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 $\mathbf{m_l = \{l, (l-1), \ldots, 0, \ldots, (-l+1), -l\}}$ … *and* $\mathbf{l = \{0, 1, \ldots n - 1\}}$

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. $\mathbf{m_s}$).

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**: $[Ar]3d^2 4s^2$ 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**: $[Ar]3d^{10} 4s^2$ *(diamagnetic)* **Li**: $[He]2s^1$ *(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**: $[Ar]3d^{10} 4s^1$ $\mathbf{Cu^{2+}}$: $[Ar]3d^9$

Periodic Table Blocks