Chapter 6 – Electron Structure

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- Atoms gain and lose energy in set amounts – Quantized
- Lower energy level is ground state, higher is called excited
- An atom which gains energy moves electrons to a higher energy level
- An atom which loses energy has electrons move back down (Electron Jumping)
- Energy is seen as different wavelengths of light in a flame test
- $E = hV$ and $C = \lambda V$, where h is Planck's constant, V is the frequency, λ is the wavelength, C is the speed of light, and E is energy
- Bohr's Model – Electrons orbit the nucleus, and, when they gained energy, jump up to a new level
- Quantum Mechanical Model – It is unknown how electrons move, but we know where they probably are, which is demonstrated in probability maps
- Probability Maps – Orbitals (Four Types) s, p, d, and f (sometimes called sublevels)
- s forms a circular probability, p forms a 2 leaf clover, d forms 4 leaf clover, and f is technically 6, but is hard to map out
- 2 Electrons per orbital
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Type	Orbitals	Electrons
s	1	2
p	3	6
d	5	10
f	7	14
- Electron configuration and Box diagrams (often called Orbital Diagrams)

- Quantum Numbers:

  1. Energy Level (n)
  2. Sublevel (l): Type (s=0; p=1; d=2; f=3)
  3. Box number (Number of orbitals, m_l): $-l \leq m_l \leq l$
  4. Spin (m_s): $-\frac{1}{2} \leq m_s \leq \frac{1}{2}$

- Hund's Rule − Electrons spread out

- Pauli Exclusion Principle − No two electrons have the same 4 quantum numbers

- Noble Gas Short-hand:

  1. Start with the noble gas from the row above (for example, for As it would be Ar)
  2. Then, start counting from the s type, with energy level + 1 of the noble gas (4 s for As)

- Amount of valence electrons is determined by the column an element is in (5 for As)

- When an ion is given, electrons are first pulled from the s orbital

  1. Example: Fe to Fe^{3+}
  2. Fe: $[Ar]\,4\,s^2\,3\,d^6$
  3. Fe^{3+}: $[Ar]\,3\,d^5$

- Periodic Trends:

  1. Radius (size)
     (a) The closer to the bottom left, the bigger the size of the atom
     (b) Greatest: Francium (Fr)
     (c) Less size due to more protons, causing greater strong and weak nuclear forces
     (d) Electron Shielding: Electrons in lower energy levels shield outer electrons from nuclear charge
     (e) Negative ions are larger than a neutral atom, while positive are smaller
     (f) Ex. Fe^{2+} vs Fe^{3+}: Fe^{2+} is large since more electrons mean more electron repulsion making it larger
  2. Ionization
     (a) The closer to the top right, the more ionized an atom (needs more energy to remove an electron)
     (b) Greatest: Fluorine (F)
     (c) Two Exceptions: Be and Mg need larger ionization energy than B and Al, respectively, because 2p is higher in energy than 2s, which means it takes less energy to remove from B

3. Electronegativity
    (a) The closer to the top right, the more electronegative an element (has more of a relative pull for shared electrons in a bond)
    (b) Greatest: Fluorine (F)

- Bonds:
    1. Ionic – e^- are transferred (atoms bonded this way will be far across the periodic table from each other). Atoms become polar (dipole).
    2. Polar Covalent – e^- are shared unevenly (atoms bonded this way will be relatively close to each other). Atoms also become polar (dipole)
    3. Covalent – e^- are shared evenly (no plus or minus end to it).