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
  Ex: Na⁺ + F⁻ → Na-F
  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⁻ and 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