Learning Guide Module

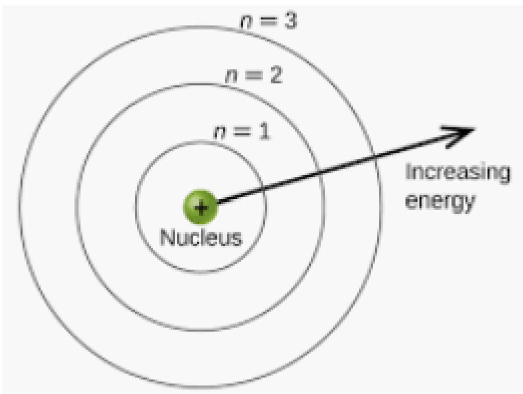
Subject Code	Chem 1	General Inorganic Chemistry
Module Code	3.0	*Electronic Structure of Atoms and the Periodic Properties of Elements*
Lesson Code	3.2	*Quantum Numbers*
Time Limit		30 mins.

Components	**Tasks**	TA[a]	ATA[b]
Target 	By the end of this module, the students will have been able to 1. Apply Hund's Rule, Aufbau Principle and Pauli's Exclusion Principle to the distribution of electrons around the nucleus. 2. Use the electronic configuration to determine the position of an element in the periodic table. 3. Relate the electron configuration with the 4 quantum numbers.	1 min.	
Hook 	Have you ever watched fireworks and wondered what made the colors, what made them crackle, and fly? Well, I wondered about that too when I still don't have knowledge in chemistry.  Source: Largest Firework Display. Retrieved from guinessworldbookrecord.com **Figure 1.** Fireworks Display in the Philippines Electrons are actually the star of fireworks. Electrons in the atoms consisting the fireworks are excited to a higher energy level by application of heat. However, an electron can only remain in a higher-energy state for a very short time. It emits light when it returns to a lower energy and that makes fireworks glow.	2 mins.	

<table>
<tr>
<td></td>
<td>In the previous lesson, you have already learned the different theories that describe the atomic structure. In this module, we will focus on one of the most revolutionary discoveries of the twentieth century—the quantum theory, which would give us an understanding of how electrons behave in atoms.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ignite
</td>
<td>
We will tackle first a very important term in quantum theory, quantum numbers. These numbers are used to describe atomic orbitals, which are regions in an atom where the electrons are most likely to be found. It is also used to characterize electrons that reside in them. It is the same as describing your permanent address. You start with the most specific (Block and Lot No.) to general (Region No.) e.g. I reside at <u>Blk. 6 Lot 3, Purok Agan Homes, Koronadal City, Region 12</u>. In the case of electrons, three quantum numbers are generally required to describe the distribution or location of electrons in atoms. These are the principal quantum number (n), the angular momentum quantum number (ℓ), and the magnetic quantum number (m_ℓ). A fourth quantum number, the spin quantum number (m_s), completes the description of electrons in atoms.

Let’s learn more about these quantum numbers.

1. Principal quantum number (n) is a positive integer (1, 2, 3, and so forth). It indicates the relative size of the orbital and so the relative distance from the nucleus. The principal quantum number specifies the energy level of an electron in atoms: the higher the n value, the higher the energy level and the farther it is from the nucleus. Note that when the electron of H atom, for example, occupies an orbital with n =1, the H atom is in its ground state and has lower energy than when the electron occupies the n = 2 orbital (first excited state) (Silberberg, 2009).



Source: Development of Quantum Theory. Retrieved from chemlibretexts.org
</td>
<td>20

mins.</td>
<td></td>
</tr>
</table>

Figure 2. The larger the value of principal quantum number, n, the further away the electron is from the nucleus and the higher its energy level

2. **Angular momentum quantum number, (*ℓ*)** is a quantum number that can only have integral values from 0 to (n-1) for each value of principal quantum number, n. This number defines the **shape of the orbital**. (Remember, orbitals are not electrons but regions in space where electrons can be and each orbital can hold up to two electrons only.)

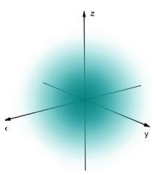
The value of ***ℓ*** for a particular orbital is generally designated by the letters s, p, d, and f, * corresponding to ***ℓ*** values of 0, 1, 2, and 3:

(Brown et.al, 2012)

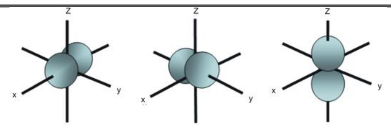
Value of l	0	1	2	3
Letter used	s	p	d	f
Maximum number of electrons	2	6	10	14

Table 1. Subshells' designation and maximum number of electrons each subshell can accommodate

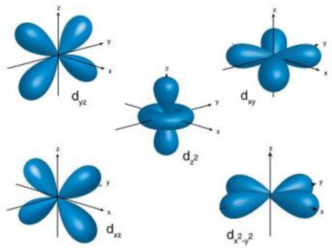
A collection of orbitals with the same value of n is frequently called a shell. One or more orbitals with the same n and ***ℓ*** values are referred to as a subshell. The s subshell (***ℓ*** =0) has one s orbital only. Therefore, it can hold a maximum of two electrons, while the p subshell (***ℓ***=1) has actually three p orbitals (p_x, p_y, p_z) which can accommodate 2 electrons for each orbital (maximum of 6 electrons). The d subshell (***ℓ*=**2) has five d orbitals and the f subshell (***ℓ***=3) has seven f orbitals (see Figure 3). **(Chang, 2008)**



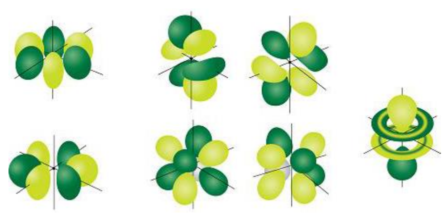
s orbital



p orbitals



d orbitals

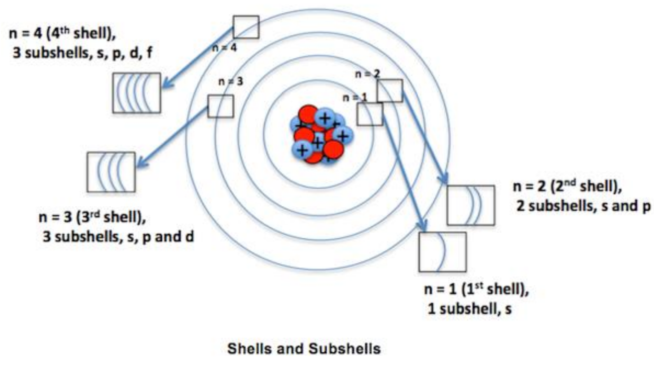


f orbitals

Source: Shape of orbitals. Retrieved from wikibook.org

Figure 3. The different shapes of s, p, d, and f orbitals

Let's have an example. For the shell with **n=2**, it is composed of two subshells**, *ℓ*=0** and ***ℓ*=**1 (the allowed values for n=2). These subshells are called the **2s** and **2p** subshells where 2 denotes the value of n, and s (***ℓ*** =0) and p (***ℓ*** =1) denote the values of ***ℓ***.



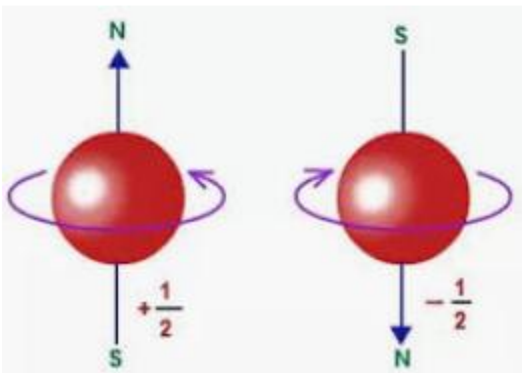
Source: What are shells, subshells, orbit and orbital? Retrived from quora.com

Figure 4. Shells and subshells

3. **Magnetic quantum number (m_ℓ)** is a quantum number that describes the **orientation of the orbital** in the space around the nucleus. It is also called the orbital-orientation quantum number or azimuthal quantum number. The possible values of an orbital's magnetic quantum number, m_ℓ, are integers from - ℓ through 0 to + ℓ. For example, an orbital with ℓ=0 can only have m_ℓ=0 and so the orbital has only one orientation in space. However, an orbital with ℓ=1 can have three m_ℓ values, -1, 0, or 1; thus, there are three possible orientation of orbitals with ℓ=1. Note that the number of possible m_ℓ values equals the total number of orbitals for each subshell or value of ℓ, which is given by this formula: $2\ell+1$ **(Chang, 2008).**

To conclude our discussion of these three quantum numbers, let us consider a situation in which **n=2 and ℓ=1**. The values of n and ℓ indicate that an electron is possibly located in second energy level or shell (n=2) and within this shell, it is probably in p subshell (ℓ=1). A short notation would be in a **2p subshell**, and in this subshell we have three 2p orbitals in different orientations ($2p_x$, $2p_y$, $2p_z$) because there are three values of m_ℓ, given by -1, 0, and 1 **(Chang, 2008).** Now let's proceed to the last quantum number.

4. **Spin quantum number (m_s)** is the last quantum number needed to be introduced to take the electron spin into account. This will only have a value of either ½ or – ½ no matter what **(Chang, 2008).** This is because any orbital can only hold up to two electrons and even the two electrons in the same exact orbital will have opposite spin values. (1)



Source: What is the spin quantum number. Retrieved from quora.com

Figure 5. The (a) counterclockwise and (b) clockwise spins of an electron. The magnetic fields generated by these two spinning motions are analogous to those

from the two magnets. The upward and downward arrows are used to denote the direction of spin.

FORMATIVE ASSESSMENT (NON-GRADED)

A. What values of the angular momentum (ℓ) and magnetic (m_ℓ) quantum numbers are allowed for a principal quantum number (n) of 3? **(Silberberg, 2009)**

B. How many orbitals exist for n=3?

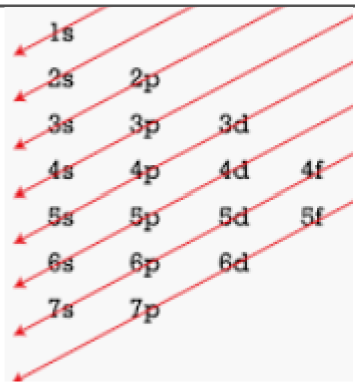
Take note that every electron in an atom has a unique set of quantum numbers. As Linus Pauling, an American Chemical Engineer and one of the founders of the fields of quantum chemistry and molecular biology, in his Exclusion Principle said that *no two electrons in an atom can have precisely the same four quantum numbers.*

Electronic Configuration

Now that you are already familiar with the different quantum numbers, you are equipped to learn how electrons are distributed in an atom and that is what we call **electronic configuration**. In this case we will focus on determining the ground state electron configuration of atoms.

Recall, that the more electrons an atom has, the more orbitals it will need to accommodate them all.

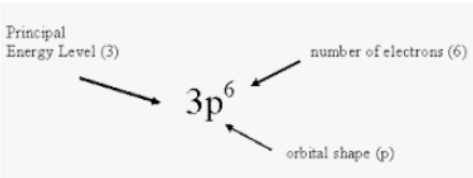
The **Aufbau Principle** (from the German word *aufbeem*, meaning "building up") tells us that orbitals are filled up in order of increasing energy and from Pauli's principle, there should be no more than two electrons per orbital. Orbitals that are nearer to the nucleus, have lower potential energy so 1s will be filled first then 2s then 2p and so forth. We can use the diagonal rule below to determine the electron configuration of any atom. **(Brown et.al, 2012)**



Source: Aufbau Principle. Retrieved from courses.lumenlearning.com

Figure 6. Aufbau Principle states that electrons will fill up the lowest energy level first which is 1s, then 2s, 2p, 3s so on and so forth.

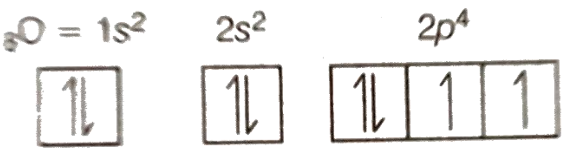
We can represent any electron configuration by writing the symbol for the occupied subshell and adding a superscript to indicate the number of electrons in that subshell.



Source: Chapter 10. Retrieved from mrwiggersci.com

Figure 7. Important Parts in Writing Electron Configuration

For example, a neutral oxygen atom has 8 electrons. We determine the electron configuration or distribution of these 8 electrons by starting filling up the lowest energy orbital (which is 1s) first and then climb upwards. We can use **orbital box diagrams** to visually depict how the orbitals are filled and they look like as in Figure 7. Each **box indicates an orbital** (which can hold a maximum of 2 electrons or paired electrons with opposite spins) while **electrons are depicted as arrows**. (2)



Source: Electrons in oxygen Atom. Retrieved from doubtnut.com

Figure 7. Electronic Configuration and Orbital Box Diagram of Oxygen atom.

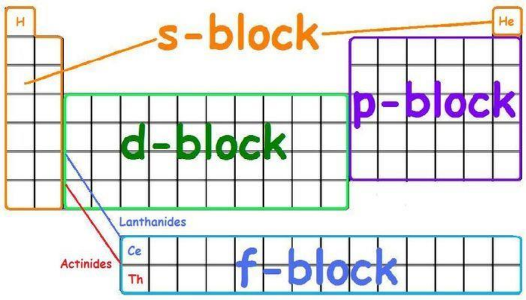
The 1s orbital gets 2 electrons, one spin up the other spin down. After it is full, we move on to the next orbital, 2s and the last four will go in 2p orbitals. Notice the filling up pattern in Figure 7. In determining electron configuration, we must determine which one will give the greatest stability. The answer is provided by *Hund's rule, which states that the most stable arrangement of electrons in subshells is the one with the greatest number of parallel spins. In other words, Hund's Rule says that for electrons of the same energy (or electrons in same n and* ℓ *quantum numbers) you put one electron in each orbital first before doubling them up* **(Chang, 2008).** This is because electrons repel one another and if they were in the same orbital, they would be close together. For example, orbitals in the 2p subshell are degenerate - in other words the 2px, 2py, and 2pz orbitals are equal in energy so you fill up 1 electron each before doubling them up. (Friedrich Hund is a German physicist known for his work on the electronic structure of atoms and molecules.) **(Britannica, 2020).**

To summarize, oxygen has an electron configuration of:

$$1s^2\ 2s^2\ 2p^4$$

There are two main exceptions to electron configuration however: chromium and copper. In these cases, a completely full or half full d sub-level is more stable than a partially filled d sub-level. **(Brightstorm Inc, 2020)**

If you want a quicker way to determine electron configuration just know which areas of the periodic table correspond to which types of orbitals.

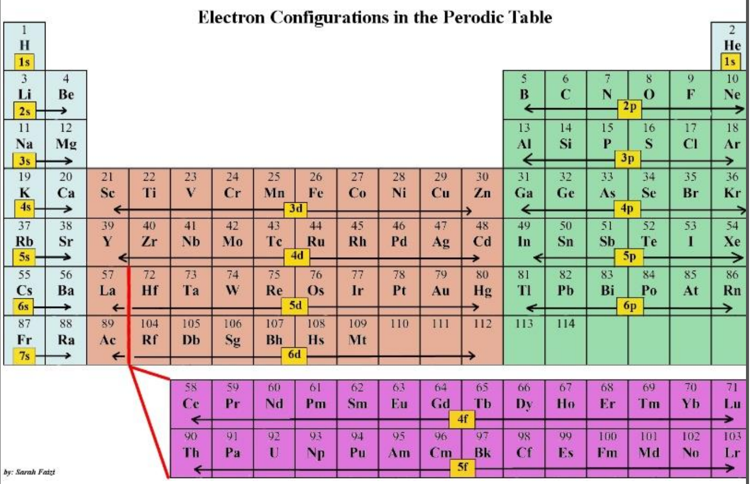


Source: A periodic table for learning quantum numbers. Retrieved from pinterest.com

Figure 8. A block of the periodic table is a set of elements unified by the type of orbitals their valence

electrons lie in.

When we are determining the electron configuration for an atom of a given element, the easiest way to do it is just travel left to right and top to bottom on the table until you get to the element since each element has one more electron than the last.



Source: Valence Electron. Retrieved from pinterest.com

Figure 9. Electronic configuration of elements under same block shows that their valence electrons occupy same type of orbitals.

Looking at oxygen again, the top row is just a 1s orbital that's full so $1s^2$. Going to the second row where n=2, we fill the 2s orbital which holds 2 electrons; $2s^2$. Hence two elements, and lastly we get to the 2p's and count 1 2 3 4 until we arrive to the oxygen atom; $2p^4$. Overall, we passed through $1s^2\ 2s^2\ 2p^4$.

A common abbreviation is using the condensed form by writing the noble gas from the row above (list it in brackets) and copy the rest like this: [He] $2s^2\ 2p^4$. (3)

In summary, the number and letter pairs in an electron configuration represent two of the electron's four quantum numbers. These quantum numbers tell us more information about the properties of electrons and their orbitals.

<table>
<tr>
<td>Navigate
</td>
<td>FORMATIVE ASSESSMENT(GRADED):

Answer the following questions in a ½ sheet of paper, have it scanned and submit through our assigned learning platform.

Give the name, magnetic quantum numbers, and number of orbitals for each sublevel with the given quantum numbers: (Tabulate your answers.)

(a) $n = 3, \ell = 2$ (7pts)
(b) $n = 2, \ell = 0$ (3pts)
(c) $n = 5, \ell = 1$ (5pts)
(d) $n = 4, \ell = 3$ (9pts)</td>
<td>5 mins.</td>
<td></td>
</tr>
<tr>
<td>Knot
</td>
<td>Four quantum numbers characterize each electron in an atom: the principal quantum number (n) identifies the main energy level, or shell, of the orbital; the angular momentum quantum number (ℓ) indicates the shape of the orbital or subshell; the magnetic quantum number m_ℓ specifies the orientation of the orbital in space; and the electron spin quantum number m_s indicates the direction of the electron's spin on its own axis. These quantum numbers help us to define the electron configuration of an atom. There are three important principles to apply when determining electron configuration of an atom and these are:

• Pauli's Exclusion Principle: No two electrons may have the same set of four quantum numbers.

• Aufbau Principle: Orbitals are filled up in order of increasing energy. Fun fact: Aufbau is German for building up.

• Hund's Rule: When electrons are in orbitals with the same energy or degenerate orbitals, they have to fill up the empty orbitals first before doubling them up.

In addition, the position of an atom in the periodic table also helps us to easily define the electron configuration of any</td>
<td>2 mins.</td>
<td></td>
</tr>
</table>

	atom. In the next lesson, we will focus our study mainly on periodic table of elements so please study in advance about the history of the periodic table.		

[a] suggested time allocation set by the teacher

[b] actual time spent by the student (for information purposes only)

Endnotes

(1) Quantum Numbers and Electronic Configuration.(https://www.youtube.com/watch?v=Aoi4j8es4gQ&t=382s)
(2) How to Draw Orbital Diagrams and Hund's Rule. (https://www.youtube.com/watch?v=Y89s89kEsJY)
(3) How to Write Electron Configuration.(https://www.youtube.com/watch?v=iFN9agJVea4)

References

Brown, T.L., LeMay, H.E., Bursten, B.E., Murphy, C.J., and Woodward, P.M. (2012) Chemistry: The Central Science, (12th Edition). Pearson Publishing Inc.

Chang Raymond (2008) General Chemistry: The Essential Concepts (5th Edition). McGrawHill Higher Education.

Silberberg, Martin S., (2006) Chemistry: The Molecular Nature of Matter and Change (4th Edition). McGraHill Higher Education

Quantum Numbers and Electronic Configuration. (n.d.). Retrieved from https://chemed.chem.purdue.edu/genchem/topicreview/bp/ch6/quantum.html

Friedrich Hund. March 27, 2020. Retrived from https://www.britannica.com/biography/Friedrich-Hund

Exceptions to Electron Configuration Concept. Brightstorm Inc, 2020. Retrieved from https://www.brightstorm.com/science/chemistry/the-atom/exceptions-to-electron configuration

Prepared by: Carren Anne Y. Dalisay
Position: Special Science Teacher II
Campus: PSHS-SRC

Reviewed by: PRINCESS ANN B. DIGNENENG
Position: Special Science Teacher II
Campus: PSHS-CLC