Atomic Electron Configurations and Periodicity

Quantum #'s n, ℓ, m_ℓ do a good job of describing the tendency of electrons to be in a given place at a given time (orbital shape)

- BUT, one more identifying factor is needed:
  - Electrons behave as though they have a spin
  - This spin is quantized

- Defined by **electron spin magnetic quantum number, m_s**
  - m_s is either +½ or -½

- Distribution of spins determines magnetic properties

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Magnetism and Electron Spins

Behavior of a material in a magnetic field is due primarily to electron characteristics
- **Paramagnetism**: Attracted to a magnetic field
- **Diamagnetism**: Repelled (weakly) by a magnetic field

If only a single electron occupies a given orbital, it (regardless of m_s) is attracted to a magnetic field.

If two electrons occupy the same orbital they could have:
- Opposite spin (*paired spins*): **Diamagnetic**
- The same spin (*unpaired):* **Paramagnetic**

Special case of paramagnetism: **Ferromagnetism**
- Clusters of unpaired electrons where spins are aligned with one another

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Electronic Configuration

Pauli Exclusion Principle: no two electrons in an atom can have the same set of quantum numbers (n, ℓ, m_ℓ, m_s).

- Since m_s can only have two values, this means…

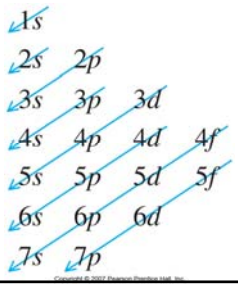
How do we determine which subshell an electron occupies?

Subshell Energies and Electron Configurations

- Current picture developed as a result of experiment.

2 General Rules:

1. Electrons are assigned to subshells in order of increasing "n + ℓ" values
2. If 2 subshells have the same "n + ℓ", electrons are assigned first to the subshell with the smallest n



Electron Bookkeeping and Notation

In describing electron configurations, the following rules apply:

1. The notation must account for all electrons
2. The Pauli exclusion principle must be satisfied
3. Orbitals are filled by increasing energy (***aufbau***)
4. Hund's Rule must be obeyed: most stable arrangement of electrons is that with the maximum number of unpaired electrons. Orbitals are filled one electron at a time until all orbitals of a subshell contain one electron, then any remaining electrons are added to complete the shell.

Electron Configuration Notation (*spectroscopic notation*) shows:

- Values for n and ℓ for each orbital that is occupied by at least one electron
- Number of electrons in each orbital

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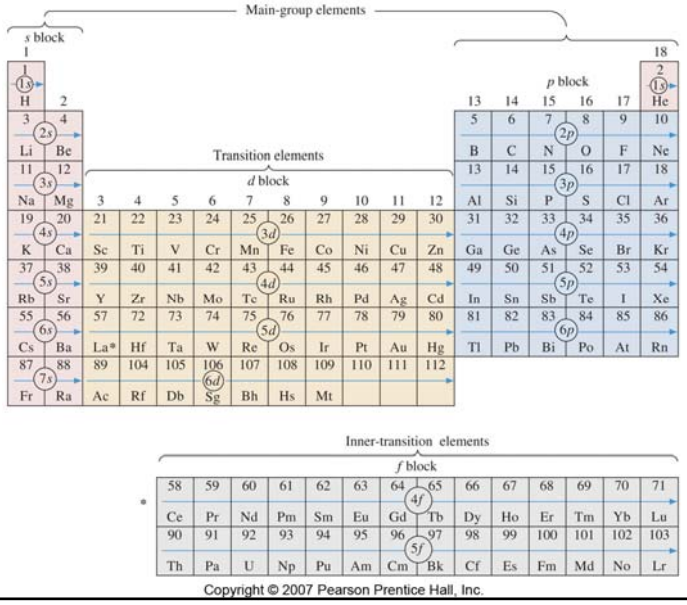
Electron Bookkeeping and Notation



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Electron Bookkeeping and Notation Examples

Carbon (6 e$^-$)

2p ____ ____ ____

2s ____

1s ____

Scandium (21 e$^-$)

3d ____ ____ ____ ____ ____

4s ____

3p ____ ____ ____

3s ____

2p ____ ____ ____

2s ____

1s ____

Spectroscopic Notation = ?

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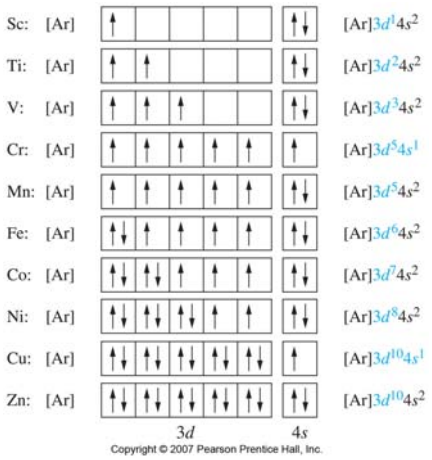
Misc. Notes on Electron Configuration

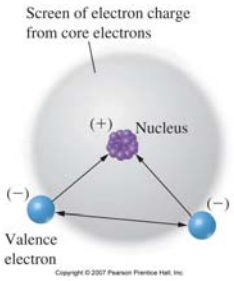
- Shortcuts using Noble Gas Configuration
- Core e- vs valence e-
- Dealing with ions
- Exceptions to the rule



Effective Nuclear Charge (Z*)

What does an electron "see" from its orbital? What keeps it in the orbital?

- *Penetration and Shielded Nuclear Charge:*



- Depends on:
  - Number of Protons (charge) in nucleus
  - Number of electrons in lower energy shells

Effective Nuclear Charge: Average shielded charge "felt" by an electron.

- The higher the Z*, the stronger the attraction for the nucleus.

$$Z_{eff} = Z - S$$

S = # of "shielding" electrons (core e^-)

More qualitative than quantitative

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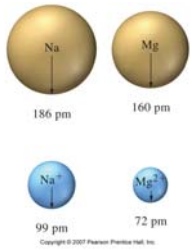
Periodic Trends

- **Atomic Radius:**



- **Ionic Radius:**

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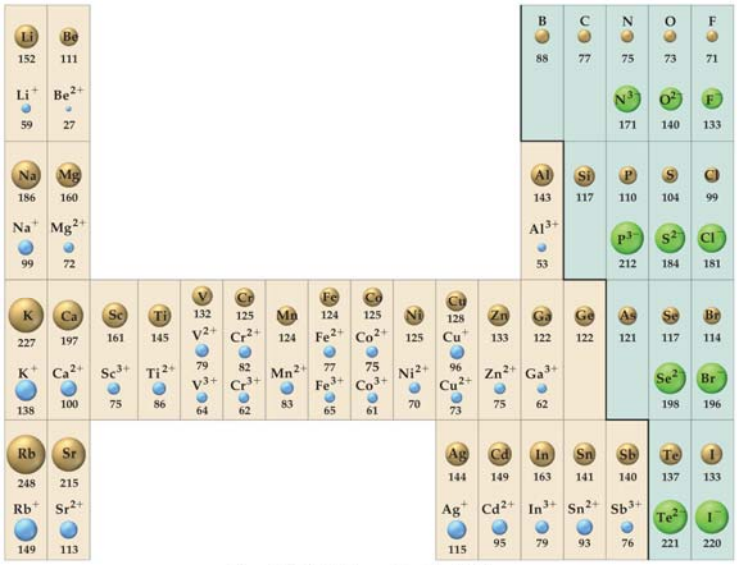
Periodic Trends



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Periodic Trends

- **Ionization Energy:**

$$X(g) \rightarrow X^+(g) + e^-$$

- **Electron Affinity:**

$$X(g) + e^- \rightarrow X^-(g)$$

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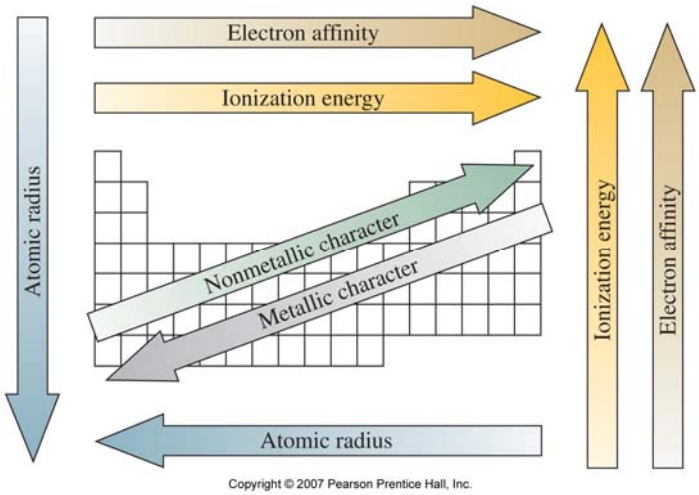
Periodic Trends



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