Chapter 6 Part 3

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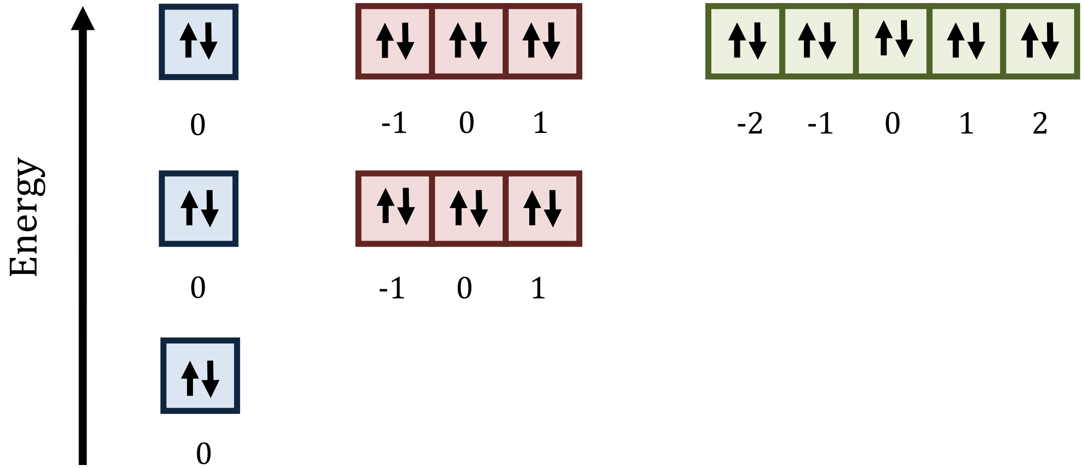
Orbital Energy Diagrams

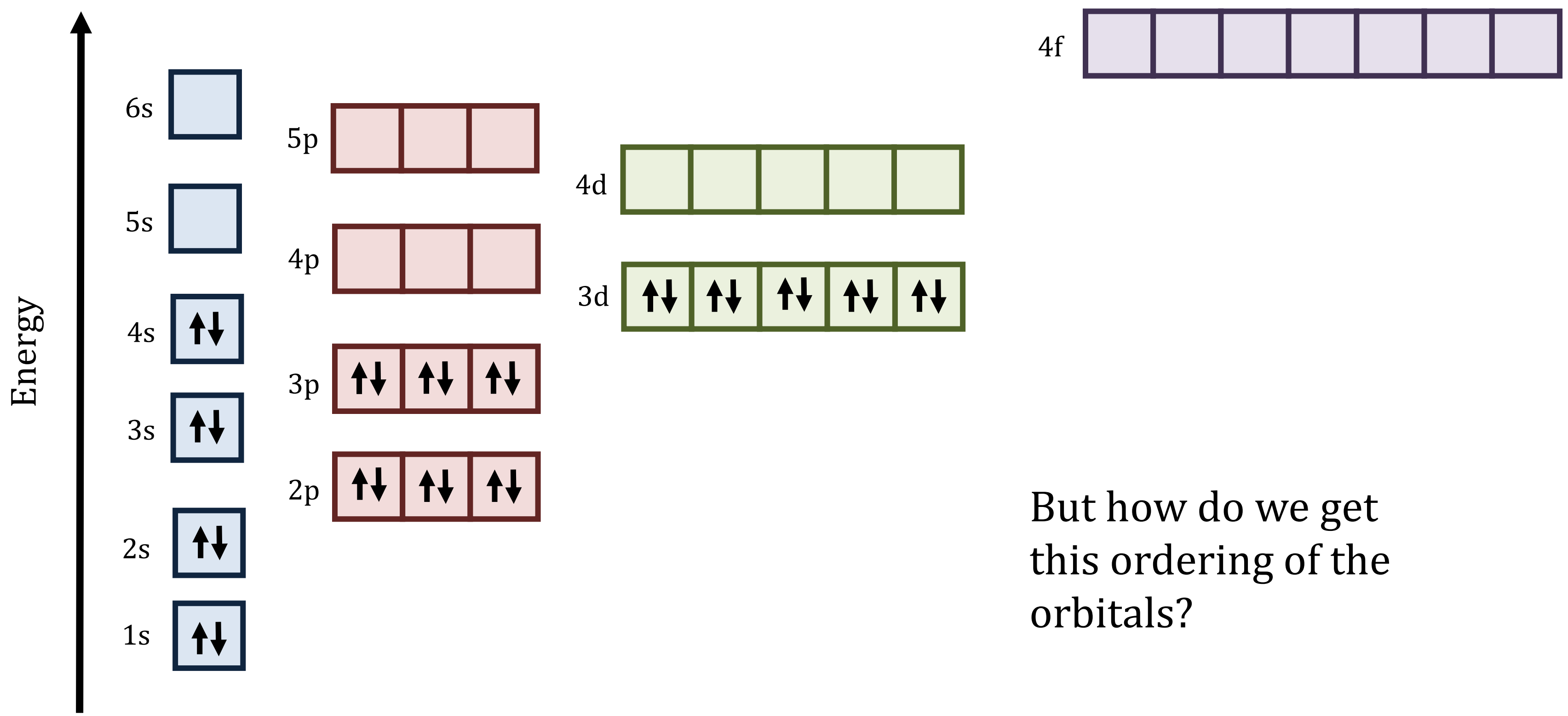
Which sets of orbitals would you expect to have the same energy?



Diagrams for multielectron systems

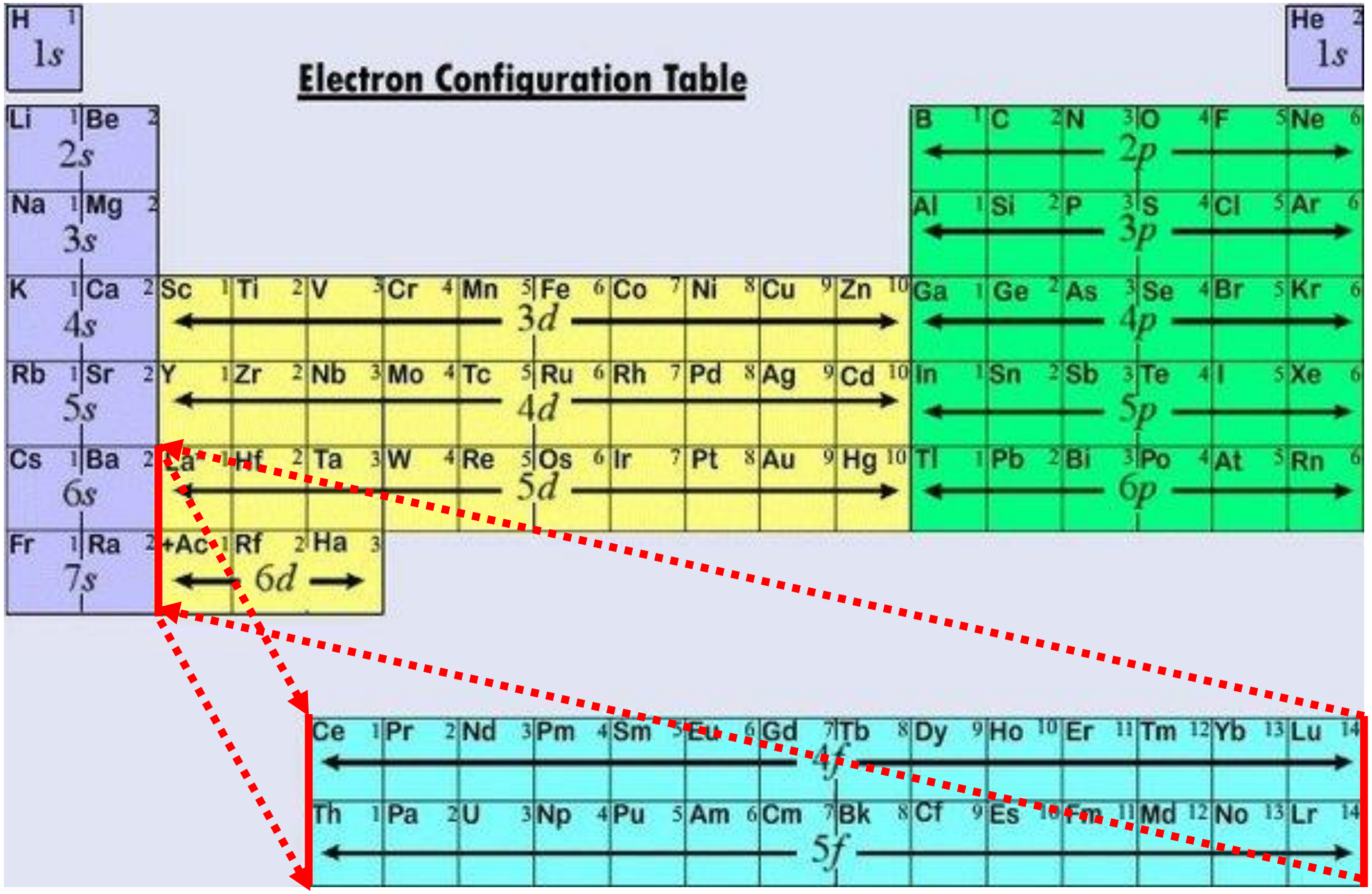
- For multielectron systems, the energy depends on the quantum numbers n and l
- As n increases, orbital energy increases for orbitals of the same type
  - 4s > 3s > 2s > 1s
  - 4p > 3p > 2p
- As l increases within a shell, orbital energy increases
  - Ex. 4f > 4d > 4p > 4s
- As n increases, the subshell energies become closely spaced and overlapping occurs

Diagrams for multielectron systems



But how do we get this ordering of the orbitals?

Electron configurations and the periodic table



Comparison to our Periodic Table

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** 1.008																	2 **He** 4.0026
3 **Li** 6.94	4 **Be** 9.0122											5 **B** 10.81	6 **C** 12.011	7 **N** 14.007	8 **O** 15.999	9 **F** 18.998	10 **Ne** 20.180
11 **Na** 22.990	12 **Mg** 24.305											13 **Al** 26.982	14 **Si** 28.085	15 **P** 30.974	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.948
19 **K** 39.098	20 **Ca** 40.078	21 **Sc** 44.956	22 **Ti** 47.867	23 **V** 50.942	24 **Cr** 51.996	25 **Mn** 54.938	26 **Fe** 55.845	27 **Co** 58.933	28 **Ni** 58.693	29 **Cu** 63.546	30 **Zn** 65.38	31 **Ga** 69.723	32 **Ge** 72.630	33 **As** 74.922	34 **Se** 78.97	35 **Br** 79.904	36 **Kr** 83.798
37 **Rb** 85.468	38 **Sr** 87.62	39 **Y** 88.906	40 **Zr** 91.224	41 **Nb** 92.906	42 **Mo** 95.95	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.76	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57-71 *	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.84	75 **Re** 186.21	76 **Os** 190.23	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** (226)	89-103 #	104 **Rf** (265)	105 **Db** (268)	106 **Sg** (271)	107 **Bh** (270)	108 **Hs** (277)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (280)	112 **Cn** (285)	113 **Nh** (286)	114 **Fl** (289)	115 **Mc** (289)	116 **Lv** (293)	117 **Ts** (294)	118 **Og** (294)

* Lanthanide series	57 **La** 138.91	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 151.96	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.05	71 **Lu** 174.97
# Actinide series	89 **Ac** (227)	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

Aufbau Principle

- Says that electrons fill the lowest available energy levels before filling higher levels
- By filling orbitals of lowest energy first, you usually get the lowest total energy (ground state) of the atom

spdf notation

- When using *spdf* notation, fill the orbitals using the previously stated rules.
- A superscript is used to show how many electrons are in an orbital
  - Ex. Li: $1s^2 2s^1$ & F: $1s^2 2s^2 2p^5$

Noble gas notation

- Noble gas notation is often used to represent filled shells
- In noble gas configuration, write the symbol of the most recent noble gas in brackets, and then list the remaining electrons as previously done in *spdf* notation
  - Ex. Si: [Ne]$3s^2 3p^2$ & In [Kr]$5s^2 4d^{10} 5p^1$
- The electron configuration for chromium is [Ar]$4s^1 3d^5$, and the electron configuration for copper is [Ar]$4s^1 3d^{10}$. These are irregular and should be memorized.

Electron configurations

Give the electron configurations for the following atoms using *spdf* notation and noble gas notation.

A. Sulfur
B. Cobalt
C. Californium (noble gas notation only)

Electron configurations

K is in the 1st column of the periodic table, and the 4th row. The electron configuration of K ends in

A. $1s^4$
B. $4s^1$
C. $1p^4$
D. $4p^1$
E. $4d^1$

Excited State Electron Configurations

Which of the following is not an excited state electron configuration of carbon?

A. $1s^2 2s^2 2p^2$
B. $1s^2 2s^2 2p^1 3s^1$
C. $1s^2 2s^2 3s^2$
D. $1s^2 2s^1 2p^3$

Adding electrons to form anions

- Electrons are filled in according to the energy ordering of the orbitals and Hund's rule.

Ion Formation

Give the electron configuration of a N^{2-} ion.

Ion Formation

What is the electron configuration for the P^{3-} ion?

A. $1s^2 2s^2 2p^6 3s^2$
B. $1s^2 2s^2 2p^6 3s^2 3p^0$
C. $1s^2 2s^2 2p^6 3s^2 3p^3$
D. $1s^2 2s^2 2p^6 3s^2 3p^6$

Loss of Electrons

- Electrons are removed corresponding to the shells with the greatest n value because electrons with greater values are farther from the positive pull of the nucleus

Ion Formation

From which orbital would electrons be lost from Zr in the formation of formation of Zr^{2+}?

A. 4p
B. 5s
C. 4d
D. 5p

Ion Formation

What is the noble gas electron configuration of Zr^{2+}?

Hund’s rule of maximum multiplicity

- Hund’s rule states that the lowest-energy electron configuration is the one where the maximum number of electrons is unpaired.

Orbital Box Notation

- In orbital box notation, boxes are used to represent orbitals and arrows are used to represent electrons
- In comparison to the *spdf* method, this one places more emphasis on the spin of the atom
- The electron configuration of nitrogen using orbital box notation is

1s	2s	2p		
$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow$	$\uparrow$	$\uparrow$

- The electron configuration of fluorine using orbital box notation is

1s	2s	2p		
$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow$

Orbital Box Notation

Write the electron configuration of silicon using orbital box notation.