Electron Configuration & Chemical Periodicity

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Each **orbital** can be described by **three quantum numbers**. But, each **electron** in an atom can be described by **four quantum numbers**.

- **m(sub(s)):** describes spin, either positive or negative 1/2
  - *Up arrow:* "spin up" electron
  - *Down arrow:* "spin down" electron
- **Pauli Exclusion Principle:** two **electrons** in the *same* atom **cannot** have the same four **quantum numbers**. In each orbital, there is on electron with m(sub(s)) = +1/2 and on with m(sub(s)) = -1/2.
- **Electron configuration:** describes how electrons are distributed in atomic orbitals
  - H - 1s (n = 1, l = 0, ml = 0, m(sub(s)) = -1/2
  - He - 1s2 (n = 1, l = 0, ml = 0, m(sub(s)) = 1/2
- Sublevel energies and order for filling electrons:
  - 3d is higher than 4s
  - Electrons enter orbitals of lower energy first
  - Orbitals are filled before moving to a higher energy orbital (**Aufbau Principle**)
- **Ground state:** lowest energy config.
- **Excited state:** higher energy config. than ground state
- **Hund's Rule:** when orbitals of equal energy are available, the ground state has the max number of unpaired electrons with parallel spins
- **Paramagnetic:** atom with one or more unpaired electrons, attracted to magnetic field

Exceptions to Aufbau Principle (know just these 5 exceptions)

Cr: [Ar] 4s2 3d4 is wrong, really [Ar] 4s1 3d5 (also true for Mo); half-filled d orbitals are lower energy than d4

Cu: [Ar] 4s2 3d9 is wrong, really [Ar] 4s1 3d10 (also Ag and Au); completely filled d orbitals are lower energy than d9

Some definitions . . .

- Fe: [Ar] 4s2 3d6
  - **Inner (core) electrons** - in common w/ previous noble gas and any filled d