

Predicting Atomic Electron Configurations

- 1) Electrons occupy the lowest energy orbitals available - *the n+l Rule*
Begin assigning electrons at **1s** and continue in the following order:
1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p etc.
Examples: **Li:** $1s^2 2s^1$ **Na:** $1s^2 2s^2 2p^6 3s^1$ **Ca:** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
- 2) **s** orbitals have one subshell; **p** orbitals have three subshells; **d** orbitals have five subshells; **f** orbitals have seven subshells. Or:
Number of subshells = 2l+1
where $m_l = \{l, (l-1), \dots, 0, \dots, (-l+1), -l\}$... and $l = \{0, 1, \dots, n - 1\}$
- 3) No two electrons in an atom can have the same set of four quantum numbers - *Pauli Exclusion Principle*. Each subshell can hold only two electrons, and the two electrons must have opposite values of spin (i.e. m_s).
- 4) The most stable arrangement of electrons is that with the maximum number of unpaired electrons - *Hund's Rule*. Single electrons must occupy every subshell in an orbital before they "pair up" or are "spin paired".
Example: **Ti:** $[\text{Ar}]3d^2 4s^2$ Titanium has two unpaired electrons
- 5) *Paramagnetic* compounds contain unpaired electrons.
Diamagnetic compounds contain electrons that are exclusively "spin paired." No unpaired electrons exist in diamagnetic compounds.
Examples: **Zn:** $[\text{Ar}]3d^{10} 4s^2$ (*diamagnetic*) **Li:** $[\text{He}]2s^1$ (*paramagnetic*)
- 6) *Atomic ion configurations* can be assigned using the rules given above and while remembering that the electrons easiest to remove will generally come from the highest energy orbital available.
Examples: **Cu:** $[\text{Ar}]3d^{10} 4s^1$ **Cu²⁺:** $[\text{Ar}]3d^9$