

ALE 5. Ground-State Electron Configurations

(Reference: Sections 8.1 – 8.3 - *Silberberg 5th edition*)

Where are the electrons in an atom in its lowest electronic state?

The Model: Orbital Designations

The designation of an **orbital** is a number representing the **energy level (or shell)** (given by the **principal quantum number, n**) immediately followed by the lowercase letter (*s, p, d, or f*) which represents the **sublevel (or subshell)** (given by the **orbital angular momentum quantum number, ℓ**).

Examples:

$1s$ = the orbital in the 1st shell (which is required to have $\ell = 0$)

$3d$ = an orbital in the 3rd shell having $\ell = 2$

$4f$ = an orbital in the 4th shell having $\ell = 3$

Key Questions

- Provide the orbital designation for each of the following:
 - an orbital in the 2nd shell having $\ell = 1$
 - the orbital in the 5th shell having $\ell = 0$
- Provide a description of the orbital (what are n and ℓ equal to?) for each of the following:
 - $6p$
 - $4d$

The Model: The Number of Orbitals Within a Subshell

Within a subshell, the **magnetic quantum number, m_ℓ** , differentiates orbitals that have the same shape. For example, when $\ell = 1$, m_ℓ may be equal to either +1, or 0, or -1. This tells us that within a *p* subshell there are three *p* orbitals, which differ from each other due to their orientation in 3-D space (one points in the *x*-direction, a second in the *y*-direction, and the last in the *z*-direction).

Key Questions

- What are the allowed values of m_ℓ for:
 - $\ell = 0$?
 - $\ell = 2$?
 - $\ell = 3$?

4. How many orbitals are there within:
- an s subshell?
 - a d subshell?
 - an f subshell?

The Model: The Number of Electrons in an Orbital

The electron is what is referred to as a “spin- $\frac{1}{2}$ particle”. The electron may be thought of as “spinning like a top”. If it is spinning in one direction, we call it “spin up” and it is assigned a **spin quantum number, m_s** , of $+\frac{1}{2}$. If it is spinning in the other direction, we call it “spin down” and it is assigned $m_s = -\frac{1}{2}$.

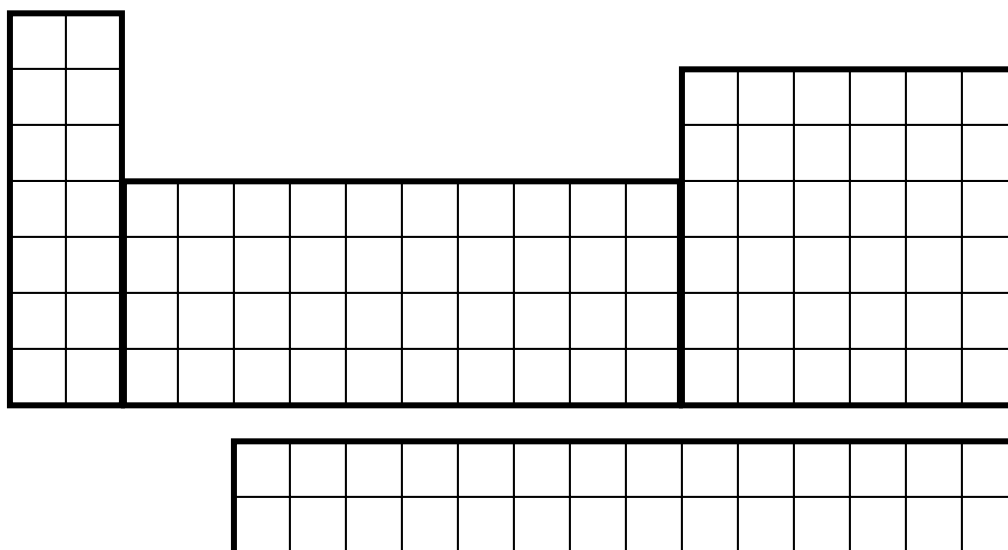
The **Pauli Exclusion Principle** dictates that two electrons within the same orbital must have different values of m_s . Therefore, there may be 0, or 1, or 2 electrons in an orbital, ***BUT NO MORE THAN TWO ELECTRONS MAY EVER OCCUPY ANY ORBITAL.***

Key Questions

5. How many electrons can occupy:
- an s subshell?
 - a p subshell?
 - a d subshell?
 - an f subshell?

The Model: “Blocks” of the Periodic Table

If the “square” on the Periodic Table containing helium (He) were shifted over to Group IIA (the Alkaline Earth Metal family), the Periodic Table may be sketched as:



Each “square” on the Periodic Table represents one combination of the four quantum numbers (n , ℓ , m_ℓ , and m_s). The “square” in the upper left of the Periodic Table (the square containing hydrogen) is our “starting point” in that it has the combination of n , ℓ , m_ℓ , and m_s that minimizes the energy of our electron. As we move to “squares” corresponding to larger atomic numbers, the combination of n , ℓ , m_ℓ , and m_s changes so that the trend is that electronic energy increases.

The four outlined rectangular regions are referred to as the s block, the p block, the d block, and the f block. Two (or more) “squares” within the same period are in the same subshell if they are also in the same block. Since two electrons can occupy an orbital, two “squares” within the same block represent one orbital within a subshell.

Key Questions

6. Use your answers to Questions 4 and 5 to determine which outlined region of the [periodic table on page 2](#) is the s block, which is the p block, which is the d block, and which is the f block. Label the representation of the Periodic Table [on page 2](#) with the labels “ s block”, “ p block”, “ d block”, and “ f block”.
- 7a. What is the minimum value of n needed for the shell to have an s subshell? Label the representation of the Periodic Table with that number in the uppermost “square” of the s block. Label the square underneath it with the next counting number, and continue labeling in this fashion until you get to the bottom row of the block.
- b. What is the minimum value of n needed for the shell to have a p subshell? Label the representation of the Periodic Table with that number in the uppermost “square” of the p block. Label the square underneath it with the next counting number, and continue labeling in this fashion until you get to the bottom row of the block.
- c. What is the minimum value of n needed for the shell to have a d subshell? Label the representation of the Periodic Table with that number in the uppermost “square” of the d block. Label the square underneath it with the next counting number, and continue labeling in this fashion until you get to the bottom row of the block.
- d. What is the minimum value of n needed for the shell to have an f subshell? Label the representation of the Periodic Table with that number in the uppermost “square” of the f block. Label the square underneath it with the next counting number, and continue labeling in this fashion until you get to the bottom row of the block.

The Model: Ground State Electron Configurations

An **electron configuration** is written to inform the reader in which orbitals the electrons may be found within an atom. The **ground state** is the lowest energy state of an atom. So the ground state electron configuration simply tells the reader what orbitals are occupied when the atom has a minimum electronic energy.

For example, hydrogen in the ground state has a single electron in the $1s$ orbital. So the ground state electron configuration of hydrogen is: $1s^1$. The superscript “1” tells the reader how many electrons are in the $1s$ subshell.

Exercises

A. Use the Periodic Table as a guide to write *full* ground state electron configurations for:

- ① Fluorine
- ② Iron
- ③ Rubidium
- ④ Tin
- ⑤ Europium

The Model: Electron Configurations of Ions

Once the ground state electron configuration of a neutral atom is determined, it is easy to predict what the ground state electron configuration of an ion of the atom will be.

If the ion is an anion: Until the number of added electrons reflects the magnitude of the negative charge, electrons are added to the very next orbital in order to minimize the energy.

If the ion is a cation: Until the number of removed electrons reflects the magnitude of the positive charge, electrons are removed from the orbitals of highest energy.

EXCEPTION: If the atom is a transition metal, then the first electrons to be removed come from the *s* subshell of highest principal quantum number. If the atom is a main-group heavy metal (*e.g.*, Lead), then the first electrons to be removed come from the *p* subshell of highest value of *n* and the next electrons to be removed come from the *s* subshell of highest value of *n*.

Example 1: Sulfide ion (anion)

The ground state electron configuration of Sulfur is: $1s^2 2s^2 2p^6 3s^2 3p^4$. The Sulfide ion is S^{2-} , which means it has two more electrons than a Sulfur atom. Since a *p* subshell can accommodate up to six electrons, the ground state electron configuration of Sulfide is: $1s^2 2s^2 2p^6 3s^2 3p^6$.

Example 2: Potassium ion (cation)

The ground state electron configuration of Potassium is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$. The Potassium ion is K^+ , which means it has one less electron than a Potassium atom. So the ground state electron configuration of a Potassium cation is: $1s^2 2s^2 2p^6 3s^2 3p^6$.

Example 3: Iron(II) ion (cation)

The ground state electron configuration of Iron is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$. The Iron(II) ion is Fe^{2+} , which means it has two less electrons than an Iron atom. Since Iron is a transition metal, the ground state electron configuration of Iron(II) is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^0 3d^6$.

Exercises

B. Write full ground state electron configurations of the following ions:

① Bromide ion

② Nitride ion

③ Magnesium ion

④ Chromium(III) ion

⑤ Tin(IV) ion