Guided Practice Activities

**Module:** Atomic Theory

**Section:** Electron Configurations of Many Electron Atoms - Key
Activity 1

The purpose of this activity is to recognize that while the Schrödinger equation accurately solves for the energies and orbitals for a one-electron system such as hydrogen, several issues arise when solving for atoms containing more than one electron. Check your comprehension of the consequences of multi-electron systems by working through the following problems.

1. Explain the concept of the “loss of degeneracy within energy levels.”

   The loss of degeneracy within energy levels occurs for multi-electron atoms. This means that the 2s orbital will not share the same energy as the 2p orbitals even though they are both in the level 2 shell. The three 2p orbitals, however, each share the same energy. As a result, the electrons will reside in the lowest energy orbitals and “build up” from the lowest energy orbitals to the higher energy orbitals (meaning an atom in its ground state will have electrons that occupy the lowest energy orbitals possible).

2. The Pauli exclusion principle states:

   The Pauli exclusion principle states that no two electrons of the same atom can share the same set of quantum numbers.

3. Because of the Pauli exclusion principle and the fact that each orbital can accommodate up to two electrons, a fourth quantum number must be included when fully describing the electron configuration of an electron in a multi-electron atom.

4. Define ms and state its allowed values.

   ms is the “electron spin” quantum number. This is the fourth quantum number that is required by the Pauli Exclusion principle. It can have the value of −½ or +½.

5. Hund’s rule states:

   Hund’s rule states that electrons must be placed into separate degenerate orbitals before they are placed in the same degenerate orbital. For example, there are three 2p orbitals. In order to fill these orbitals with four electrons (say for the oxygen atom), first we place a single atom into each of the three orbitals before we pair up the fourth electron into an orbital.

6. The Aufbau principle states:

   Aufbau’s principle states that the electron configuration of an atom can be determined by “building up” electrons from the lowest energy orbital to the highest energy orbital. So every atoms will begin with an electron placed in the “1s” orbital. From there, we can place
the rest of the atom’s electrons into orbitals by ascending energy level until we have run out of electrons. (caveat: for excited atoms we would need to move valence electrons from their ground state energy levels to higher ones).

Activity 2

1. One can use the periodic table to help determine the order of filling for the ground state electron configurations of all the elements. Label the s, p, d and f blocks on the periodic table to the right. Explain how to use the periodic table to write electron configurations.

By moving across the periods, we can count up and classify the number of electrons in a given atom. The color blocks tell us which type of orbital a particular electron resides in while the period number tells us the energy level.

For electrons in d orbitals, we subtract “1” from the period number to find the energy level. For electrons in f orbitals, we subtract “2” from the period number to find the energy level.
2. The orbital notation of the electron configuration of selenium is $1s^22s^22p^63s^23p^64s^23d^{10}4p^4$. Without looking at the periodic table, we know that selenium is in the fourth row based on the energy level (principle quantum number) value of its highest energy electrons.

Ground State Electron Configurations

Activity 1

The purpose of the following activity is to check your understanding of writing electron configurations of elements and using multiple methods, relating the electron configuration to the orbital notation and to recognize and explain the exceptions to the Aufbau principle, and to think about the electron configurations in the context of where the electrons are in an atom.

1. Write the ground state electron configuration for calcium.

   \[1s^22s^22p^63s^23p^64s^2\]

2. Write the ground state electron configuration for calcium using the noble gas shorthand method.

   \[[\text{Ar}]4s^2\]

3. Write the ground state electron configuration for calcium using the orbital method in which you show the spins on the electrons using arrows.

   \[\begin{array}{cccc}
   \uparrow & \uparrow & \uparrow & \uparrow \\
   1s & 2s & 2p & 3s \\
   \hline
   \uparrow & \uparrow & \uparrow & \uparrow \\
   3p & 4s & \\
   \end{array}\]

4. State all the possible quantum numbers for an electron found in the 2p subshell of calcium.

   \[n = 2\text{ (second energy level)}\]
   \[l = 1\text{ (in the p orbital)}\]
   \[m_l = -1, 0, 1\text{ (}m_l\text{ can range from }[-1,0,1]\text{ so there are three possibilities)}\]
   \[m_s = -\frac{1}{2}, +\frac{1}{2}\text{ (}m_s\text{ is either }-\frac{1}{2}\text{ or }+\frac{1}{2}\text{)}\]
There are six unique sets of quantum numbers available to describe an electron found in a 2p orbital of Calcium:

\[(2, 1, -1, -\frac{1}{2}) ; (2, 1, -1, +\frac{1}{2}) ; (2, 1, 0, -\frac{1}{2}) ; (2, 1, 0, +\frac{1}{2}) ; (2, 1, 1, -\frac{1}{2}) ; (2, 1, 1, +\frac{1}{2})\]

5. State all the possible quantum numbers for an electron found in a 2p orbital of Ca. Explain why this list is the same of different from the list you wrote in problem #4.

\[
\begin{align*}
n &= 2 \text{ (second energy level)} \\
\ell &= 1 \text{ (in the d orbital)} \\
m_\ell &= -1, 0, 1 \text{ (}m_\ell\text{ can range from} -\ell \text{ to} +\ell \text{ so there are three possibilities)} \\
m_\text{s} &= -\frac{1}{2}, +\frac{1}{2} \text{ (}m_\text{s}\text{ is either} -\frac{1}{2} \text{ or} +\frac{1}{2})
\end{align*}
\]

There are six unique sets of quantum numbers available to describe an electron found in a 2p orbital of Calcium:

\[(2, 1, -1, -\frac{1}{2}) ; (2, 1, -1, +\frac{1}{2}) \text{ OR} \]
\[(2, 1, 0, -\frac{1}{2}) ; (2, 1, 0, +\frac{1}{2}) \text{ OR} \]
\[(2, 1, 1, -\frac{1}{2}) ; (2, 1, 1, +\frac{1}{2})\]

There are six possible sets of quantum numbers in the 2p subshell. These sets are identical for electrons in a 2p orbital EXCEPT that a particular 2p orbital has only one possible \(m_\ell\) value. So if we are just looking at electrons in a single 2p orbital there are only two choices from quantum number sets. The question didn’t define which 2p orbital we are in though, so we still had to list all six quantum number sets, but notice how the list was broken differently. Once we identified the \(m_\ell\) value of the 2p orbital we are focusing on, then we only have two choices for quantum number sets.

For example, if we were looking at the particular 2p orbital with the \(m_\ell\) value of 1, then an electron in that particular 2p orbital could have the quantum numbers of either \((2, 1, 1, -\frac{1}{2})\) or \((2, 1, 1, +\frac{1}{2})\).
**Activity 2**

1. Write the ground state electron configuration for gallium.
   
   $1s^22s^22p^63s^23p^64s^23d^{10}4p^1$

2. Write the ground state electron configuration for gallium using the noble gas shorthand method.
   
   $[Ar]4s^23d^{10}4p^1$

3. Write the ground state electron configuration for gallium using the orbital method in which you show the spins on the electrons using arrows.

4. State all the possible quantum numbers for an electron found in the 3d subshell of gallium.
   
   $n = 3$ (third energy level)
   
   $\ell = 2$ (in the d orbital)
   
   $m_\ell = -2, -1, 0, 1, 2$ ($m_\ell$ can range from $-\ell$ to $+\ell$ so there are five possibilities)
   
   $m_s = -\frac{1}{2}, +\frac{1}{2}$ ($m_s$ is either $-\frac{1}{2}$ or $+\frac{1}{2}$)

   There are ten unique sets of quantum numbers available to describe an electron found in a 3d orbital of Gallium:

   $$(3, 2, -2, -\frac{1}{2})$; (3, 2, -2, $+\frac{1}{2}$); (3, 2, -1, $-\frac{1}{2}$); (3, 2, -1, $+\frac{1}{2}$); (3, 2, 0, $-\frac{1}{2}$); (3, 2, 0, $+\frac{1}{2}$); (3, 2, 1, $-\frac{1}{2}$); (3, 2, 1, $+\frac{1}{2}$); (3, 2, 2, $-\frac{1}{2}$); (3, 2, 2, $+\frac{1}{2}$)

**Activity 3**

1. There are some exceptions to the ground state electron configurations predicted by this periodic table filling method. Those exceptions occur when:

   The exceptions generally occur when an atom has an almost half-filled or almost completely filled “d” or “f” subshell. For example, a half-filled “d” subshell is much more stable than a “d” subshell with only 4 electrons. So the “d” subshell will borrow a “s” electron from the nearest “s” subshell.

   The same process would occur if the “d” subshell had only 9 electrons instead of the full 10 electrons.
2. Write the ground state electron configuration for chromium.

\[ 1s^22s^22p^63s^23p^64s^13d^5 \]

3. Write the ground state electron configuration for chromium using the noble gas shorthand method.

\[ [\text{Ar}]4s^13d^5 \]

4. Write the ground state electron configuration for chromium using the orbital method in which you show the spins on the electrons using arrows.

\[
\begin{align*}
1s & \quad 2s \\
& \quad 2p \\
& \quad 3s \\
& \quad 3p \\
& \quad 4s \\
& \quad 3d
\end{align*}
\]

**Activity 4**

1. The most important electrons for the chemists are the electrons on the surface of the atom. These electrons are in the highest energy level and are the electrons that are involved in bonding.

Electrons on the surface of the atom are called **valence** electrons.

Electrons below the surface of the atom are called **core** electrons.

2. Explain the shell model of the atom in the context of the location of the electrons.

Sometimes it's helpful to think of an atom as a nucleus surrounded by different shells of electron density. Almost like a nesting doll, the shells increase in size as we look farther away from the nucleus. In this way the outmost shell engulfs all the inner shells and the nucleus sits down in the center of it all. So valence electrons are the electrons in the shell farther out from the nucleus, the electrons in the inner shells make up the "core” of the atom.

For example, silicon has a total of 14 electrons: \( 1s^22s^22p^63s^23p^2 \)

Its valence electrons are in the 3rd energy level and there are a total of four valence electrons: two in the “3s” orbital and two in the “3p” orbital. As we can see in the diagram below, the third energy level is larger than both the first (lowest; most inward) and the second (middle).