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Learning Guide Module

Subject Code **Chem** 1 Chemistry 1
LG Code 3.0 Electronic Structure of Atoms and the Periodic Table
Lesson Code 3.1 Dalton's Atomic Theory -atomic nature of matter Part 1-A and 1-B
Time Frame 30 minutes

Components	Tasks	TA (min)[a]	ATA (min)[b]
Target 	After completing this learning guide, students are expected to: 1. Calculate the atomic mass of an element given its isotopes' percent natural abundance. 2. Apply Pauli's Exclusion Principle, Hund's Rule, and Aufbau Principle to the distribution of electrons around the nucleus. 3. Use the electron configuration to determine the position of an element in the periodic table and vice versa.	1	
Hook 	**What is the difference between the atomic mass and atomic weight of an element?** **Atomic mass** is the mass of an atom which is the sum of protons and neutron. Electrons contribute so little mass that they are not counted. The atomic mass of an element won't change, for example Carbon -6 atom has an atomic mass of 12 (*Helmenstine, 2019, para.1*). The **atomic weigh**t of an element is the weighted average of the isotopic masses, according to the naturally occurring abundances of the isotopes of the element (*Petrucci, Herring, Madura, Bissonnette, 2011, p. 48*). In nature, most elements occur as a mixture of two or more isotopes. Each isotope of an element has a fixed mass and a natural percent abundance. Atomic weights of elements are indicated in periodic tables. **Electron configurations.** In an atom, electrons and the nucleus interact to make the most stable arrangement possible. And the ways in which electrons are arranged in various orbitals around the nucleus of atoms are called electron configurations (*Brown, LeMay, Murphy, Woodward, 2009*). The position of an element in the periodic table can be determined by its electron configuration.	3	

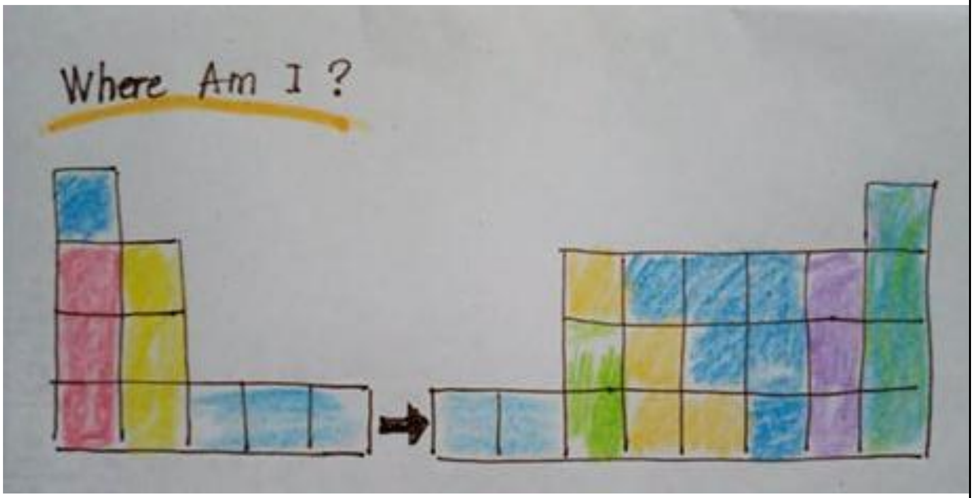

Figure 1: Empty Periodic Table

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How to calculate the atomic weight of an element which has two or more natural isotopes and has relative percent abundance?

Formula in calculating the atomic weight of an element which has two or more natural isotopes and has relative percent abundance.

$$\sum_{i=1}^{n} x_i y_i = x_1 y_1 + x_2 y_2 + \ldots x_n y_n$$

where: x is the mass of the isotope.

y is the % natural abundance of the isotope.

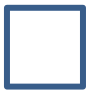
A sample problem is discussed in the navigate section found on pages 8 & 9 part a.

What are the three rules for writing the electron configurations?

The three rules are as follows:

1.1. **Pauli's Exclusion Principle.** It states that an atomic orbital may describe at most two electrons. **No two electrons in an atom can have all four quantum numbers alike.** To occupy the same orbital, two electrons must have opposite spins; that is, the electron spins must be paired.

Spin is a quantum mechanical property of electrons and may be thought of as clockwise or counterclockwise. The orbitals are filled in order of increasing energy, with no more than 2 electrons per orbital (*Wilbraham, Staley, Matta, Waterman, 2005*).

A box represents an orbital. 

A vertical arrow represents an electron and its direction of spin (↑or ↓).

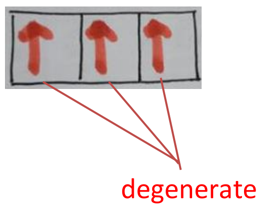
1s

A completely filled-out orbital. 

1.2. **Hund's Rule.** It states that electrons occupy orbitals of the same energy in a way that makes the number of electrons with the same spin direction as large as possible.

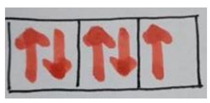
For example, three electrons would occupy three orbitals of equal energy or ***degenerate orbitals.***

2p



If there are five electrons, the remaining electrons will occupy the first and second orbitals so that their spins are paired. Thus the first 2 orbitals can have two electrons in opposite spins.

2p



Important!

According to Hund's Rule, the most stable electron configuration is the one with the most number of unpaired electrons. Wherein the electrons are in their lowest possible states. This is called the ***ground-state electron configuration***. The electron configuration of the same atom that enters the ***excited state*** is not stable.

1.3. **Aufbau Principle.** According to this principle, electrons occupy the orbitals of lowest energy first. The orbitals for any sublevel of a principal energy level are always of equal energy. Within a principal energy level the **s** sublevel is always the lowest - energy sublevel. The range of energy levels within a principal energy level can overlap the energy levels of another principal energy level.

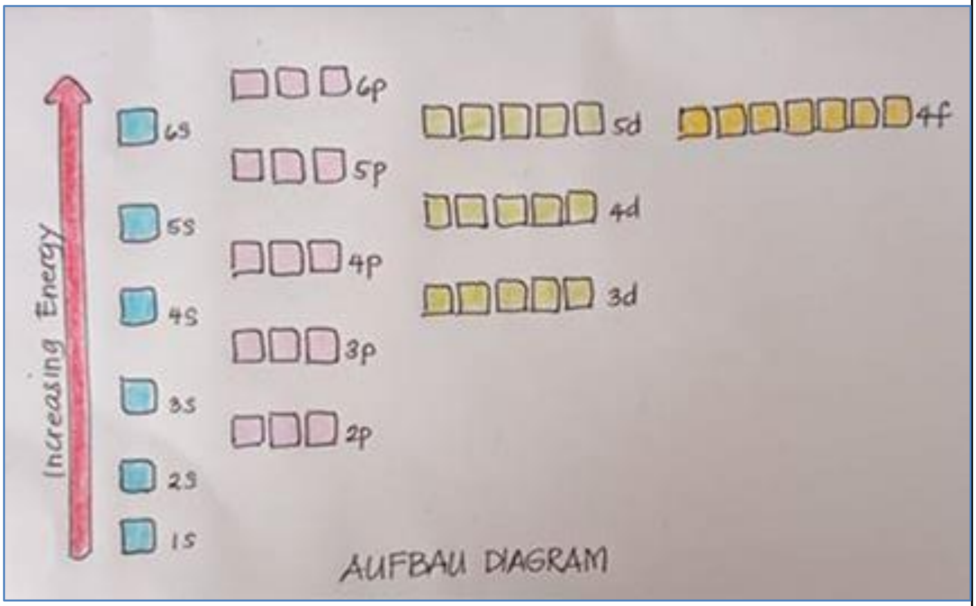


Figure 1: Aufbau Diagram which shows the energy level of various atomic orbitals.

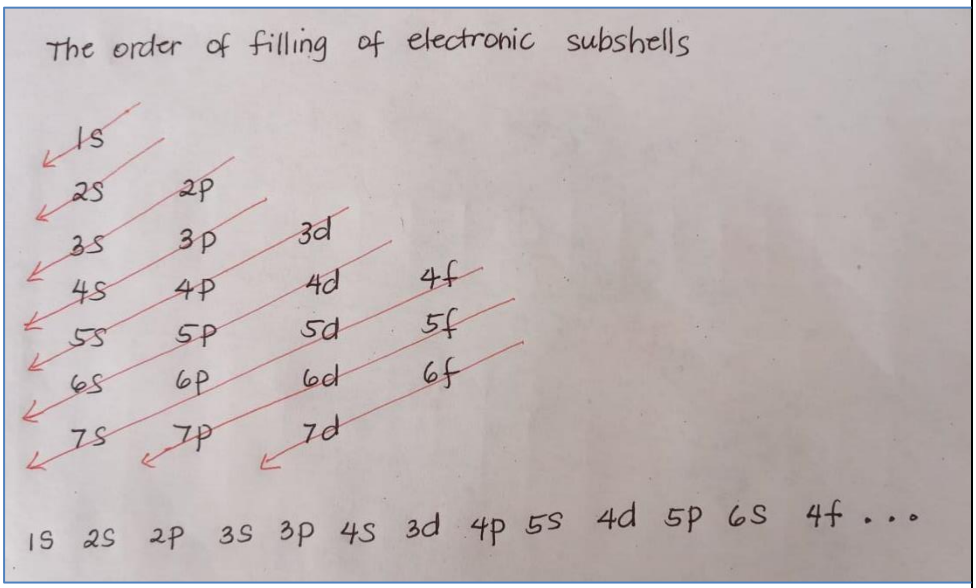


Figure 2: The order of filling of electronic configuration

In writing the electron configurations using the **aufbau process,** we follow the order of filling of electronic subshells as shown in figure 2.

s subshell (has one orbital) can hold a maximum of 2 electrons.

p subshell (has three orbitals) can hold a maximum of 6 electrons.

d subshell (has 5 orbitals) can hold a maximum of 10 electrons.

f subshell (has 7 orbitals) can hold a maximum of 14 electrons.

For example, write the electron configuration of Chlorine (Z=17)

$1s^2$ $2s^2$ sp^6 $3s^2$ $3p^5$

Methods in representing electron configurations:

spdf notation (condensed) : Carbon (Z=6) $1s^2\ 2s^2\ 2p^2$

spdf notation (expanded) : Carbon (Z=6) $1s^2\ 2s^2\ 2px^1\ 2py^1$

(***Hund's rule is reflected)***

Orbital diagram : Carbon (Z=6)

↑↓	↑↓	↑	↑	
$1s^2$	**$2s^2$**		**$2p^2$**	

Condensed electron configuration. In writing the condensed electron configuration of an element, the electron configuration of the nearest noble-gas element of lower atomic number is represented by its chemical symbol in brackets. For example,

$$\mathbf{C:\ [He]2s^2\ 2p^2}$$

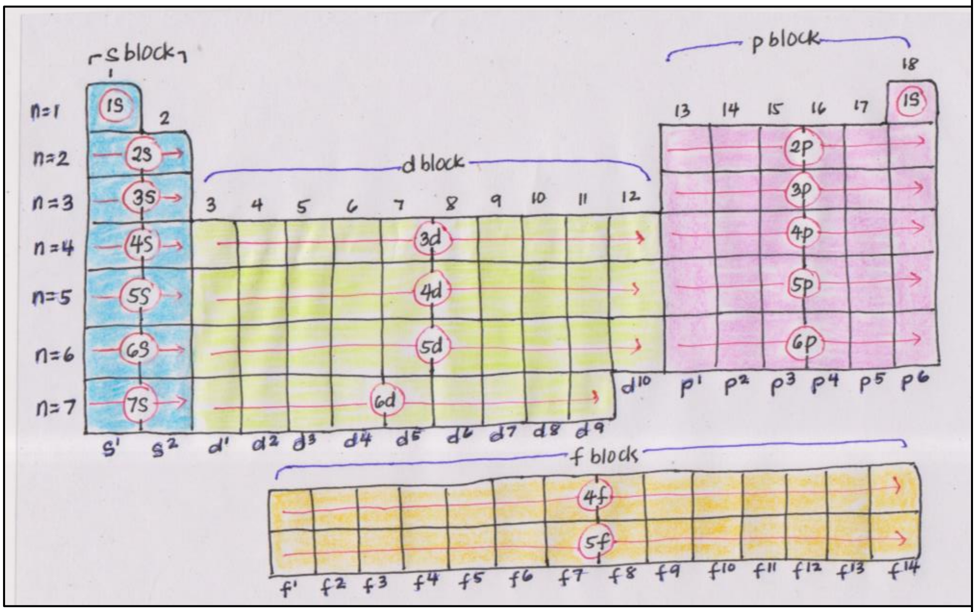


Figure 3: Electron configuration and the periodic table

How to determine the position of an element like carbon in the periodic table using the electron configuration **Carbon (Z=6) $1s^2\ 2s^2\ 2p^2$** ?

Refer to figure 3:

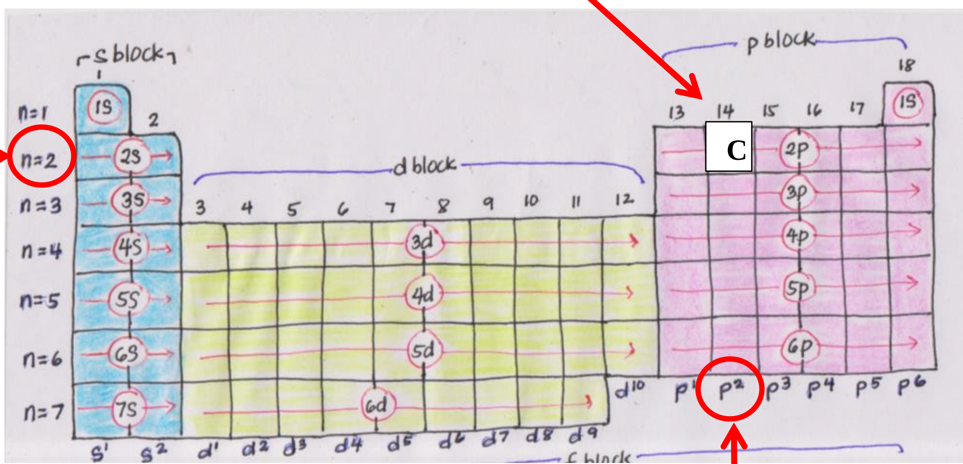
*The **s** block (the filled **s** orbital of highest principal quantum number ,**n**) consists of group 1 and 2 plus He in group 18.*

*The **p** block (the filled **p** orbital of highest principal quantum number ,**n**) consists of groups 13 -18 except He.*

*The **d** block (the filled **d** orbital next to the outermost energy level ,**n-1**) includes groups 3-12.*

*The **f** block (the filled **f** orbital, **n-2**) are the Lanthanides and Actinides.*

Carbon (Z=6) $1s^2$ $2s^2$ $\underline{2p^2}$



[a] TA – time allocation suggested by the teacher
[b] ATA- actual time allocation spent by the student (for information purposes only)

Subject Code Chem 1 Chemistry 1
LG Code 3.0 Electronic Structure of Atoms and the Periodic Table
Lesson Code 3.1 Dalton's Atomic Theory - atomic nature of matter Part 1-B
Time Frame 30 minutes

Components	Tasks	TA (min)[a]	ATA (min)[b]
Navigate	To further understand the concepts learned, try to solve the practice problem given below. ***Formative Assessment (Non-graded Student's Activity):*** Boron, atomic number 5, occurs naturally as two isotopes. The isotope with a mass of 10.012 amu (**^{10}B)** has a natural abundance of 19.91%. The isotope with a mass of 11.009 (**^{11}B)** has a natural abundance of 80.09% (*Brown et al, 2009, p. 243*). a. Calculate the atomic mass of this element. b. Write the electron configuration in spdf notation of Boron. c. Draw the orbital diagram. d. Write the condensed electron configuration. e. Determine the position of Boron in the periodic table. **Steps:** 1. Analyze the problem and write down the given data. Isotope ^{10}B: mass = 10.012 amu, % natural abundance = 19.91% Isotope ^{11}B: mass = 11.009 amu, % natural abundance = 80.09% 2. Write down what is/are asked for. a. atomic mass of Boron b. electron configuration in spdf notation of Boron. c. orbital diagram. d. condensed electron configuration. e. position of Boron in the periodic table. 3. Calculate or solve for the unknown/do what is asked. a. Using the summation formula:	15	

$$\sum_{i=1}^{n} x_i y_i = x_1 y_1 + x_2 y_2 + \ldots x_n y_n$$

= (10.012 amu)(0.1991) + (11.009)(0.8009)

= 10.810 amu

b. **spdf** notation: $\mathbf{1s^2\ 2s^2\ 2p^1}$

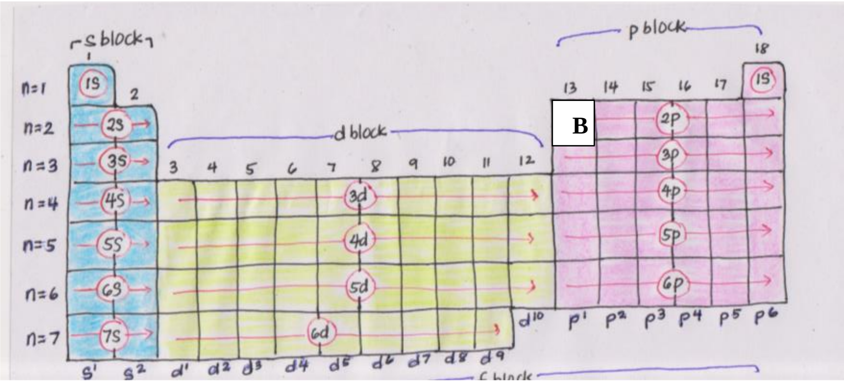
c. orbital diagram: ↑↓ ↑↓ ↑

$\mathbf{1s^2}$ $\mathbf{2s^2}$ $\mathbf{2p^1}$

d. Noble-gas element of lower atomic number: **He**

Condensed electron configuration: $\mathbf{[He]2s^2\ 2p^1}$

e. Locate the exact position of Boron in the periodic table. $\mathbf{1s^2\ 2s^2\ 2p^1}$



4. Evaluate your answer/s (Do/does the result make sense?)

Note: Check each result as a self-assessment.

Formative Assessment (Graded): (8 points)

Write the electron configurations of Chromium and copper in spdf notation and condensed electron configuration. Apply Pauli's exclusion principle, Hund's rule, and Aufbau principle in writing their ground-state (most stable) electron configurations.

Submit your answers with complete solution three days after finishing the non-graded student activity.

(40% of your grade in this activity is your compliance and 60% is from your score. Please submit your answers on or before the deadline. Corresponding deductions will be applied for late submission or for noncompliance).

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As a summary, the atomic mass of an element is the sum of protons and neutron while atomic weight is the weighted average of the isotopes' masses based on their percent natural abundances.

The electron configurations of elements may be presented in **spdf notation**, **orbital diagram**, and **condensed electron configuration** by applying the ***Pauli's exclusion principle***, ***Hund's rule***, and ***Aufbau principle***. The location or position of elements in the periodic table can be determined by their electron configurations.

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Problem Set: (10 points)

Please answer the following problems given below. This is a graded problem set for this quarter. **40% of your grade in this activity is your compliance and 60% is from your score**. Please submit your answers on or before the deadline which is a week after completing this Learning Guide. Corresponding deductions will be applied for late submission or for noncompliance.

1. The element oxygen has naturally occurring isotopes with mass numbers of 15.995 amu, 16.995 amu, and 17.999 amu . Their natural abundances are 99.759%, 0.037%, and 0.204% respectively. Calculate the atomic weight of oxygen. **(3 points)**

2. Without looking at your periodic table, which list of elements would fall within the same group in the periodic table?
(Hint: write the electron configurations of the given elements). **(3 points)**
Atomic numbers: **$_{2}$He, $_{18}$Ar, $_{54}$Xe, $_{8}$O, $_{9}$F, $_{10}$Ne, $_{19}$K, $_{37}$Rb, $_{56}$Ba, $_{1}$H, $_{3}$Li**

   a. He, Ar, Xe
   b. O, F, Ne
   c. K, Rb, Ba
   d. H, He, Li

3. Use the periodic table to write the condensed electron configurations in their ground states. **(3 points)**

   a. Bromine
   b. Cobalt
   c. Tellurium

4. Identify the element with the electron configuration of **$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$.** **(1 point)**

[a] TA – time allocation suggested by the teacher
[b] ATA- actual time allocation spent by the student (for information purposes only)

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.

Helmenstine, Anne Mary, (2019) The Difference Between Atomic Weight and Atomic Mass. (*thoughtco.com*), updated on July 1, 2019, paragraph 1.

Petrucci, Ralph H., Herring, F. Geoffrey., Madura, Jeffrey D., Bissonnette, Carey, (2011) General Chemistry: Principles and Modern Applications (10th Edition) Pearson Canada Inc.

Silverberg, Martin S., (2006) Chemistry: The Molecular Nature of Matter and Change (4th Edition) Mc Graw Hill Higher Education.

Wilbraham, Anthony C., Staley, Dennis D., Matta, Michael S., Waterman, Edward L. (2005) Chemistry (4th Edition) Prentice-Hall Inc.

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