Lecture 8, Jan 28, 2022

Atomic Bonding

- Atoms try to achieve the most stable (lowest energy) configuration by achieving closed shells
- Two types of bonds: primary (ionic, covalent and metallic) and secondary (dipole interactions); we will focus on the former
  - Primary bonds have a high separation energy
  - Ionic: electron exchange, covalent: electron sharing, metallic: delocalized sharing
    * Delocalized electrons are not bound; they have enough energy that they're not stuck in a potential well
- Most elements have unstable electron configurations; e.g. Lithium $1s^2 2s^1$, while elements such as Neon $1s^2 2s^2 2p^6$ have fully filled outer shells and are stable
  - Recall: s orbital has 2 electrons, p orbital has 6, d has 10, f has 14, which comes from the constraints on the quantum numbers and spin up/down electrons
  - *Electropositive* elements readily give up electrons to become positive ions; *electronegative* elements readily acquire electrons to become negative ions
- Electronegativity is quantified using the Pauling scale, ranging from 0.7 (francium) to 4.0 (fluorine)
- Ionic bond: local and discrete (an absolute amount of charge is moved), between metal (gives electron to become +) and nonmetal (takes electron to become -)
  - Happens when two elements differ greatly in electronegativity
  - Electrons are transferred from one atom to another, creating a positively charged and negatively charged atom, and a Coulomb attraction force
  - Example: $\mathrm{NaCl}, \mathrm{MgO}$