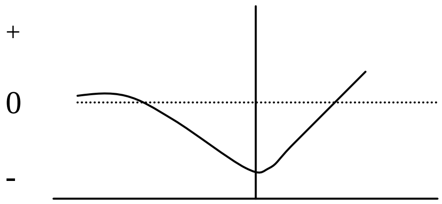
Chemical Bonding
Honors Chemistry Lesson

12.1-5
Linus Pauling:

Bonding: Measurement of force of attraction between 2 atoms. A bond has a lower potential energy than when separate.



Valence electrons positioned between the atoms.

Electronegativity: Ability of an atom within a bond to attract electrons to itself. P.332
Trend

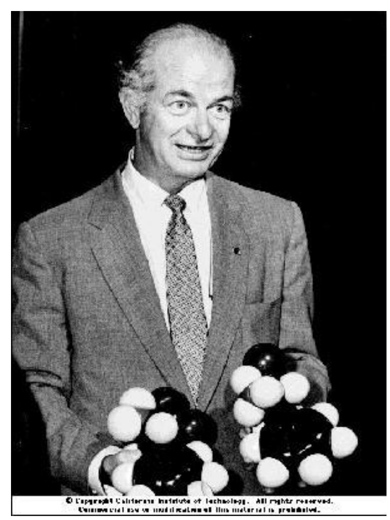
In almost every case in which a bond is formed between two different atoms the resulting bond will be polar.

In the 1930's, Linus Pauling (1901 - 1994), an American chemist who won the 1954 Nobel Prize, recognized that bond polarity resulted from the relative ability of atoms to attract electrons. Pauling devised a measure of this electron attracting power which he called "*electronegativity*" which he defined as the "power of an atom in a molecule to attract electrons to itself." Electronegativity only has meaning in a bond.

The table below presents the electronegativities for the main group elements.

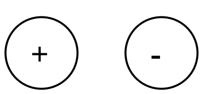
Electronegativity

H = 2.1	x	x	x	x	x	x
Li = 1.0	Be = 1.5	B = 2.0	C = 2.5	N = 3.0	O = 3.5	F = 4.0
Na = 0.9	Mg = 1.2	Al = 1.5	Si = 1.8	P = 2.1	S = 2.5	Cl = 3.0
K = 0.8	Ca = 1.0	Ga = 1.6	Ge = 1.8	As = 2.0	Se = 2.4	Br = 2.8
Rb = 0.8	Sr = 1.0	In = 1.7	Sn = 1.8	Sb = 1.9	Te = 2.1	I = 2.5
Cs = 0.7	Ba = 0.9	Tl = 1.8	Pb = 1.9	Bi = 1.9	Po = 2.0	At = 2.2

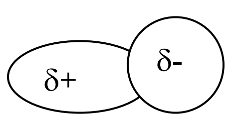


Generally, the electronegativity increases moving left to right across a row, and decreases going down the table. Notice that this trend is violated by the Group 13 metals for which the electronegativity drops from B to Al as expected, but hen rises slightly going down to Tl. This effect is due to the intervention of the d electrons and other effects that come into play with very large atoms.

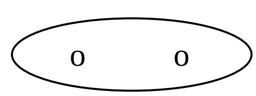
Large difference (>ΔEN 1.7)
Ionic



Small difference
Polar Covalent



No difference
Covalent



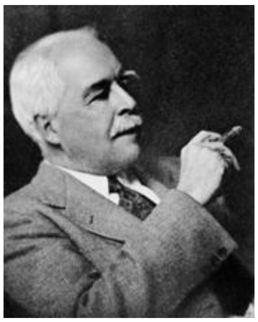
Types of bonds:

Ionic: Salts, transfer of 1 or more electrons from metal with low ionization energy to non-metal with high ionization energy.
Crystal lattice of repeating unit
Solids, high melting points, conduct electric current when melted or dissolved
Empirical formula

Covalent: Molecule, shares 1, 2, o r3 pairs of electrons among 2 non-metals with similar electro negativities.
Solids, liquids or gases. Have strong – weak intermolecular attractions.

Homework Practice: P. 360 # 1-15, 17, 19, 26, 27, 29, 31-33, 36, 37, 48

Lewis Structures: Gilbert Newton Lewis:



(Dot Structures): involves only valence electrons.
Examples using electron configuration and orbital filling diagrams

Octet Rule and stable (noble gas) configurations ands charges of ions:

Guideline to determine the number of bonds in a compound:
A.) Determine number of electrons needed to satisfy Octet (duet w/ H) count:
B.) Number of actual valence electrons:
Subtract these two (A-B) and divide by 2 = number of bonds

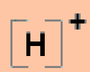
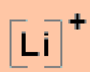
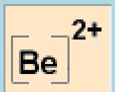
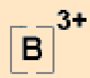
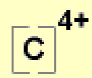
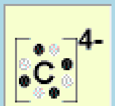
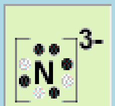
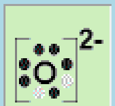
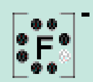
Lewis Structures of atoms

- The chemical symbol for the atom is surrounded by a number of dots corresponding to the number of valence electrons.

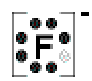
Number of Valence Electrons	1		2		3	4	5	6	7	8
Example	Hydrogen	Group I (Alkali metals)	Helium	Group II (alkali earth metals)	Group III	Group IV	Group V	Group VI	Group VII (Halogens)	Group VIII except Helium (Noble Gases)
Lewis Structure (electron dot diagram)	H•	Li•	He:	Be:	B	C	N	O	F	Ne

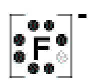
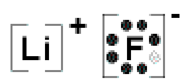
Lewis Structures for Ions of Elements

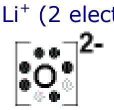
- The chemical symbol for the element is surrounded by the number of valence electrons present in the **ion**. The whole structure is then placed within square brackets, with a superscript to indicate the charge on the ion.
- Atoms will gain or lose electrons in order to achieve a stable, Noble Gas (Group VIII), electronic configuration.
- Negative ions (anions) are formed when an atom gains electrons.
- Positive ions (cations) are formed when an atom loses electrons.

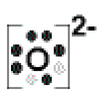
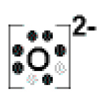
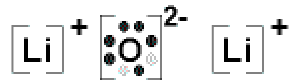
Charge on Ion	1+		2+	3+	4+	4-	3-	2-	1-	
No. electrons gained or lost	1e lost		2e lost	3e lost	4e lost	4e gained	3e gained	2e gained	1e gained	
Example	H^{+}	Group I $^{+}$ (Alkali metals)	Group II $^{2+}$ (alkali earth metals)	Group III $^{3+}$	Group IV $^{4+}$	Group IV $^{4-}$	Group V $^{3-}$	Group VI $^{2-}$	Group VII $^{-}$ (Halogens)	H^{-} (hydride)
Lewis Structure (electron dot diagram)	 OR H^{+}	 OR Li^{+}	 OR Be^{2+}	 OR B^{3+}	 OR C^{4+}					

Lewis Structures for Ionic Compounds

- The overall charge on the compound must equal zero, that is, the number of electrons lost by one atom must equal the number of electrons gained by the other atom.
- The Lewis Structure (electron dot diagram) of each ion is used to construct the Lewis Structure (electron dot diagram) for the ionic compound.
- Examples
  - Lithium fluoride, LiF
    Lithium atom loses one electron to form the cation Li^{+}
    
    Fluorine atom gains one electron to form the anion F^{-}
    Lithium fluoride compound can be represented as

    Li^{+}  OR 

  - Lithium oxide, Li_2O
    Each lithium atom loses one electron to form 2 cations Li^{+} (2 electrons in total are lost)
    
    Oxygen atom gains two electrons to form the anion O^{2-}
    Lithium oxide compound can be represented as

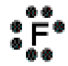
    $2Li^{+}$  OR Li^{+}  Li^{+} OR 

Lewis Structures for Covalent Compounds

- In a covalent compound, electrons are shared between atoms to form a covalent bond in order that each atom in the compound has a share in the number of electrons required to provide a stable, Noble Gas, electronic configuration.
- Electrons in the Lewis Structure (electron dot diagram) are paired to show the bonding pair of electrons.
- Often the shared pair of electrons forming the covalent bond is circled
- Sometimes the bond itself is shown (-), these structures can be referred to as *valence structures*.
- Examples

- hydrogen fluoride, HF

Hydrogen atom has 1 valence electron

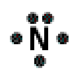


Fluorine atom has 7 valence electrons
Hydrogen will share its electron with fluorine to form a bonding pair of electrons (covalent bond) so that the hydrogen atom has a share in 2 valence electrons (electronic configuration of helium) and fluorine has a share in 8 valence electrons (electronic configuration of neon)
Lewis Structure (electron dot diagram) for hydrogen fluoride

OR

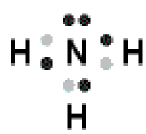
Valence Structure for hydrogen fluoride

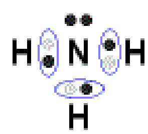
- ammonia, NH_3



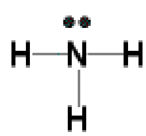
Nitrogen atom has 5 valence electrons

Hydrogen atom has 1 valence electron
Each of the 3 hydrogen atoms will share its electron with nitrogen to form a bonding pair of electrons (covalent bond) so that each hydrogen atom has a share in 2 valence electrons (electronic configuration of helium) and the nitrogen has a share in 8 valence electrons (electron configuration of neon)
Lewis Structure (electron dot diagram) for ammonia



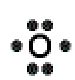


OR



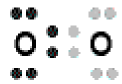
Valence Structure for ammonia

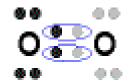
- oxygen molecule, O_2



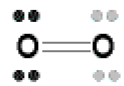
Each oxygen atom has 6 valence electrons
Each oxygen will share 2 of its valence electrons in order to form 2 bonding pairs of electrons (a double covalent bond) so that each oxygen will have a share in 8 valence electrons (electronic

configuration of neon).
Lewis Structure (electron dot diagram) for the oxygen molecule





OR



Valence structure for the oxygen molecule

List of Compounds:
Draw Lewis structures for all compounds that obey the octet.

1.) NaCl
2.) H_2O
3.) NH_3
4.) $MgBr_2$
5.) CH_4
6.) CaO
7.) CO
8.) CO_2
9.) CO_3^{2-}
10.) CH_3OH
11.) SO_2Cl_2
12.) OCl_2
13.) SO_2
14.) SO_3
15.) ClO_3^-
16.) N_2O
17.) O_3
18.) C_2H_6
19.) C_2H_4
20.) C_2H_2
21.) N_2
22.) O_2
23.) Cl_2
24.) HCN
25.) NF_3
26.) CF_2Cl_2
27.) H_2CO
28.) NH_4^+
29.) NO_2^-
30.) NO_3^-
31.) ClO_2^-
32.) SO_4^{2-}
33.) SO_3^{2-}
34.) Na_2SO_4
35.) N_3^-
36.) NaOH
37.) PO_4^{3-}
38.) KNO_3
39.) C_6H_6
40.) C_2H_5OH

<u>Bond length and strength:</u>

	Single	Double	Triple
# electrons:	2	4	6
Length:	longest -----------		shortest
Strength:	weakest ----------		strongest

<u>Exceptions to the octet:</u>

Be, B e- deficient: Ex. $BeCl_2$, BF_3

Expanded octet: Period 3 or greater

To draw Lewis structures.

Satisfy octet rule for terminal atoms. Place extra lone pairs on central atom

.Ex. PF_5, Br_3^-, SF_4, XeF_4, ClF_5, SF_6, KrF_2, SeF_4, ClF_3, XeO_3, $RnCl_2$, ICl_4^-

Molecular Geometry -VSEPR (Valence Shell Electron Pair Repulsion)

Help Link: **http://www.shef.ac.uk/chemistry/vsepr/**
Excellent Information: http://www.faidherbe.org/site/cours/dupuis/vseprev.htm
Text, Table 12.4

Lewis structures and bond angles

Total Number of **Atoms + Lone pairs** around central atom	Geometry			
No Central Atom	Linear O_2 N_2	With Lone Pairs O_2 N_2		
2	Linear CO_2			
3	Trigonal Planar BF_3	Bent SO_2		
4	Tetrahedral CH_4	Pyramidal NH_3	Bent H_2O	
5	Trigonal Bipyramidal PF_5	See-saw SF_4	T-Shaped ClF_3	Linear XeF_2
6	Octahedral SF_6	Square Pyramidal BrF_5	Square Planar XeF_4	

Dipole Moments and Polarity:

Polar molecules: Dipole moment (μ)

- The polarity of a molecule is the vector sum of the individual bond's polarities.
- Molecular polarity can be can be measured and is expressed in units of Debye (D), higher D means nigher molecular polarity.
- To examine the polarity of a molecule with a dipole moment the center of positivity and negativity do not coincide.

Covalent bonds are polar when the bonded atoms have rather large differences in electronegativity. In the A–B bond, if A is more electronegative, the pair of electrons in that bond is closer to A than to B. Symbols are used to illustrate the slightly electron-rich atom A (δ–) and the equally electron-poor atom B (δ+). The polar A–B bond is a bond **dipole**, which can be represented by using a vector arrow pointing to the negative end of the dipole.

$$\begin{array}{c} A - B \\ \delta- \;\; \delta+ \end{array}$$

Just as bonds can be polar, entire molecules can also be polar. Molecules can be polar because there is a net sum of individual bond polarities. You can think about it this way: Assume there is a center of mass of all the positive charges (nuclei) in a molecule, and a center of mass of all negative charges (electrons). If these two centers do not coincide, the molecule is polar.

Determine the polarity of all Lewis structures