Chemical Bonds

Chemical bonds are interactions of electrons leading to strong forces of attraction which holds atoms together in molecules and compounds. Atoms may **transfer** or **share** electrons, and either process may provide for a stable arrangement of electrons between the atoms that results in the formation of molecules.

Rules for Electron Dot Structures and Bonding Structures

The central atom follows the Octet Rule (usually) and in most cases the least electronegative nonmetal is surrounded by the other atoms. Other atoms follow the Octet Rule whenever electrons are available (see exceptions in class). Check to see that every atom has the influence of 8 electrons and the total number of electrons is correct for that molecule.

Drawing bonding structures (called Lewis structures)

1. select a reasonable "skeleton" for the molecule or polyatomic ion
   a. the LEAST electronegative element is usually the central element, except that hydrogen never is
      example: S C S in compound CS_2
   b. oxygen atoms do not bond to each other except in a few cases such as O_2 and O_3
2. calculate the total number of outer shell electrons available in all the atoms of the molecule or ion
3. draw a single bond to represent each pair of shared electrons in the skeleton
4. allowing 2 electrons for each shared pair, subtract the total number of electrons already used
5. distribute the remaining electrons in such a fashion as to give each element an octet, if possible
6. for ions, be sure to add (for negative ions) or to subtract (for positive ions) the number of electrons indicated by the charge on the ion
7. remember that you can use double or triple bonds in order to give elements an octet, but only when necessary
8. if there are any electrons "left over", place these additional unshared (lone) pairs of electrons into the skeleton to fill the octet of every group 1,2, 13, 14, 15, 16, 17 element (except hydrogen, which can only share 2 electrons).

Ionic Bonds

- metals and nonmetals react chemically by the ***TRANSFER*** of electrons (from metals to nonmetals)

- metals form positive ions by losing valence electrons to the nonmetals which then form negative ions

- positive ions are strongly attracted to the negative ions by the electrostatic attraction that exist between unlike charges

- the new substance formed does not resemble either of the original atoms

- this attraction binding unlike ions together is called ionic bonding

example: CaF_2

see classroom drawing

Covalent Bonds

- two or more atoms both of which tend to gain electrons during reactions (nonmetals) may combine by <u>sharing 1 or 2 or 3 pairs of electrons</u>

- the force holding the atoms together is due to the attraction of each atom for the electrons that are held jointly (a stable condition)

- **<u>HYDROGEN, CARBON, NITROGEN, AND OXYGEN</u>** are noted for forming covalent bonds

single covalent bond: see examples in classroom

double covalent bond:

triple covalent bond:

Homework/Test Problems

First determine if the molecule is ionic or covalently bonded. Then draw the electron dot structures showing an acceptable bonding structure.

1. H_2S	2. F_2	3. HF	4. H_2O	5. AlF_3
6. MgO	7. NH_3	8. PBr_3	9. CCl_4	10. CS_2
11. CO_2	12. K_2S	13. CH_4	14. C_2H_2	15. $MgCl_2$
16. SiO_2	17. NF_3	18. HCl	19. $CHCl_3$	20. C_2F_2
21. C_2H_6	22. C_2H_4	23. CHN	24. Si_2F_4	*25. BF_3

Bonding Information

The difference in the electronegativities of two elements can be used to predict the nature of the bond. When this difference is small, the bond is primarily covalent. As the difference increases, the covalent bonds become increasingly polar. When the difference becomes even greater, the bond becomes ionic.

Generally the line is drawn at 1.7. When the differences in electronegativities is greater than 1.7 the bond is ionic (and less than 1.7 is covalent). Another boundary often is drawn at a difference of 1.0 (sometimes 0.8) to separate polar bonds from nonpolar bonds.

When a molecule behaves as if one end were negative and the opposite end positive, the molecule is said to be polar. Polar molecules are known as dipoles. A molecule is polar when there is an uneven distribution of electrons in the molecule.

When two atoms of the same element form a molecule, the shared electrons are equidistant from the nuclei of the two atoms. This makes the bond nonpolar.

HCl is an example of a two-atom polar molecule. The shared electron pair is attracted toward the highly electronegative chlorine atom and away from the hydrogen atom. The resulting concentration of negative charge is closer to the chlorine atom and that end of the molecule will be slightly negative. The other end will be slightly positive but the molecule as a whole will be neutral.

Summary:

1) compounds or bonded atoms in molecules are polar is the center of positive charge does not coincide with the center of negative charge.

2) when a covalent bond is formed between atoms of different electronegativities, the pair of electrons will be more closely associated with the more electronegative atom, and the resulting covalent bond will be somewhat polar.

3) the greater the difference between the electronegativities of the atoms involved in the bond, the greater the polarity of the bond.

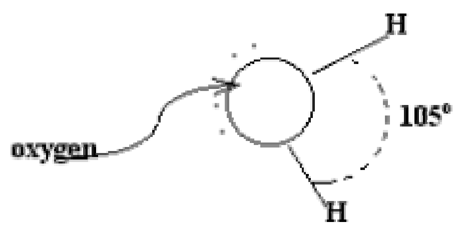
4) if the difference in electronegativity is too large, the electrons will be transferred and ionic bonding will result instead.

5. if both atoms in covalent bond have identical ionization potentials and electronegativities, no ions are formed and there is no polarity.

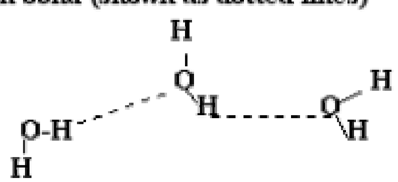
Hydrogen bonds: In compounds such as water, ammonia (NH_3), and hydrogen fluoride (HF), the hydrogen atoms are bonded to small atoms of high electronegativity (oxygen, nitrogen, and fluorine, respectively). The hydrogen atom has only a very small share of the electron pair that forms the bond. Such molecules are highly polar. In fact, each hydrogen atom acts largely as exposed proton. It can be attracted to, and form a weak bond with, the highly electronegative atom of a neighboring molecule. This is called a hydrogen bond. It is more than just an electrostatic attraction between opposite charges. It actually has some covalent character.

Hydrogen bonding is responsible for a number of unusual properties. Hydrogen bonding occurs between water molecules. Water must therefore be raised to a much higher temperature before the kinetic energy of its molecules becomes great enough to break the hydrogen bonds between the molecules. Breaking these hydrogen bonds is necessary in order to boil water. X ray studies show that the three-dimensional structure caused by hydrogen bonding gives ice crystals a crystalline arrangement with many hexagonal openings. This open structure accounts for the low density of ice.

Shape of water molecule (called bent or angular shape)



Hydrogen bond (shown as dotted lines)



Metallic Bonds

Most metals have only one or two valence electrons and low ionization energies. The valence electrons do not seem to belong to any individual atom but move easily from one atom to another. Metals can be thought of as positive ions immersed in a "sea" of mobile electrons. The attractive forces that bind metals atoms together are called metallic bonds. The ease with which the valence electrons move within the crystal distinguishes the metallic bond from ionic or covalent bonds.

a) metals are good conductors of heat and electricity because of the mobility of their valence electrons.

b) High luster of metals is the result of the way in which valence electrons absorb and re-emit light energy that strikes them

c) Metals can be flattened out or stretched out into a wire because the electrons and ions can move into other positions without breaking up the essential structure.

Summary:
The forces between ions are very strong; so that ionically bonded substances have high melting and boiling points, and are usually solids at room temperature. Water is usually capable of dissolving them.

Atoms in covalently bonded substances are electrically neutral, do not conduct electricity, have low melting and boiling points, and are gases or volatile liquids at room temperatures. Organic solvents will often dissolve them.

Extension Information on Bonds

Ionic Bonds (have large differences in electronegativity

Ionic Crystals – electrostatic attractions between ions, NO MOLECULES. Nondirectional bonds; localized electrons on ions. Examples: $NaCl$, K_2SO_4, NH_4Cl, $(NH_4)_2SO_4$

Crystal properties:

1. medium high melting point (600 - 2000° C)

2. medium high boiling points

3. hard and brittle

4. nonconductor of electricity

5. poor conductor of heat

Molecular Crystals – small individual molecules held internally together by covalent directional bonds. The electrons are localized on molecules. The molecules are attracted to each other by (1) dipole attraction (2) Van der Waal forces (3) hydrogen bonds. Examples: HCl, SO_2, CO_2, CH_4, H_2SO_4, H_2O

Crystal properties:

1. very low melting point (-370 to 300° C)

2. very low boiling point

3. soft

4. nonconductor of electricity

5. poor conductor of heat

Covalent Bonds (only very small differences in electronegativity)

Covalent Crystals – all atoms in the crystal are inter bonded by covalent bonds to make one large crystal. The electrons are usually localized in the bonds. Examples: diamond, SiC, SiO_2, graphite

Crystal properties:

1. very high melting point (1200 - 3500° C)

2. very high boiling point

3. very hard and brittle

4. usually a nonconductor of electricity

5. usually a poor conductor of heat

Metallic Bonds

Metallic Crystals – positive nuclei lattice in a cloud of delocalized electrons. Examples: Hg, Cu, Au, Fe, alloys.

Crystal properties:

1. very low to very high melting point

2. very low to very high boiling point

3. very soft to very hard

4. ductile and malleable

5. good conductor of heat and electricity

Bonding Geometry

Total Pairs	Bonding Pairs (central atom excluding double bonds	Nonbonding pairs (central atom)	Overall Geometry	Molecular Geometry	Example	Hybridization
1	1	0	linear	linear	H_2	none
2	2	0	linear	linear	BeF_2	sp
3	3	0	trigonal planar	trigonal planar	BF_3	sp^2
	2	1		bent	(SO_2)	
4	4	0	tetrahedral	tetrahedral	CH_4	sp^3
	3	1		pyramidal	NH_3	
	2	2		bent	H_2O	
	1	3		linear	HCl	
5	5	0	trigonal bypramid	trigonal bipyramid	PF_5	dsp^3
	4	1		distorted tetrahedron	SF_4	
	3	2		T-shaped	ClF_3	
	2	3		linear	XeF_2	
6	6	0	octahedral	octahedral	SF_6	d^2sp^3
	5	1		square pyramid	ClF_5	
	4	2		square linear	XeF_4	