Section 5 Notes H-Pre-AP Chemistry Chapter 7

This **section 5 notes** will not have a HW assigned.
However, it does have a **video** to watch, **Periodic Trends.**
Please read the notes, then watch the video.
Although you will not have an assignment for this section, there may be questions on the test concerning this information.

Noble Gas Configuration:

There is a shortened form of electron configuration called noble gas configuration. We do not include orbital diagrams (the lines and arrows under the subshell notation) with this configuration. It is very helpful in finding the valence electrons which we will discuss later and in the next chapter.

All elements have an "inner" electron configuration that is identical to all other elements in the same period that follow a certain noble gas.
For example, let's consider the 3rd period elements.
Na, Mg, Al, Si, P, Cl, and Ar all follow the noble gas Neon (Ne). All these elements have Neon's electron configuration for their "inner" electrons.
Neon's electron configuration is $\underline{1s^2\ 2s^2\ 2p^6}$.
Na is $\underline{1s^2\ 2s^2\ 2p^6}\ 3s^1$ or **Na:[Ne] $3s^1$**
Mg is $\underline{1s^2\ 2s^2\ 2p^6}\ 3s^2$ or **Mg:[Ne] $3s^2$**
Al is $\underline{1s^2\ 2s^2\ 2p^6}\ 3s^2\ 3p^1$ or **Al:[Ne] $3s^2\ 3p^1$**
Si is $\underline{1s^2\ 2s^2\ 2p^6}\ 3s^2\ 3p^2$ or **Si:[Ne] $3s^2\ 3p^2$**
P is $\underline{1s^2\ 2s^2\ 2p^6}\ 3s^2\ 3p^3$ or **P:[Ne] $3s^2\ 3p^3$**
S is $\underline{1s^2\ 2s^2\ 2p^6}\ 3s^2\ 3p^4$ or **S:[Ne] $3s^2\ 3p^4$**
Cl is $\underline{1s^2\ 2s^2\ 2p^6}\ 3s^2\ 3p^5$ or **Cl:[Ne] $3s^2\ 3p^5$**
Ar is $\underline{1s^2\ 2s^2\ 2p^6}\ 3s^2\ 3p^6$ or **Ar:[Ne] $3s^2\ 3p^6$**

(Notice Ar is a noble gas but its noble gas configuration must include the noble gas that comes before in on the table.)

For "d-block" elements, the same holds true.
Example: **Fe** would have a normal electron configuration of $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$
But, the noble gas configuration would go "back" to the last noble gas which is Ar and proceed from there—**Fe: [Ar] $4s^2\ 3d^6$**
Other random examples:
N:[He] $2s^2\ 2p^3$
Se: [Ar] $4s^2\ 3d^{10}\ 4p^4$
Sn: [Kr] $5s^2\ 4d^{10}\ 5p^2$

As stated above, Noble Gas Configuration is helpful in showing the **valence electrons**.
Valence electron are very, very important to be able to identify.
Valence Electrons are the outer shell electrons in the "s" and "p" orbitals. (**d** e^-s **not** included)
In the next chapter we will be learning how atoms **bond** and the shapes they take when bonding together to make molecules. The only **electrons involved in bonding** are **valence electrons.** As seen above, Fe has 2 val e^-, Se has 6 val e^-, N has 5 val e^-, and Sn has 4 val e^-.

We will discuss the importance of valence electrons in bonding in the next chapter. The noble gas configurations always show the **s and p electrons of the outer shell** so this is one reason that it is important.

<u>Electron Configurations of ions:</u>

Sometimes you will be asked to show the electron configuration of ions (+ and – particles).
If you have the positive ion of an element, you subtract or remove electrons from the neutral atom. If you have the negative ion of an element, you add electrons to the neutral atom.
Example: Carbon (C) has 6 electrons, C^{+2} ion has 4 electrons and C^{-2} ion has 8 electrons.
The electron configuration for
C is $1s^2 2s^2 2p^2$ or C: [He] $2s^2 2p^2$
C^{+2} is $1s^2 2s^2$ or C^{+2} [He] $2s^2$
C^{-2} is $1s^2 2s^2 2p^4$ or C^{-2} [He] $2s^2 2p^4$

Note: d-block elements (**transition metals**) will become + ions and will **always lose their s^2 electrons before the d's.**
Example:
Fe^{+2} ion would have a noble gas configuration of Fe^{+2}: [Ar] $3d^6$ (The Fe atom loses the $4s^2$ electrons first before the d-electrons.
Fe^{+3} ion would be Fe^{+3}: [Ar] $3d^5$ (The Fe atom loses the $4s^2$ then one electron from the 3d orbitals.
The transition metals lose their s-electrons first when they ionize because the s-electrons are further out than the d's. The s-electrons are valence electrons unlike the d's.

<u>Tendencies in the Periodic Table</u>

Atomic Radius (Covalent Radius)
<u>Across</u> (left to right) a period—radius **decreases—**nuclear charge increases, valence shell doesn't, pull on outer electrons gets stronger.
<u>Down</u> a group—radius **increases**—valence shells increase making the valence electrons further from nucleus,inner electrons increase and block the pull of the nuclear charge (shielding effect).

Ionization Energy (Energy needed to remove out electron.)
<u>Across</u> (left to right) a period—ionization energy **increases**—(same reason as atomic radius).
<u>Down</u> a group—ionization energy **decreases**—(same reason as atomic radius).
There are exceptions. Watch the video.

Ionic Radius (Radius of atom's most common ion)
<u>Across</u> a period—the **bigger the negative ion charge the bigger** the radius, and the **bigger the positive ion charge** the **smaller** the radius.
Negative charges have electrons added, the nuclear charge is smaller than the electrons' charges so pull is less on outer electrons and radius is bigger.
Positive charges have electrons removed, the nuclear charge is bigger than the electrons' charges so the pull is more on the outer electrons and the radius is smaller.

<u>Down</u> a group—the ionic radius is **increases** regardless of charge—(reason is the same as the down reason for atomic radius).

Electron Affinity (The energy [release] associated with an addition of an electron to an atom.)
<u>Across</u> a period—electron affinity **increases**—pull of nuclear charge gets stronger and the addition of electrons make the atom more stable (noble gas configuration).
<u>Down</u> a group—electron affinity decreases—pull of nucleus gets weaker due to distance from the nucleus and the shielding effect making it harder to add an electron to the atom.