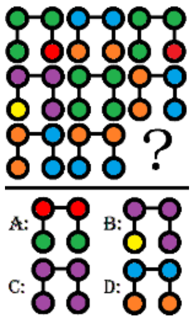
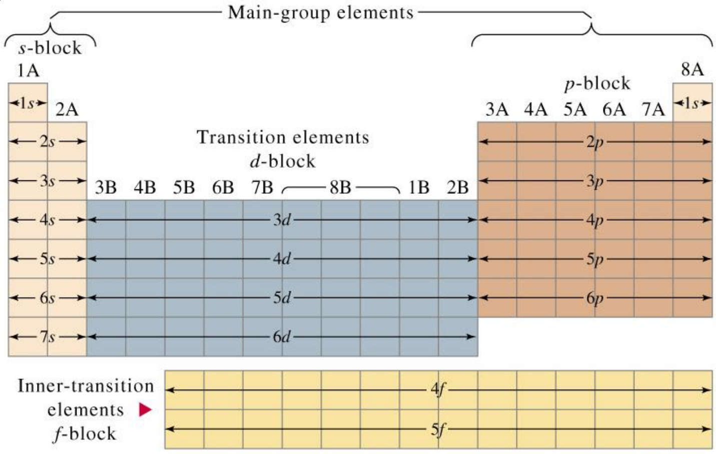
Learning Guide Module

Subject Code	Chem 1	General Inorganic Chemistry
Module Code	3.0	*Electronic Structure of Atoms and the Periodic Properties of Elements*
Lesson Code	3.3	*Periodic Table*
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. Determine the valence electron based on element's position in the periodic table. 2. Predict the properties of an element based on its position in the periodic table 3. Relate the physical and chemical properties of the elements with the observed trends in the periodic table.	1 min.	
Hook 	Let's have a brain exercise and guess the next pattern to the figure below.  Mind-boggling isn't it? But if you look and study the picture closely, the answer is ….. (oops! Answer will be revealed later. ☺) Let's focus first on our lesson today. Patterns! Patterns! Patterns everywhere! From the soles of your feet to stunning feathers of peacocks. That's how nature operates. It is somewhat a congregation of patterns that when observed closely, makes you amazed and appreciate life. That's the same with the chemical elements in the periodic table. At first, it seems like a random arrangement of substances but the way the elements are arranged reveals many beautiful patterns that tells us about how nature operates.	2 mins.	

	 Source: Patterns in Nature Photography. Retrieved from chrlesnolder.zenfolio.com **Figure 1**. The patterns make the peacock's feathers more beautiful. We could not imagine if the elements were just randomly arranged but thanks to Dmitri Mendeleev, a Russian chemist. In the mid 1800's lots of chemists were trying to come up with a way to depict all the elements in table form. Many different formats were proposed but it was the one by **Dmitri Mendeleev** that stuck because of how well it correlates data as well as its predictive powers. He arranged the elements, according to atomic weights, into rows called *periods* and columns called *groups*. **(Silberberg, 2009)** Though another scientist named **Henry Moseley** improved the arrangement of the elements according to atomic numbers and hence our modern Periodic Table.		
Ignite 	Why does the periodic table have the structure it does? If you understand electron configurations, you will discover that the shape of the periodic table portrays the filling of the subshells with electrons.	20 mins.	



Source: Electron Configurations and the Periodic Table. Retrieved from kentchemistry.com

Figure 2. The periodic table is separated into blocks depending on which subshell is being filled for the atoms that belong in that section.

Figure 2 shows the blocks of the periodic table. The s-block is the region of the alkali metals including Helium (Groups 1 & 2), the d-block are the transition metals (Groups 3 to 12), the p-block are the main group elements from Groups 13 to 18, and the f-block are the lanthanides and actinides series, that is depicted as detached from the main body of the periodic table. It could be part of the main body, but it could look efficiently this way.

Let's go back to electron configuration and try to relate it with chemical periodicity, but first let's define some terms. The electrons in the outermost (highest-numbered, n) shell, plus any electrons in the last unfilled subshell, are called *valence electrons*; the outermost shell is called the *valence shell*. (The inner electrons are called core electrons.) The valence electrons largely control the chemistry of an atom. If we look at just the valence shell's electron configuration, we find that in each column in the periodic table, the valence shell's electron configuration is the same.

For example, take the elements in the first column of the periodic table: H, Li, Na, K, Rb, and Cs. Their electron configurations (abbreviated for the larger atoms, see link at endnotes for review on steps on writing abbreviated and non-abbreviated electron configuration) are as follows, with the valence shell electron configuration highlighted: (1)

H:	$1\mathbf{s^1}$
Li:	$1s^2 2\mathbf{s^1}$

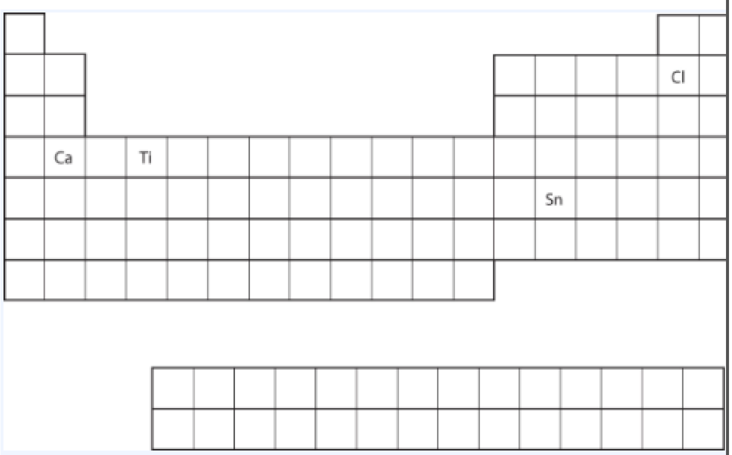
Na:	[Ne]$3s^1$
K:	[Ar]$4s^1$
Rb:	[Kr]$5s^1$
Cs:	[Xe]$6s^1$

Table 1. Electron configuration of first column elements shows that each element has 1 electron in their outermost shell.

They all have a similar electron configuration in their valence shells: a single s electron. You will know in the later topics that the properties of an element are actually influenced by valence electrons, so we would expect that elements in same column would have similar properties and they do. The same concept applies to the other columns of the periodic table.

Example Exercise 1.

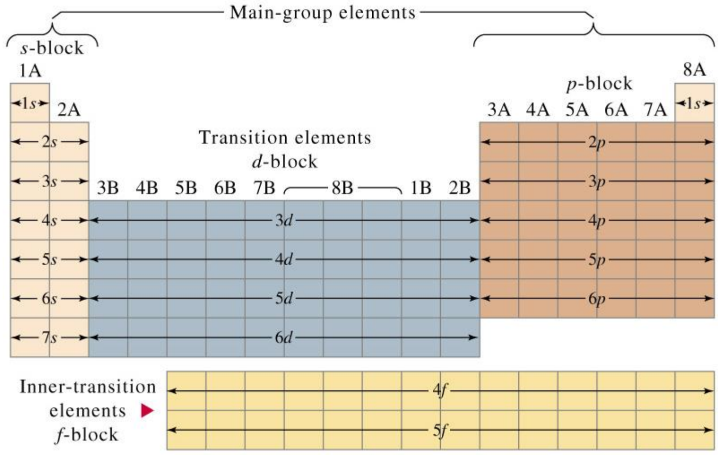
From the element's position on the periodic table, predict the valence shell electron configuration for each given atom (Figure 2)



(a) Ca

(b) Sn

Answer. (Look at Figure 2)



a. Ca is located in the second column of the *s* block. We would expect that its electron configuration should end with s^2 and the preceding noble gas core is Argon so Calcium's electron configuration is $[Ar]4s^2$.

b. Sn is located in the second column of the *p* block, so we expect that its electron configuration would end in p^2 and the preceding noble gas core is Krypton so Tin's electron configuration is $[Kr]5s^2 4d^{10} 5p^2$

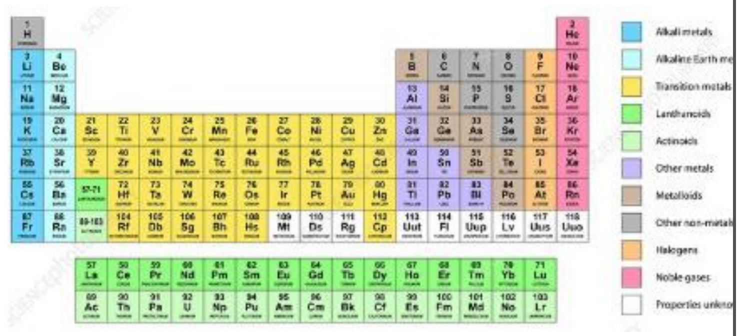
Note: The designation of d block starts with n-1 hence after 4s orbital is 3d orbital and not 4d orbital and f block starts with n-2.

FORMATIVE ASSESSMENT 1(NON-GRADED):

From the element's position on the periodic table, predict the valence shell electron configuration for each atom.

(a) Ti
(b) Cl

Going further in the study of the periodic table, let's find out how the chemical elements were successfully organized.



Source: Standard Periodic Table. Retrieved from sciencephoto.com

Figure 3. The Periodic Table of Elements

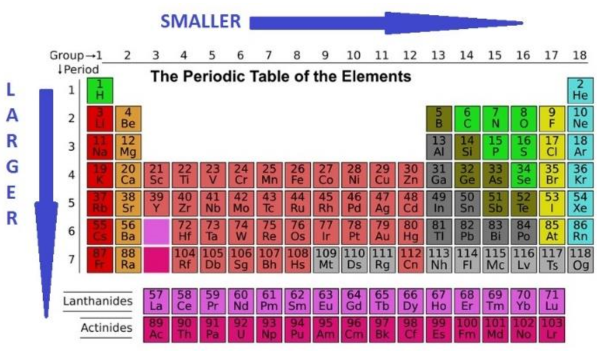
As mentioned before, it was Dmitri's arrangement that had accurately organized the chemical elements. Elements that have similar behavior were put in groups which help correlate existing data. It also predicted the existence of elements that had never been seen before. Mendeleev said there must be elements that go in the gaps in the table and he predicted some of their properties. Eventually these elements were discovered with the properties precisely as expected and now we have all the metals, metalloids and nonmetals organized nicely (see Figure 3).

It wasn't known at the time but the reason elements in the same group behave similarly is because they have the same number of valence electrons as we have pointed out previously.

This simple fact determines many characteristics about each element in which we will continue to see. There are some periodic trends that we can recognize when we look at the table.

Periodic Trends

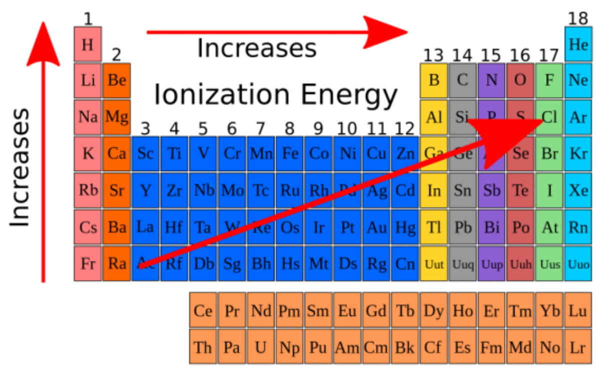
1. The first periodic trend that we will tackle is *atomic radius* or the size of the atom. As we proceed down a group, atomic size increases because the number of energy levels (n) increases, so there is a greater distance between the nucleus and the outermost orbital. As we move to the right, atomic radius decreases because we are moving within the same shell (n) and each element to the right has 1 more proton in the nucleus than the last. Therefore, there is a stronger electromagnetic attraction felt by the electrons and the radius shrinks.



Source: Understanding Atomic Radius Trends. Retrieved from blog.prepscholar.com

Figure 4. Atomic radius increases from top to bottom of a group and decreases from left to right of a period**.**

2. Next, we look at *ionization energy* trend. This is the energy required to remove an electron from the atom. It will always be an electron in the outermost shell (valence electron). The electromagnetic force that attracts the electrons to the protons in the nucleus decreases very quickly with distance. *The farther away the electron is from the nucleus, the easier it is to pull away*. This means the ionization energy trend is precisely the opposite of the atomic radius trend. Francium, a very large atom with only one valence electron will be easy to ionize because the electron is so far away from the nucleus and atoms like to have their outermost shell completely full. Losing the electron means this shell is gone and the one below is completely full so elements in Group 1 will easily lose one electron. Looking at the opposite corner with Helium there is only one shell so the electrons are very close to the nucleus and the shell is full so it is very stable. For this reason, it requires much more energy to ionize Helium. For halogens and other elements on the right side, their high ionization energy is affected by the effective nuclear charge (net positive charge experienced by the electrons) too. The ionization energy trend is summarized by the figure below.



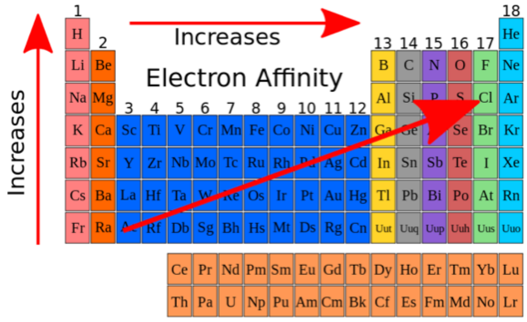
Source: Ionization Energy. Retrieved from weebly.com

Figure 5. Ionization Energy generally increases from left to right and bottom to top.

There are just a couple of exceptions to the ionization energy trend but all deviations from the ionization energy trend can be explained by discrepancies in orbital symmetry like in Oxygen and Nitrogen.

3. Next we will look at *electron affinity*. This is exactly the opposite of ionization energy since ionization energy tells us how much an atom wants to remove an electron

while electron affinity tell us how much an atom wants to gain an electron. It is actually the energy change that results from adding an electron to a gaseous atom. It increases across the period because the electrons added to energy levels (n) become closer thus more attracted to the nucleus. Remember that closer the distance, the more nuclear attraction; thus, more energy is released when an electron is added to the outside orbital. In addition, the more valence electrons an element has, the more likely it is to gain electrons to form a stable octet. The less valence electrons an atom has, the least likely it will gain electrons. Disregarding the noble gases as their shells are full, electron affinity trend goes like the figure below.

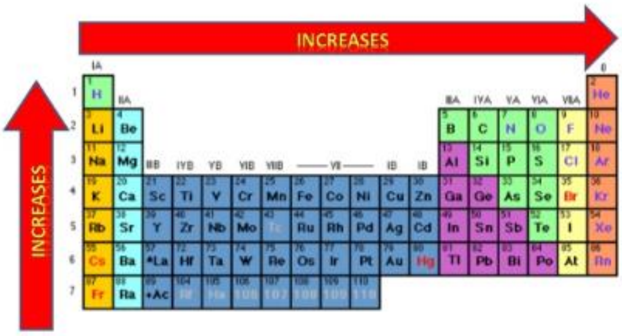


Source: Electron Affinity. Retrieved from study.com

Figure 6. Electron affinity increases from left to right and bottom to top.

Fluorine has the highest electron affinity because if it gains one electron it will have a full shell or noble gas electron configuration. Looking at the opposite corner (metals side), these elements have few electrons in their outermost shell and they would rather lose them in order to leave a stable octet in the previous shell. The exceptions to this trend happens exactly for the same reason as the exceptions to the ionization energy trend.

4. Lastly we look at *electronegativity*. This is the ability of an atom to draw electrons towards itself or hold electrons tightly. It will increase like in Figure 7. Electronegativity increases across a period (same energy level, n) because the number of protons and thus charges on the nucleus increases. This attracts the bonding pair of electrons more strongly.

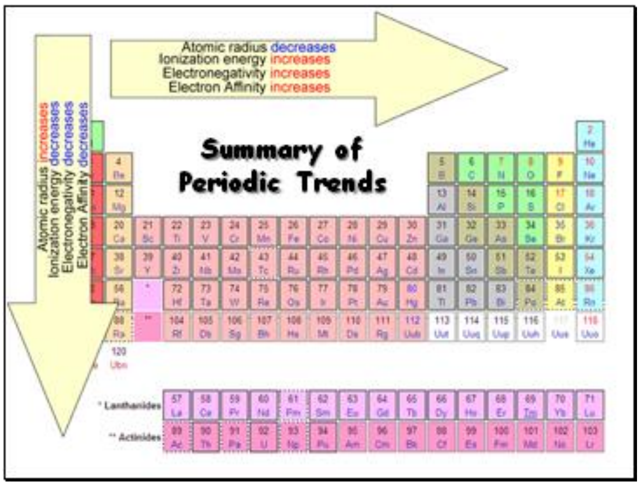


Source: Periodic Trends. Retrieved from slideshare.net

Figure 7. Electronegativity increases from left to right and bottom to top of the periodic table.

For example, smaller atom like Fluorine has more protons for its energy level so it has a higher effective nuclear charge that holds electrons strongly. (2) Again, we will disregard the noble gases for this trend. Electronegativity will be important in our next topics when we learn about chemical bonds.

To summarize, the trends to remember are: atomic radius, the only one with opposite trend, ionization energy, electron affinity and electronegativity, which all goes this way.



Source: Periodic Table of Electronegativities. Retrieved from sliderbase.com

Figure 8. Summary of all periodic trends**.**

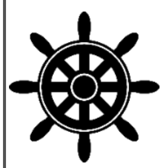
Let's check comprehension on periodic trends.

FORMATIVE ASSESSMENT 2 (NON-GRADED):

Arrange the following elements: O, C, Be, Ne in increasing:

i. Atomic radius
ii. Ionization energy
iii. Electron affinity
iv. Electronegativity

Navigate



FORMATIVE ASSESMENT (GRADED) :

5 mins.

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

Elements	Atomic Radius	Ionization Energy	Electron Affinity	Electronegativity
Na, Mg, Al				
F, Br, I				
B, C, N				

Knot

2 mins.

The electron configuration in atoms is actually related to their arrangement in the periodic table. It is because elements in each column or group in the periodic table have the same valence shell electron configurations. Keep in mind that the valence electrons largely control the chemistry of an atom.

In addition, this arrangement leads to the fact that there are certain trends or patterns that exist in the periodic table. If we look at the periodic trends of some properties of an atom, atomic radius decreases from left to right and bottom to top, while the rest of the properties like ionization energy, electron affinity and electronegativity increases from left to right and bottom to top.

[a] suggested time allocation set by the teacher

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

Endnotes

(1) How to Write Electron Configuration.(https://www.youtube.com/watch?v=iFN9agJVea4)
(2) Shielding Effect and Effective Nuclear Charge.(https://www.youtube.com/watch?v=ZPfDBu8Mhk8)

References

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J. Albarico (PSHS-CBRZ). (undated). THINK Framework. Based on Science Links by E.G. Ramos and N. Apolinario, Quezon City. Rex Bookstore.

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