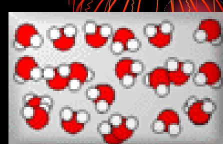
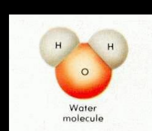
Section 6.2

Covalent Bonding
and
Molecular Compounds

Most Chemical Compounds

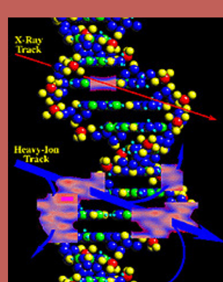
- Are *molecules*, a neutral group of atoms that are held together by covalent bonds. It is a single unit capable of existing on its own.



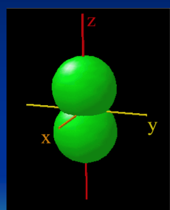


Molecular (or Covalent) Compound

- A chemical compound whose simplest units are molecules.
- They may be made up of one type of atom only (O_2), or of two or more different atoms (H_2O).



Diatomic Molecules



- A molecule containing only two atoms (O_2 or CO).

Chemical vs. Molecular Formulas

Molecular formula: Shows the types and numbers of atoms combined in a single molecule of a molecular compound.

Chemical formula: Indicates the relative numbers of atoms of each kind in a chemical compound by using atomic symbols and numerical subscripts.

So What's The Difference?

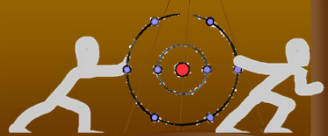
- An example for both a chemical vs. molecular formula:

H_2SO_4 (sulfuric acid)

The first (chemical) talks about moles of molecules, the second (molecular), talks about <u>one</u> molecule.

Formation of a Covalent Bond

- Using the white boards in your table groups, draw and describe the attractive and repulsive forces involved during the creation of a covalent bond.

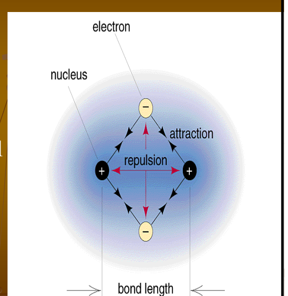


Formation of a Covalent Bond

From before, it is a trade off between the attractive and repulsive forces between electrons and the nuclei of atoms.

Attraction: Decrease in potential energy of the atoms.

Repulsion: Increase in potential energy of the atoms.

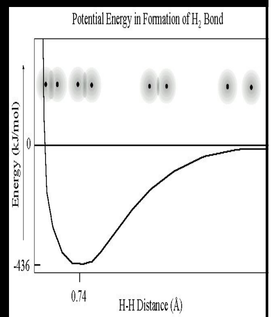


Remember the potential energy graph from the online tutorial?

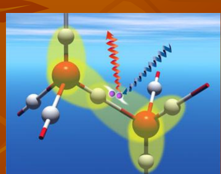
- Draw the graph on your white board. Practice explaining what it means to each other in terms of how the distance between atoms (and those attractive and repulsive forces) affects the potential energy of two atoms.

Nature favors arrangements where potential energy is minimized.

Bonding occurs at the energy minimum (the repulsion of like charges equals the attraction of unlike charges).

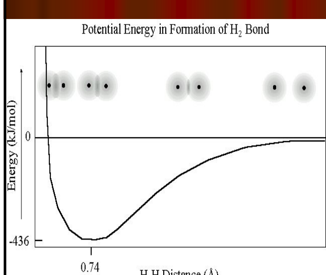


Bond Length



- The distance between two bonded atoms at their minimum potential energy (the average distance between the two bonded atoms).

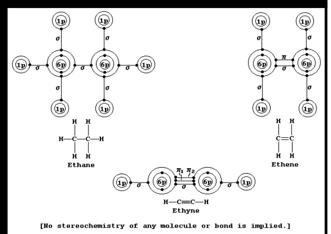
Bond Energy



The energy required to break a chemical bond and form neutral isolated atoms. Really the same amount that was released when the bond formed, but with different sign.

The Lengths and Energies Vary

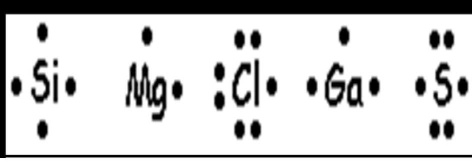
- With the types of atoms that have combined and may vary between the same two atoms, depending on what other bonds the atoms have formed.



Electron Dot Notation

An electron configuration notation in which only the valence electrons of an atom of a particular element are shown, indicated by dots placed around the element's symbol.

You put one single dot on each side first and then double up, if needed.



Practice Problem #1

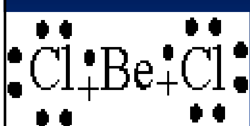
- 37. Use electron dot-notation to illustrate the number of valence electrons present in one atom of each of the following elements.
- a. Li
- b. Ca
- c. Cl
- d. O
- e. C
- f. P
- g. Al
- h. S

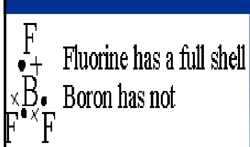
Octet Rule

Chemical compounds tend to form so that each atom, by gaining, losing, or sharing electrons, has an octet (8) electrons in its highest occupied energy level.

This fills the *s* and *p* orbitals, just like a noble gas

Exceptions to the Octet Rule

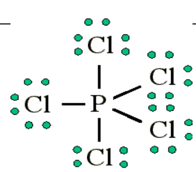


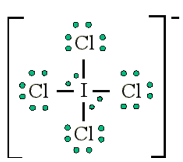


- Hydrogen (why?)
- Lithium
- Beryllium
- Boron
- Molecules with an odd number of electrons (N-O).

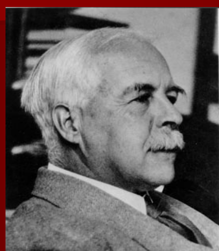
Expanded Valence (More Than 8)

- Involves bonding in *d* orbitals as well as *s* and *p* orbitals.
- Usually occurs with bonding the highly electronegative elements F, O, and Cl.





Lewis Structures

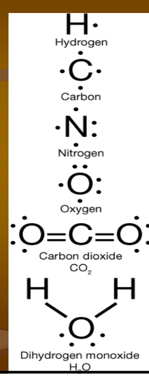


G. N. Lewis
1875 - 1946

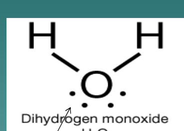
- Using electron-dot notation to represent molecules of compounds.

Lewis Structures

- Formulas in which atomic symbols represent nuclei and inner-shell electrons, dot-pairs or dashes represent electron pairs in covalent bonds, and dots adjacent to only one symbol represent unshared electrons.



Lone Pairs



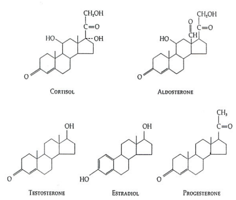
1 Lone Pair

Also called unshared pairs. A pair of electrons that is not involved in bonding and that belongs exclusively to one atom.

Represented as two dots on one side of an atom in a Lewis structure

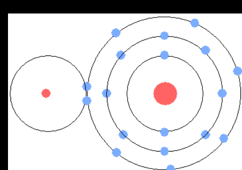
Structural formula:

- Indicates the kind, number, arrangement, and bonds, but not the unshared pairs of the atoms in a molecule.
- Ex. F-F H-Cl



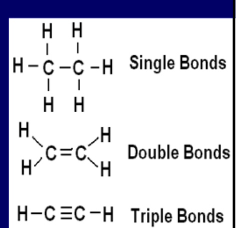
Single Bond

- A covalent bond produced by the sharing of one pair of electrons between two atoms.

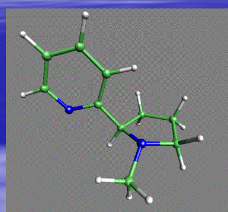


Multiple Bonds

- Double and triple bonds are possible. This happens when atoms share 2 or 3 electron pairs to form covalent bonds to conform to the octet rule. They have progressively shorter bond lengths and higher bond energies than single bonds between the same atoms.
- **Extremely** common for C, N, and O.

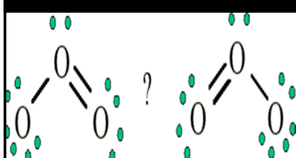


See Table 6-2, p. 173



- As should be noted from the table, multiple bonds between pairs of C, N, and O are *very* possible. Multiple bonds are needed when there are not enough valence electrons to complete the octet.

Resonance



Bonding in molecules or ions that cannot be correctly represented by a single Lewis structure.

Experimental Evidence

- Shows that there is not a mixture of single and double bonds, but that all of the bonds are identical, so there is an average of the structures. We show all possible structures, using a double-headed arrow.