Lecture 7, Jan 27, 2022

Electron Configurations

- Recall the restrictions on the atomic quantum numbers: $n > 0, 0 \leq l \leq n-1, |m_t| \leq l$
  - n designated $K, L, M, N, \cdots$, determines the "shell", i.e. average distance from the nucleus; within the principal energy state there are energy sub-states
  - l designated s, p, d, f, orbitals within the shell (angle from the z axis)
  - m_l, designated $1, 3, 5, 7, \cdots$ (number of states), (motion around the z axis)
  - m_s, $\pm\frac{1}{2}$, spin up or down
- As we go out we have more capacity for electrons ($n = 1$ has 2 spots, $n = 2$ has 8 spots, $n = 3$ has 18, etc), which matches the periodic table
- Each energy sub-level relates to a different orbital shape
  - s orbitals ($l = 0$) are all spherical, regardless of n
  - p orbitals are dumbbell shaped and oriented along the 3 spacial axes (p_x, p_y, p_z)
  - d orbitals have 5 orbitals, 4 with the same clover shape in different planes, and the 5th has its own shape
- Electrons have discrete energy states and tend to occupy the lowest energy state available
  - Energy increases as n increases
  - Within the principal quantum numbers the higher numbers of l have more energy
  - However the $3d$ shell has a higher energy than the $4s$ shell
- When a shell is completely filled (with its sub-shells) it is stable; most elements don't have a stable outer shell, the stable ones are the noble gases with the outer shells completely filled
- Electron configurations in the form e.g. $1s^2$ where the 1 indicates principal quantum number, s indicates shell (l) and the 2 is the number of electrons in that shell (spin up + spin down)