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CHEM 121.1
General Chemistry II
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Chemical Bonding Lewis

- To understand why molecules form and reactions happen, look at bonding
- When electrons are bonded they want to be paired, but if an element is just by itself you want to keep the valence electrons unpaired when possible

Summary of Topics

- Lewis structures
    - focus on valence electrons, illustrate electronic aspects of bonding
- VSEPR theory
    - shape, extends Lewis theory into 3D picture of geometry
- Valence Bond theory
    - combine VSEPR and Lewis Structures plus quantum chemistry, localized orbitals
- MO theory
    - diatomics only in this class, delocalization of atomic orbitals

Two types of bonds (focus on groups I-VIII)

- Ionic bond: results from the electrostatic attraction between metal cation and nonmetal anion

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Ex: Na. + .F: → Na-F
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Complete transfer of e - from Na to F (more EN)

- Covalent bond: results from the sharing of electrons between two nonmetals. Remember EN is how much an atom wants e- in a bond. Focuses on groups II-VIII

F-F no difference in EN , NONPOLAR covalent bond, no one wins the tug of war
F-Cl nonzero difference in EN, POLAR covalent bond, one wins the tug of war
F is $\delta-$
Cl is $\delta+$

- Lewis structures can have both covalent and ionic bonds, but mostly covalent

Relative strengths of bonds
nonpolar covalent> polar covalent> ionic

Examples of each:

- Non polar:diamond (Carbon and coal), strongest, takes the most energy to break the bond
- Polar covalent: H_2O
- Ionic: NaCl, weakest, takes the least amount of energy to break bond