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CHEM 121.1
General Chemistry II
Anne Yu

Chemical Bonding Lewis

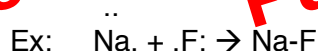
- To understand why molecules form and reactions happen, look at bonding
- When electrons are bonded they want to be paired, but if an element is just by itself you want to keep the valence electrons unpaired when possible

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



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 δ^-

Cl is δ^+

- 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: H₂O
- Ionic: NaCl, weakest, takes the least amount of energy to break bond