The periodic table:

1. Electron Configurations. written in terms of n, ℓ

1s, 2s, 2p, 3s, 3p, [4s, 3d], 4p, [5s, 4d], 5p, [6s, 4f, 5d], 6p, [7s, 5f, 6d]

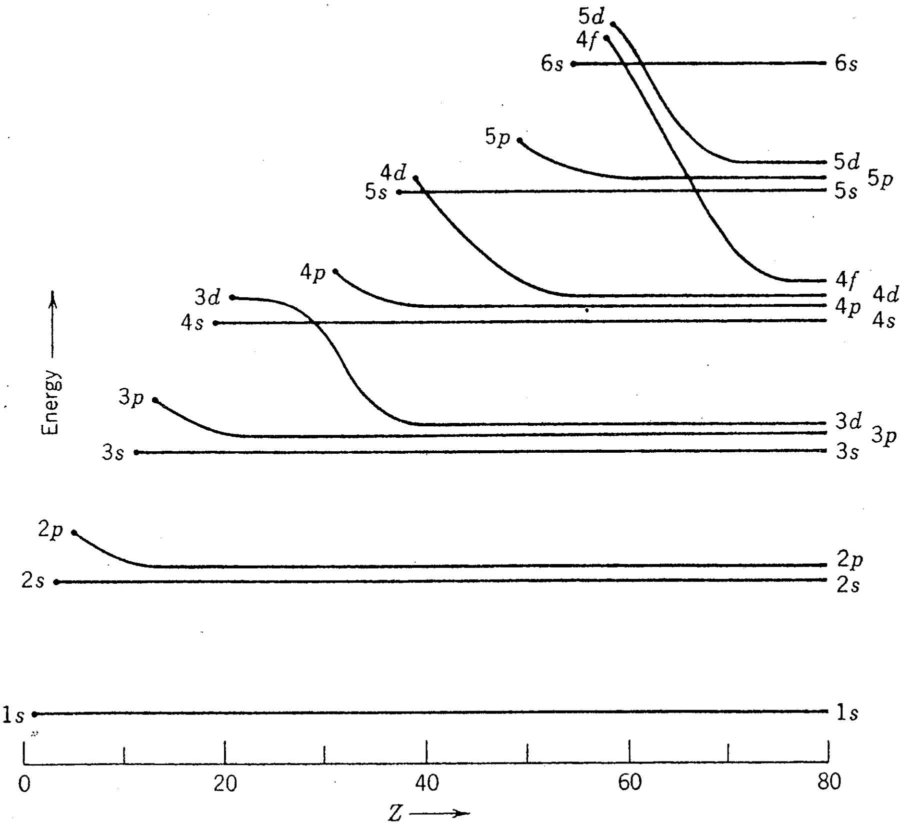

Figure 9-14 A schematic representation of the energy ordering of all the subshells in an atom, as a function of its atomic number Z. Each curve begins at the Z for which the subshell begins to be occupied. Only subshells occupied in atoms through mercury are shown, so all curves stop at $Z = 80$. The ordering of the outer filled subshells in various atoms is found on the left side of the diagram. The ordering of all filled subshells in mercury is found on the right side of the diagram. The energy scale is non-linear and, furthermore, varies with Z.

Table 9-2 The Energy Ordering of the Outer Filled Subshells

Quantum Numbers n, l	Designation of Subshell	Capacity of Subshell $2(2l + 1)$	
—	—	—	
—	—	—	
6, 2	$6d$	10	
5, 3	$5f$	14	
7, 0	$7s$	2	
6, 1	$6p$	6	
5, 2	$5d$	10	
4, 3	$4f$	14	
6, 0	$6s$	2	
5, 1	$5p$	6	↑
4, 2	$4d$	10	Increasing energy (less negative)
5, 0	$5s$	2	
4, 1	$4p$	6	
3, 2	$3d$	10	
4, 0	$4s$	2	
3, 1	$3p$	6	
3, 0	$3s$	2	
2, 1	$2p$	6	
2, 0	$2s$	2	
1, 0	$1s$	2	← Lowest energy (most negative)

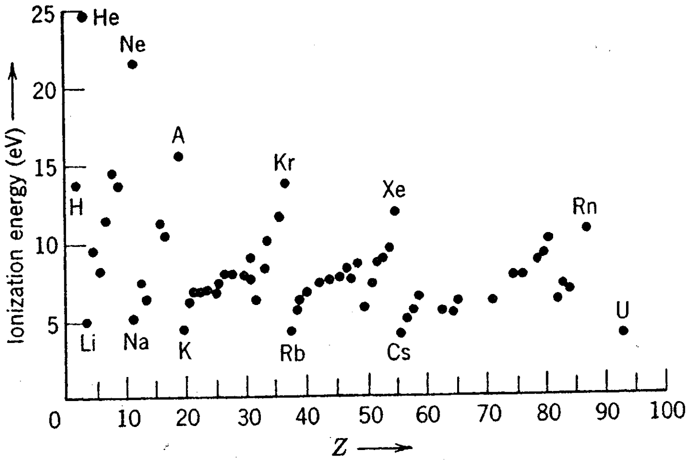

Figure 9-15 The measured ionization energies of the elements.

Carbon: $1s^2 \; 2s^2 \; 2p^2$

1.) 2 L=1 electrons in the unfilled subshell.

2.) $L=1 \otimes L=1 = |2\rangle_L + |1\rangle_L + |0\rangle_L$

3.) $M_L \Rightarrow$
$m_\ell = 1$
$m_\ell = 0 \;\; \uparrow$
$m_\ell = -1 \;\; \uparrow$

4.) What are the possible spectroscopic states? $^{2s+1}L_J$

S \ L	0	1	2
1	$^3S_1$	$^3P_{2,1,0}$	$^3D_{3,2,1}$
0	$^1S_0$	$^1P_1$	$^1D_2$

Answer $\Rightarrow$ $^3S_1 \quad ^1S_0 \quad ^3P_2 \quad ^3P_1 \quad ^3P_0 \quad ^1P_1 \quad ^3D_3 \quad ^3D_2 \quad ^3D_1 \quad ^1D_2$

Hund's Rules

**Problem 5.13

(a) **Hund's first rule** says that, consistent with the Pauli principle, the state with the highest total spin (S) will have the lowest energy. What would this predict in the case of the excited states of helium?

(b) **Hund's second rule** says that, for a given spin, the state with the highest total orbital angular momentum (L), consistent with overall antisymmetrization, will have the lowest energy. Why doesn't carbon have $L = 2$? *Hint:* Note that the "top of the ladder" ($M_L = L$) is symmetric.

(c) **Hund's third rule** says that if a subshell (n, l) is no more than half filled, then the lowest energy level has $J = |L - S|$; if it is more than half filled, then $J = L + S$ has the lowest energy. Use this to resolve the boron ambiguity in Problem 5.12(b).

(d) Use Hund's rules, together with the fact that a symmetric spin state must go with an antisymmetric position state (and vice versa) to resolve the carbon and nitrogen ambiguities in Problem 5.12(b). *Hint:* Always go to the "top of the ladder" to figure out the symmetry of a state.

In class
→ Problem 5.14 The ground state of dysprosium (element 66, in the 6th row of the Periodic Table) is listed as ${}^5I_8$. What are the total spin, total orbital, and grand total angular momentum quantum numbers? Suggest a likely electron configuration for dysprosium.

From Griffiths 5.12 (b)

Homework: Find all the spectroscopic states for the valence electrons in Nitrogen $(He)(2s)^2(2p)^3$

S \ L	0	1	2	3
3/2	${}^4S_{3/2}$	${}^4P_{\frac{5}{2}\,\frac{3}{2}\,\frac{1}{2}}$	${}^4D_{\frac{7}{2}\,\frac{5}{2}\,\frac{3}{2}\,\frac{1}{2}}$	${}^4F_{9/2\,\frac{7}{2}\,\frac{5}{2}\,\frac{3}{2}}$
1/2	${}^2S_{1/2}$	${}^2P_{\frac{3}{2}\,\frac{1}{2}}$	${}^2D_{\frac{5}{2}\,\frac{3}{2}}$	${}^2F_{\frac{7}{2}\,\frac{5}{2}}$