Announcements – 11/1/00

- **Prob Set Solutions**

  -missing pages from Ch. 13 solutions now available

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Beyond Hydrogen

- For atoms with more than one electron, we use the same orbitals (phew!) BUT their energies are not the same.
- Electron energies can be related to the *effective nuclear charge* (Z_{eff}) they experience in an orbital:

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

  -electrons in inner shells can *shield* the outer shell electrons from the full positive charge (***Z***) of the nucleus (***S = # of inner shell electrons***)

- In general: for a fixed value of ***n***, energies *increase* with increasing values of ***l***. (i.e., d > p > s)

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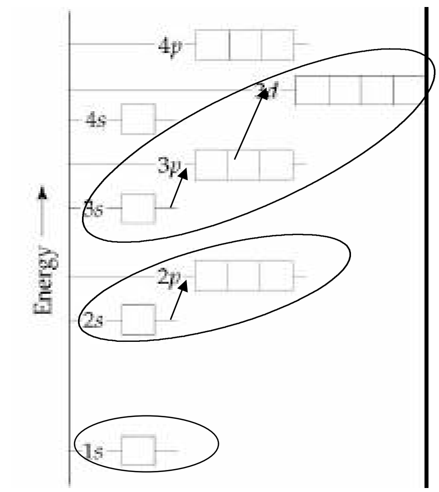
Many-electron Energy Levels

•Energies increase with increasing ***n***:

•Energies within a shell increase with increasing ***l*** of subshell:



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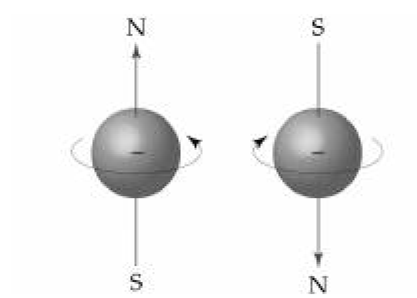
Electron Spin: The 4th Quantum Number

- Electrons also have a property called "spin":



•**The Electron Spin Quantum Number:**

$$m_s = +½ \text{ or } -½$$

-specifies a *specific electron* in an orbital

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Pauli Exclusion Principle

- **"No two electrons in an atom can possess the same set of four quantum numbers."**

  Implications:

  1. Since there are only **two** values of $\mathbf{m_s}$ possible for each orbital, *no more than two electrons (with opposite spins) can "occupy" an orbital*
  2. The four quantum numbers *uniquely identifies* an electron in an atom

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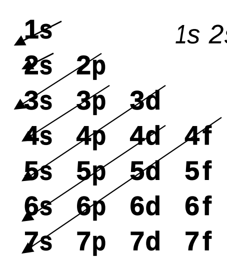
Electron Configurations

- How do we assign electrons in an atom to the various orbitals?

  1. **Fill orbitals so as to give a *minimum energy* for the atom**



1s 2s 2p 3s 3p ***4s*** *3d 4p* ***5s*** *4d 5p . . .*

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More Electron Config Rules

2. **Apply Pauli Exclusion Principle**
   - -only 2 electrons/orbital
   - -each with opposing spin

3. **Hund's Rule**
   - -fill *degenerate* orbitals so as to maximize the number of unpaired electrons *with the same spin:*

$3p^3$: ↑ ↑ ↑ ***NOT*** ↑↓ ↑ _

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The Aufbau Process

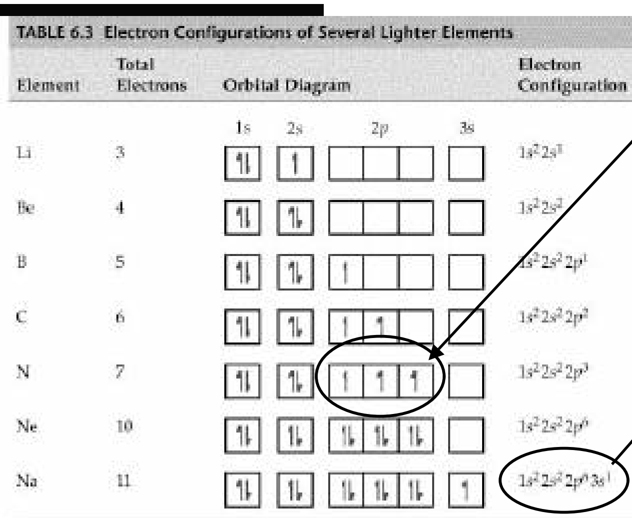

TABLE 6.3 Electron Configurations of Several Lighter Elements

Element	Total Electrons	Orbital Diagram 1s	2s	2p	3s	Electron Configuration
Li	3	↑↓	↑			$1s^2 2s^1$
Be	4	↑↓	↑↓			$1s^2 2s^2$
B	5	↑↓	↑↓	↑		$1s^2 2s^2 2p^1$
C	6	↑↓	↑↓	↑ ↑		$1s^2 2s^2 2p^2$
N	7	↑↓	↑↓	↑ ↑ ↑		$1s^2 2s^2 2p^3$
Ne	10	↑↓	↑↓	↑↓ ↑↓ ↑↓		$1s^2 2s^2 2p^6$
Na	11	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑	$1s^2 2s^2 2p^6 3s^1$

One electron in each p-orbital, before pairing

Use shorthand notation:

$[Ne]3s^1$

core e^- *valence e^-*

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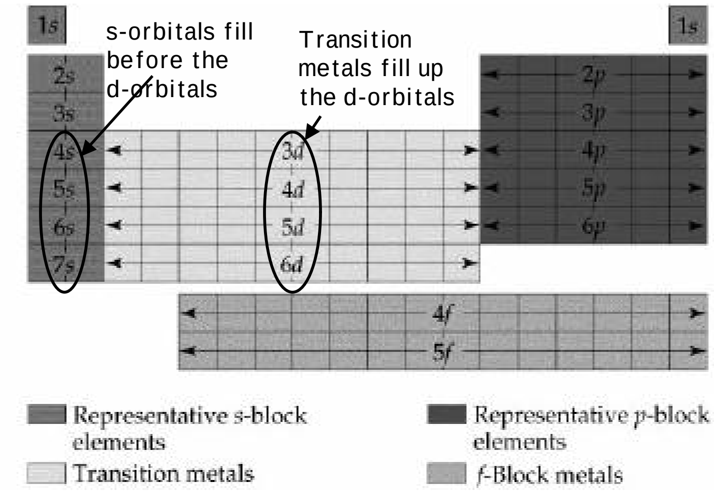
Electron Configurations and the Periodic Table



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Transition Metals

- Generally: *fill d-orbitals after filling* ***s-orbitals***

Sc: $[Ar]4s^2 3d^1$
Ti: $[Ar]4s^2 3d^2$

and on to

Zn: $[Ar]4s^2 3d^{10}$

- Of course, there are exceptions:

Cr: $[Ar]4s^1 3d^5$ - *not* $[Ar]4s^2 3d^4$
Cu: $[Ar]4s^1 3d^{10}$ - *not* $[Ar]4s^2 3d^9$

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Lanthanides and Actinides

- Generally: *fill **f-orbitals***

Lanthanides:

-fill f-orbitals *after* Lanthanum:

La [Xe]$6s^2$***$5d^1$*** ← one electron in d-orbital

Ce [Xe]$6s^2$***$4f^1$***$5d^1$ ← next, fill f-orbitals, until

Hf [Xe]$6s^2 4f^{14}$***$5d^2$*** ← resume filling d-orbitals

Actinides:

-fill f-orbitals *after* Actinium (do same as Lanthanides)

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Illustrative Example

- What is the *electron configuration* for **Se**?

Se: *34 electrons*

Se $1s^2\ 2s^2 2p^6\ 3s^2 3p^6\ 4s^2\ 3d^{10}\ 4p^4$

Se [Ar]$3d^{10}4s^2 4p^4$

4p	↑↓	↑	↑
n = 4, l = 1	p_x	p_y	p_z
m_s =	+½, -½	+½	+½

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