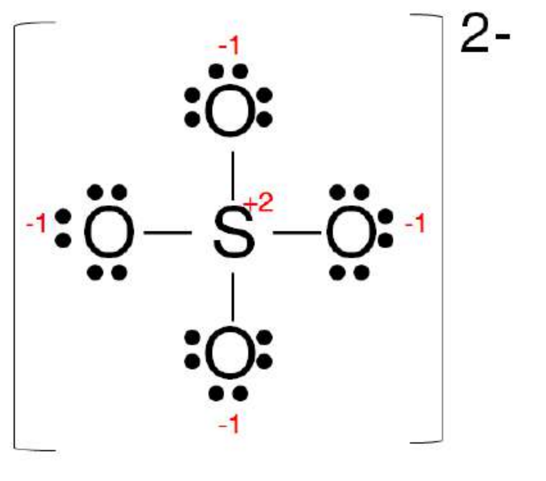
Polyatomic Ions

- When writing the Lewis dot structure for a polyatomic ion, you have to take into account the charge when you add up the number of valence electrons in step 1.
- After you draw the Lewis dot structure, you have to put the whole structure in brackets and write the charge.

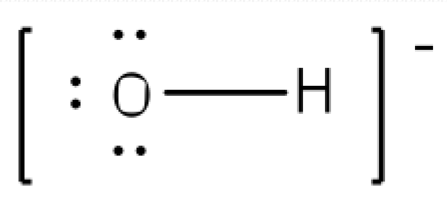


Sample Problems

- Write the Lewis dot structure for the following:
- SO_4^{-2}

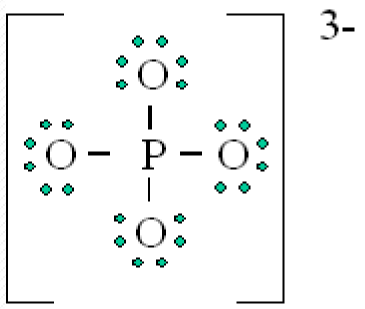


- OH^-

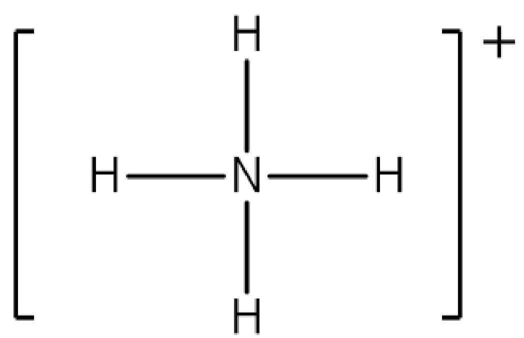


Practice Problems

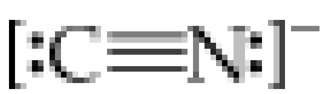
- Write the Lewis dot structure for the following:
- PO_4^{-3}



- NH_4^+



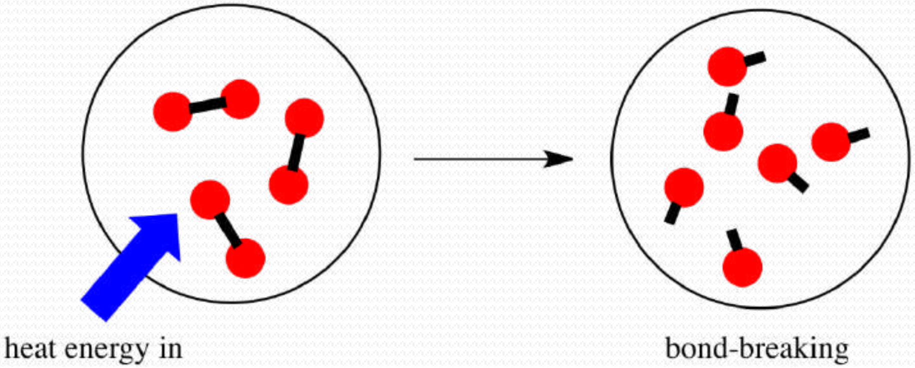
- CN^-



Bond Dissociation Energy

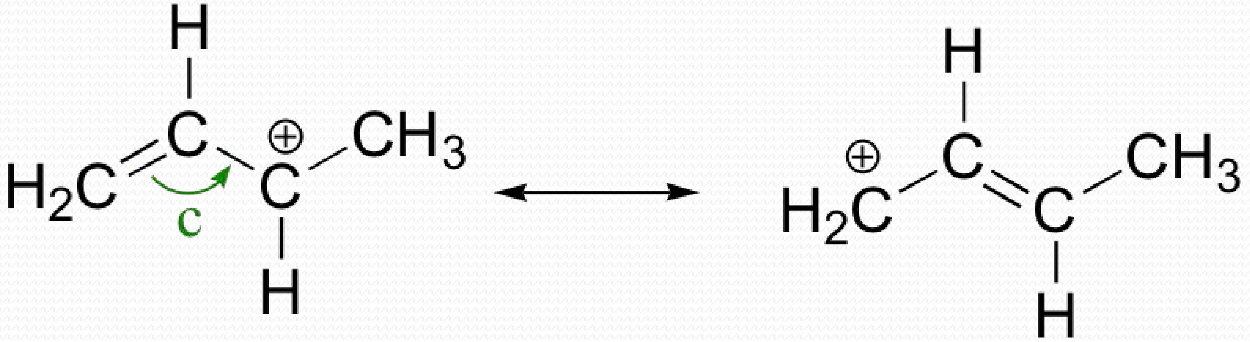
- The bond dissociation energy is the energy needed to break a bond.
- As the number of bonds increases, the bond dissociation energy increases.

Single < Double < Triple



Resonance (Honors)

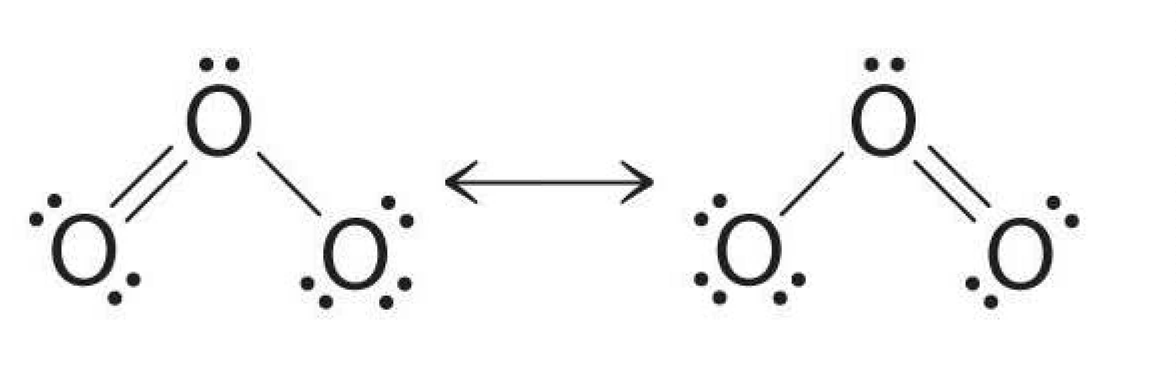
- A <u>resonance structure</u> is a structure that occurs when it is possible to draw <u>two</u> or more valid Lewis dot structures for a substance.
- Ex: O_3
- When <u>resonance</u> occurs, you should separate the possible structures with a <u>double sided arrow</u>.
- The <u>actual</u> structure is a <u>hybrid</u> of the resonance structures.



Sample Problem (Honors)

- Write the Lewis dot structures for the following:

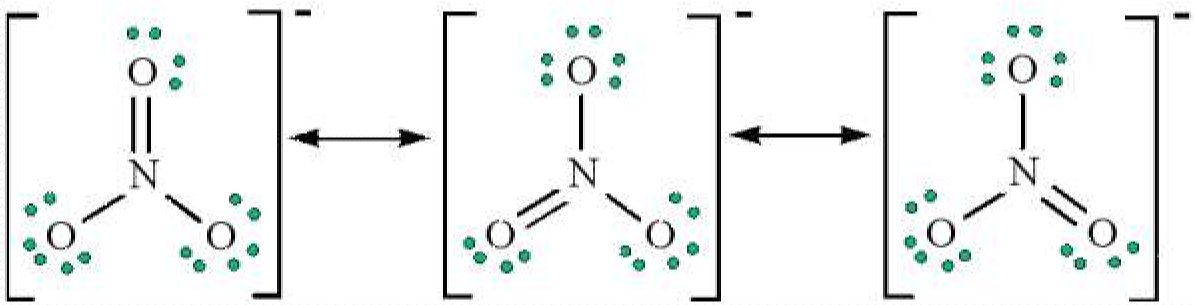
- O_3



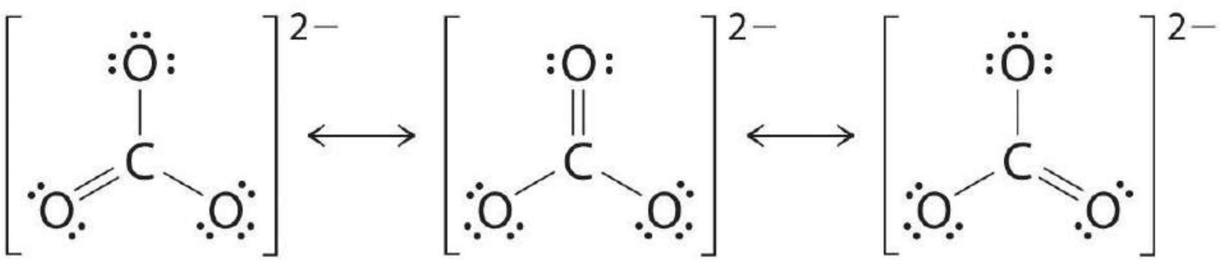
Practice Problems (Honors)

- Write the Lewis dot structures for the following:

- NO_3^-



- CO_3^{-2}

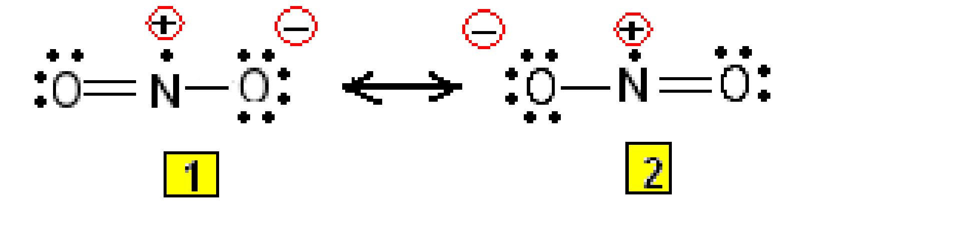


Exceptions to the Octet Rule

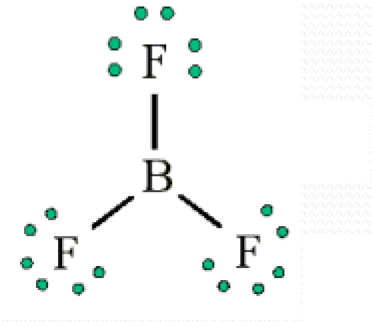
- Some molecules are exceptions to the octet rule.
- ODD NUMBER - Some atoms have an odd number of electrons. This usually occurs with nitrogen. The odd electron goes to the central atom.
- LESS THAN 8 – Some atoms have less than 8 electrons. This usually happens with elements 1-5.
- MORE THAN 8 – Some atoms have more than 8 electrons. This usually happens with S, P, the halogens, and some noble gases.

Sample Problem

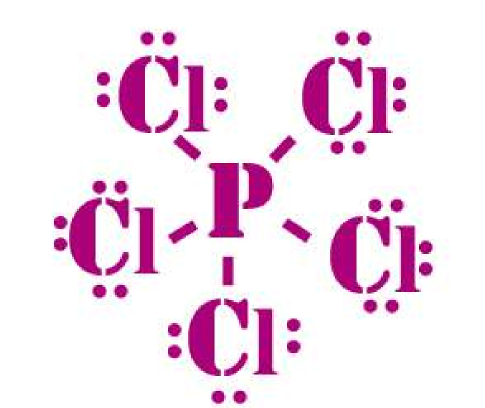
- Write the Lewis dot structure for the following:
- NO_2



- BF_3

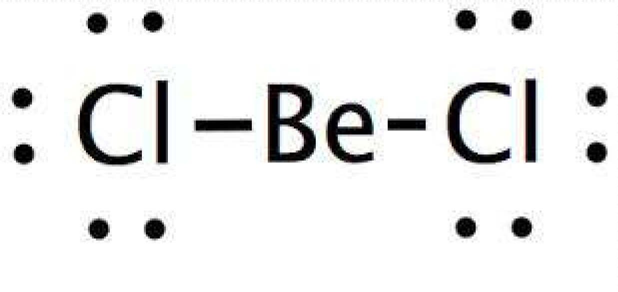


- PCl_5

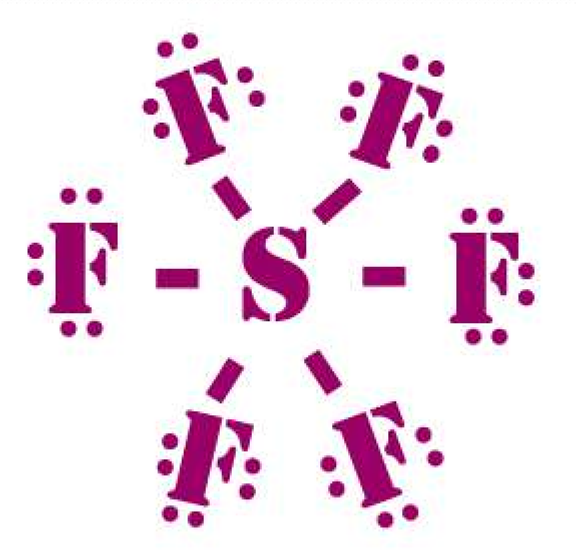


Practice Problem

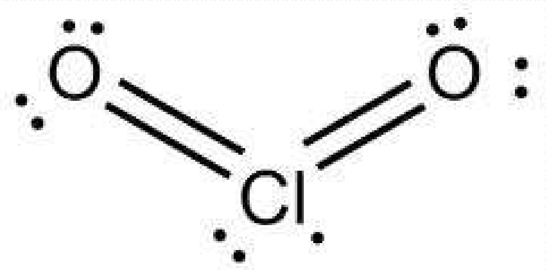
- Write the Lewis dot structure for the following:
- $BeCl_2$



- SF_6



- ClO_2



Section 8.2 Assessment

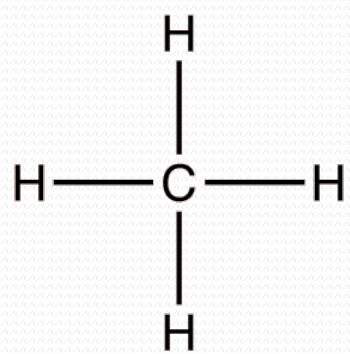
1. What electron configurations do atoms usually achieve by sharing electrons to form covalent bonds?
2. When are two atoms likely to form a double bond between them? A triple bond?
3. How is a coordinate covalent bond different from other covalent bonds?
4. How is the strength of a covalent bond related to its bond dissociation energy?

Section 8.2 Assessment

5. What kinds of information does a structural formula reveal about the compound it represents?
6. Draw the electron dot structures for the following molecules.
   a. H_2S
   b. PH_3
   c. ClF

Section 8.3 – Bonding Theories (Honors)

- The <u>VSEPR</u> (valence shell electron pair repulsion) theory states that the <u>repulsion</u> between electron pairs causes molecular <u>shapes</u> to adjust so that the valence electron pairs stay as <u>far</u> apart as possible.
- <u>Lone pair electrons</u> alter the shape more than <u>bonding</u> electrons due to the fact that they spread out more.



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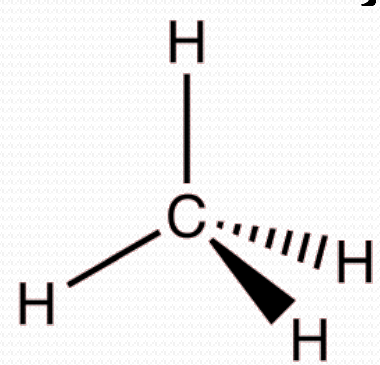
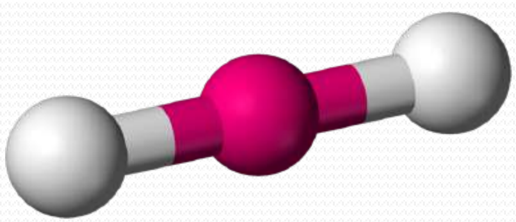


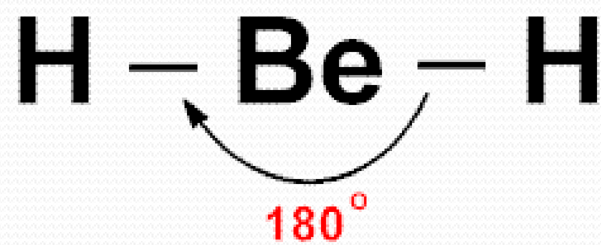
2

VSEPR Model

- Possible shapes for AB_2:

Bonding	Nonbonding	Shape
2	0	linear

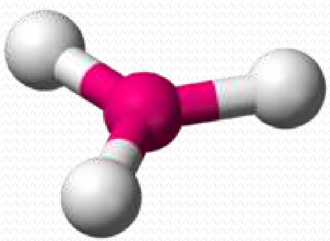


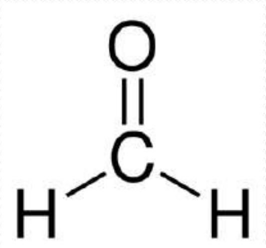


VSEPR Model

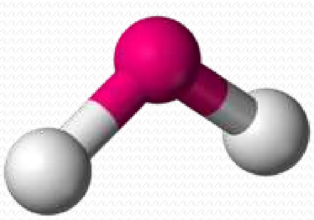
- Possible shapes for AB_3:

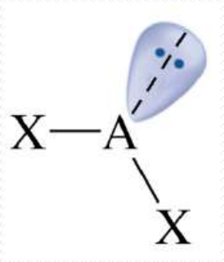
Bonding	Nonbonding	Shape
3	0	trigonal planar





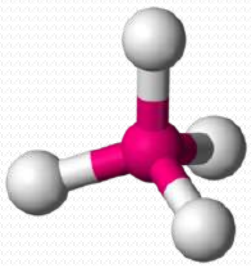
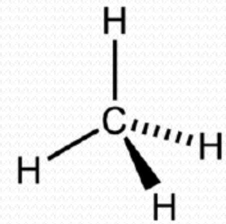
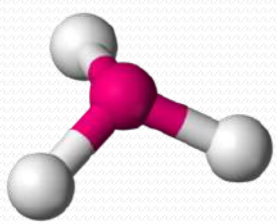
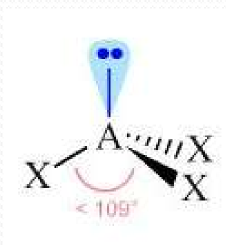
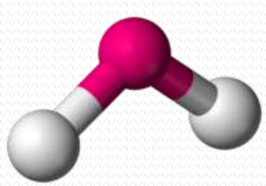
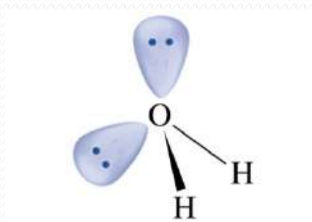
Bonding	Nonbonding	Shape
2	1	bent





VSEPR Model

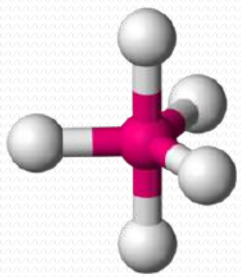
- Possible shapes for AB_4:

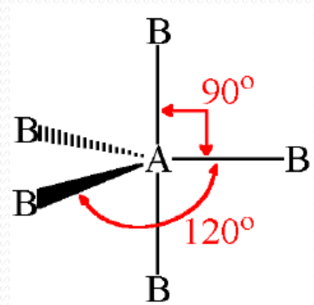
Bonding	Nonbonding	Shape
4	0	tetrahedral
		
3	1	trigonal pyramidal
		
2	2	bent
		

VSEPR Model

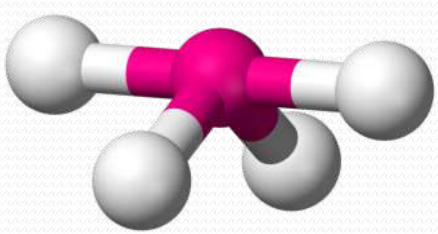
- Possible shapes for AB_5:

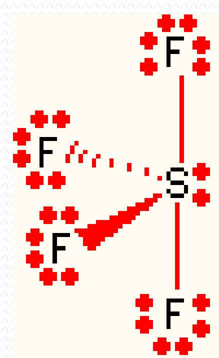
Bonding	Nonbonding	Shape
5	0	trigonal bipyramidal





Bonding	Nonbonding	Shape
4	1	seesaw

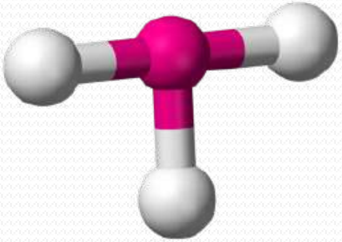


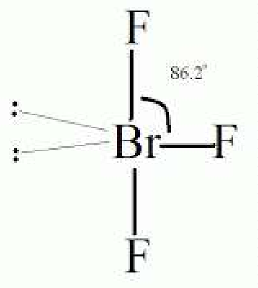


VSEPR Model

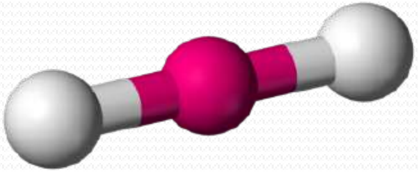
- Possible shapes for AB_5:

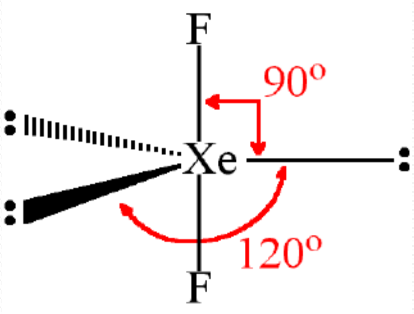
Bonding	Nonbonding	Shape
3	2	T-shaped





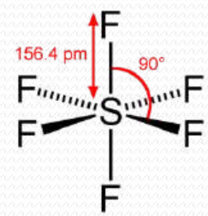
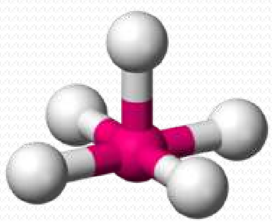
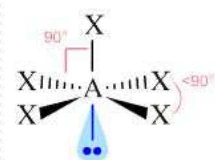
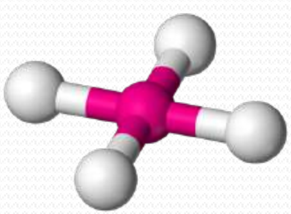
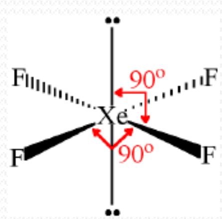
2	3	linear





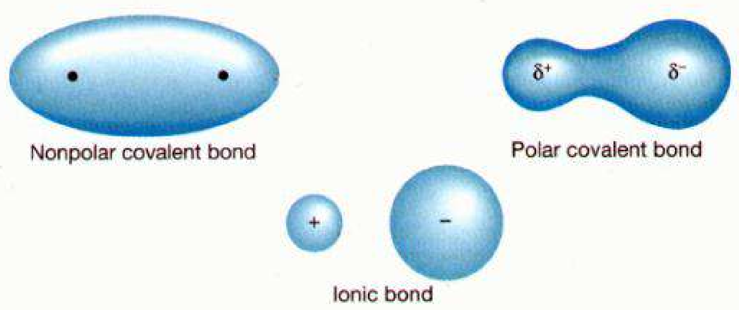
VSEPR Model

- Possible shapes for AB_6:

Bonding	Nonbonding	Shape
6	0	octahedral
		
5	1	square pyramidal
		
4	2	square planar
		

Section 8.4 – Polar Bonds and Molecules

- Covalent bonds involve sharing electrons between atoms.
- When the atoms in the bond pull equally, the bonding electrons are shared equally, and the bond is nonpolar.
- When the atoms in the bond pull unequally, the bonding electrons are pulled closer to one atom, and the bond is polar.



Polarity

- An atom's "strength" is measured by the electronegativity (the ability to attract electrons).
- The larger the electronegativity the more strongly an atom attracts electrons.
- The more electronegative elements gets a δ- (partial negative) charge and the less electronegative element gets a δ+ (partial positive) charge.

$$\overset{\delta+}{\mathrm{C}}\text{—}\overset{\delta-}{\mathrm{F}}$$

Sample Problem

- Determine the polarity of the following bonds:

- H – Cl

$\delta+$ $\delta-$

H - Cl

- F - P

$\delta-$ $\delta+$

F - P

Practice Problem

- Determine the polarity of the following bonds.

- Cl – C

δ- δ+

Cl - C

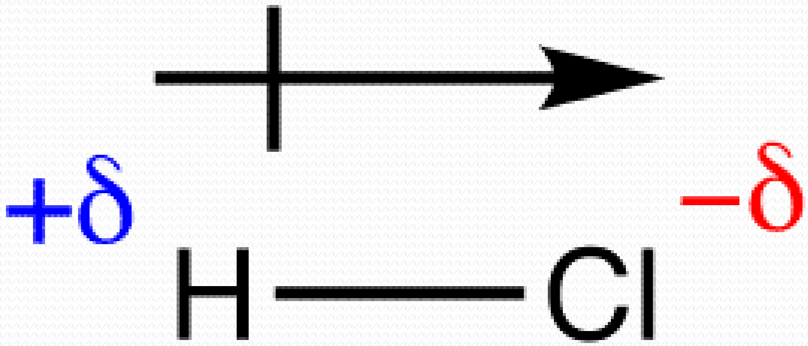
- O - S

δ- δ+

O - S

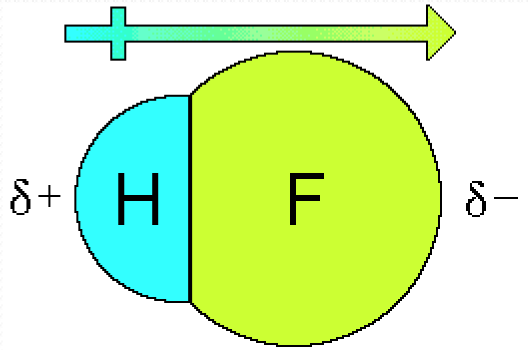
Dipole Moment

- A dipole moment occurs when a molecule is polar and has a partially positive side and a partially negative side.
- A dipole is represented by an arrow with a plus sign on one end. The arrow points toward the more electronegative element.

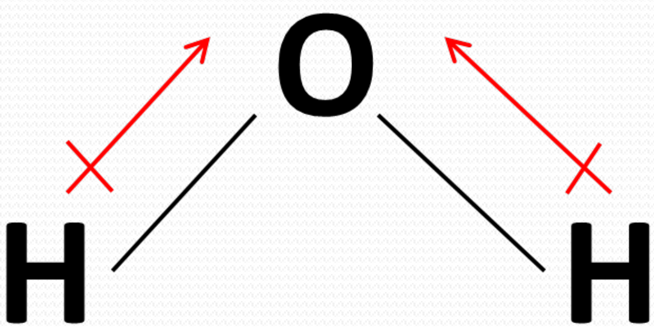


Sample Problem

- Draw the dipole moments for the following:
- H – F



- H_2O



Practice Problem

- Draw the dipole moment for the following:
- H – S

H - S

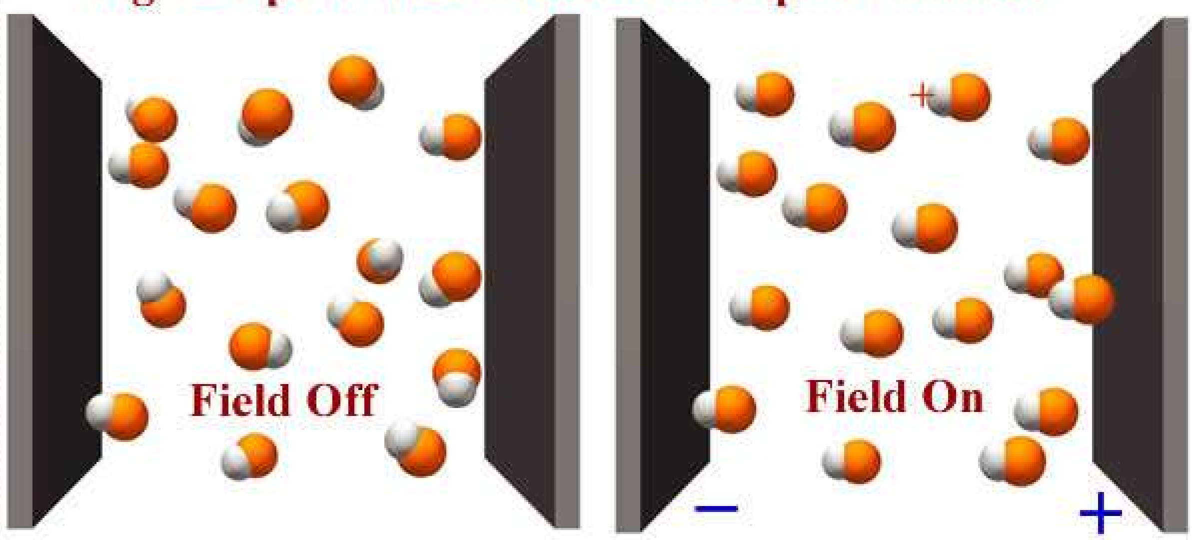
- N - H

N - H

Polar Molecules

- When polar molecules are placed between oppositely charged plates, the partially negative side is attracted to the positive plate and the partially positive side is attracted to the negative plate.

Fig 7.31 p373 Measurement of Dipole Moment



Determining Polarity

- Bond polarity is determined based on the difference in electronegativity between the to bonded atoms.

Bond Type	Electronegativity Difference
Nonpolar	<0.5
Polar	0.5 - 2
Ionic	>2

Sample Problem

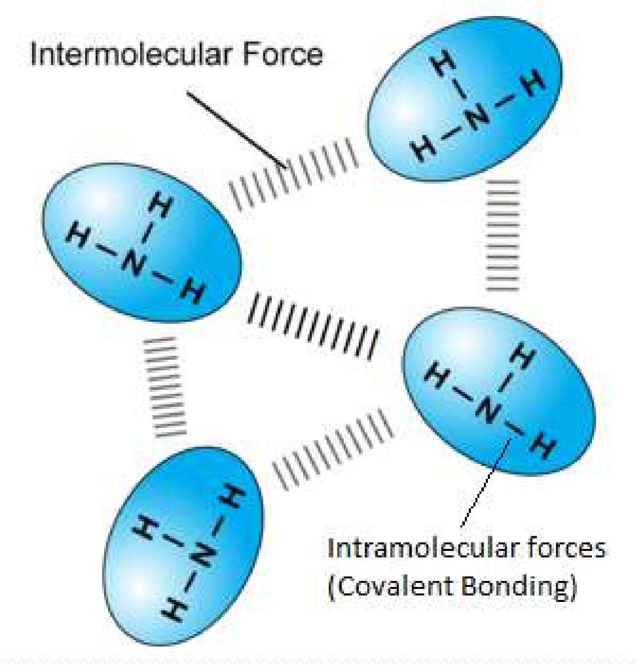
- Are the following bonds nonpolar, polar, or ionic?
- N – H polar
- F – F nonpolar
- Ca - Cl ionic

Practice Problem

- Are the following bonds polar, nonpolar, or ionic?
- H – Br polar
- C – O polar
- Li - O ionic

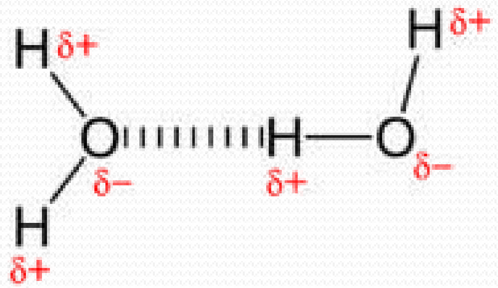
Attractions

- Intramolecular forces are the attractive forces within a single molecule. Ex: the bonds
- Intermolecular forces are the attractive forces that exist between multiple molecules. Ex: dipole attractions



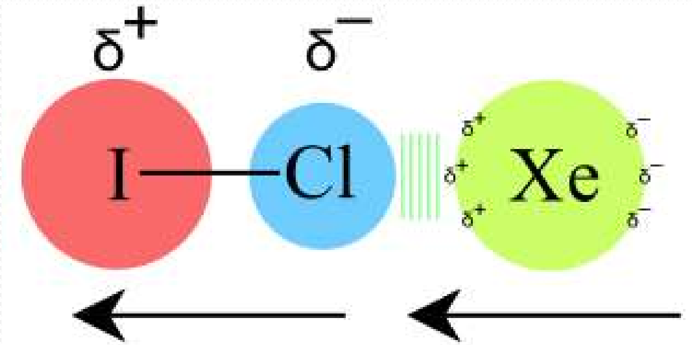
Intermolecular Forces

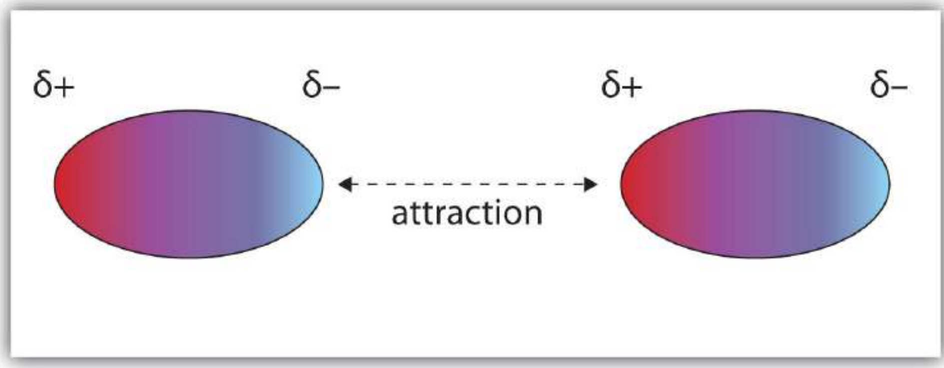
- London dispersion forces are the weakest intermolecular forces that exist between nonpolar molecules.
- Dipole-dipole attractions exist between polar molecules.
- Hydrogen bonding is a particularly strong dipole attraction that occurs between hydrogen and an extremely electronegative element (N, O, or F).

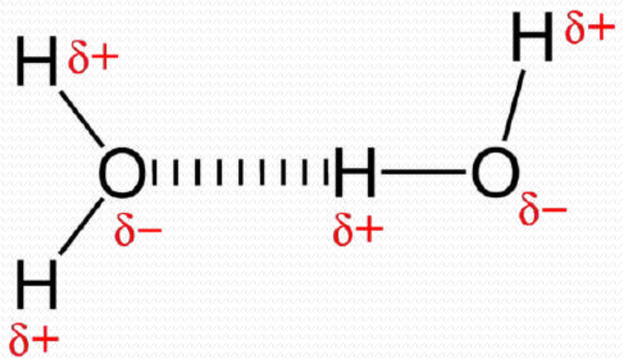


Strength of Intermolecular Forces

London Dispersion < Dipole-Dipole < Hydrogen Bonding







Sample Problem

- Determine the intermolecular forces that exist in the following molecules.

- SO_2 Dipole-dipole

- CH_4 London dispersion forces

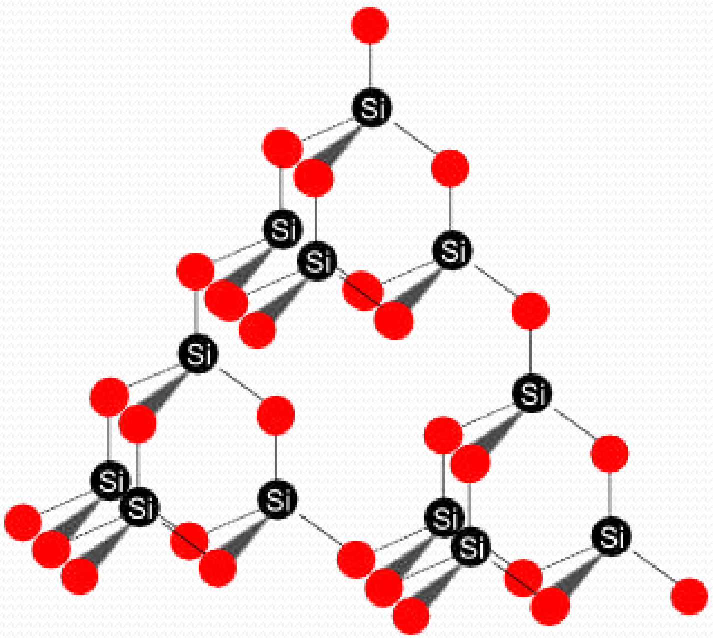
- HF Hydrogen bonding

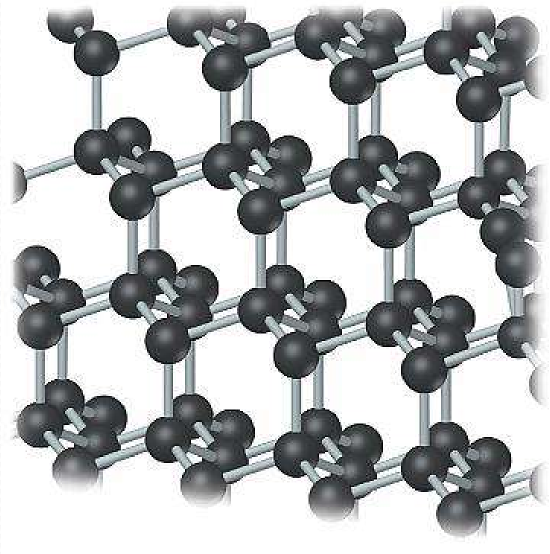
Practice Problems

- Determine the intermolecular forces that exist between the following molecules.
- NH_3 Hydrogen bonding
- O_2 London dispersion
- PCl_3 Dipole-dipole

Network Solids

- A network solid is a solid in which every atom is covalently bonded to one another.
- A network solid is extremely strong and has a very high melting point. Ex: diamond




(a) Diamond

Section 8.4 Assessment

1. How do electronegativity values determine the charge distribution in a polar covalent bond?
2. What happens when polar molecules are between oppositely charged metal plates?
3. Compare the strengths of intermolecular attractions to the strengths of ionic bonds and covalent bonds.
4. Not every molecule with polar bonds is polar. Use CO_2 as an example.

Section 8.4 Assessment

5. Draw the Lewis dot structure for each molecule. Identify the partial positive and partial negative atoms in each molecule.

a. HOOH

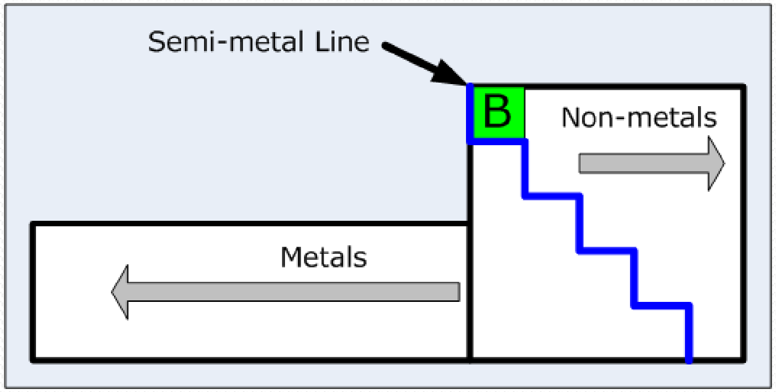
b. BrCl

c. HBr

d. H_2O

Section 9.3 – Naming with Nonmetals

- Nonmetals are to the right of the stair-step line on the periodic table.
- When naming with nonmetals, you name the first element and add –ide to the second element.
- You also must add prefixes to indicate the number of atoms of each element.
- The only time that you do not need a prefix is when the first element only has one atom.



Prefixes

Number	Prefix
1	Mono
2	Di
3	Tri
4	Tetra
5	Penta
6	Hexa
7	Hepta
8	Octa
9	Nona
10	Deca

Sample Problem

- Write the name of the following molecules.
- H_2O dihydrogen monoxide
- CO_2 Carbon dioxide
- CO Carbon monoxide

Practice Problem

- Write the name of the following molecules.
- PCl_3 Phosphorus trichloride
- N_2O_5 Dinitrogen pentoxide
- NH_3 Nitrogen trihydride

Writing Formulas with Nonmetals

- When writing the formula of a compound that starts with a nonmetal, you do not need to balance charges because the prefixes tell you the number of atoms.

Nitrogen Trichloride
= NCL_3

Sulphur Dibromide
= SBr_2

Dihydrogen Monoxide
= H_2O

Sample Problem

- Write the formula for the following compounds.

- Carbon tetrachloride

  CCl_4

- Sulfur hexafluoride

  SF_6

Practice Problem

- Write the formula for the following molecules.

- Diphosphorous trioxide

  P_2O_3

- Bromine monoiodide

  BrI

Section 9.3 Assessment

1. What information do prefixes tell you?
2. Write the names for the following molecules.
   a. NI_3
   b. BCl_3
   c. Cl_2O_7
   d. SO_3
   e. N_2H_4
   f. N_2O_3
   g. CS_2

Section 9.3 Assessment

3. Write the formula for the following compounds:
   a. carbon tetrabromide
   b. phosphorus pentachloride
   c. iodine heptafluoride
   d. chlorine trifluoride
   e. iodine dioxide
4. The name a student gives for the molecular compound $SiCl_4$ is monosilicon tetrachloride. Is this correct?

THE END