

Unit 1, Lesson 04: Electron Configurations and Trends on the Periodic Table (pages 152 – 157)

Refer to the completed chart: “Nuclear Charge and the Shielding Effect: Explaining the Trends on the Periodic Table” to help you visualize each trend.

There are two factors that affect each of the trends that will be discussed. If you **UNDERSTAND** and can visualize these factors, the trends on the Periodic Table are easily understood:

- a) **shielding effect**: the number of electrons in full shells between the nucleus and the valence electrons
- as you move from left to right across a period (\rightarrow), shielding effect is constant. Shielding effect can not explain trends across a period.
 - as you move down a group, the shielding effect increases. There are more full electron shells, so atoms become larger and the valence electrons are further from the nucleus. This effectively decreases the attraction between the electrons and the nucleus. The increase in shielding effect explains periodic trends down a group.
- b) **net nuclear attraction** (effective nuclear charge, or Z_{eff}): net nuclear attraction is calculated by taking the nuclear charge (Z = the number of protons, or atomic number) and subtracting the shielding effect
- net nuclear attraction is a relative measure of the actual attraction between the nucleus and the valence electrons in an atom
 - as you move from left to right across a period (\rightarrow), net nuclear attraction increases. This pulls the electrons more strongly into the nucleus. Net nuclear attraction is used to explain trends across a period.
 - as you move down a group, net nuclear attraction is constant. Net nuclear attraction can not explain trends down a group.

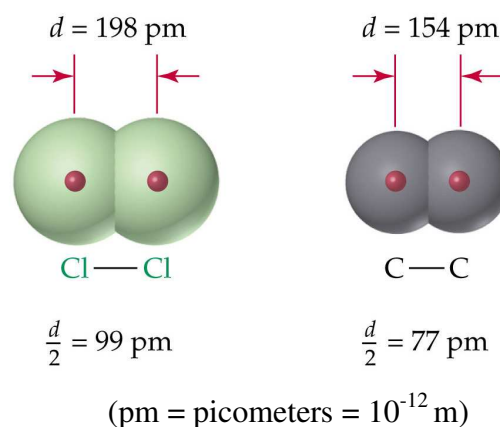
Trends across a period (\rightarrow) are due to increasing net nuclear attraction (Z_{eff}).
Trends down a group (\downarrow) are due to increasing shielding effect.

Trends

1. **Atomic Radius** indicates the size of an atom. It is calculated by measuring the distance (d) between the nuclei of two of the same type of atom, single bonded together or in the solid state, under controlled conditions.

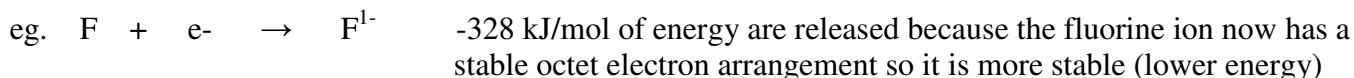
eg. for diatomic molecules (H_2 , Br_2 , F_2 , I_2 , and Cl_2). The distance between the nuclei is divided by two (2) to find the atomic radius.

- a) Atomic Radius increases down (\downarrow) a Group:
- as you move down a group, the shielding effect increases (there are more full electron shells)
 - because there are more full electron shells, the atoms are larger
- b) Atomic Radius decreases from left to right (\rightarrow) across a Period:
- as you move from left to right (\rightarrow), net nuclear attraction increases
 - because the valence electrons are more strongly attracted to the nucleus, they are pulled in closer to the nucleus, so the size of the atoms decreases as you move from left to right across a period

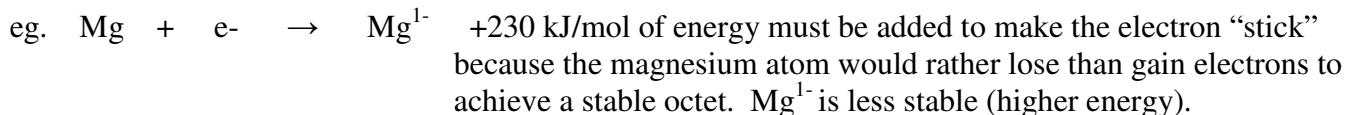


2. **First Ionization Energy** (IE_1) is defined as the amount of energy that is required to remove the outermost electron from a ground state, gaseous, neutral atom.
- a) First Ionization Energy decreases down (\downarrow) a Group:
- as you move down a group, the shielding effect increases (there are more full electron shells)
 - because the valence electrons are far away from, and less strongly attracted to, the nucleus, it requires less energy to remove a valence electron so ionization energy decreases
- b) First Ionization Energy increases from left to right (\rightarrow) across a Period:
- as you move from left to right (\rightarrow), net nuclear attraction increases
 - because the valence electrons are more strongly attracted to the nucleus, it requires more energy to remove a valence electron so ionization energy increases
3. **Electronegativity** is a measure of the relative attraction that an atom has for the electrons in a bond. It measures how strongly an atom can attract shared electron pairs to itself, relative to fluorine.
- a) Electronegativity decreases down (\downarrow) a Group:
- as you move down a group, the shielding effect increases (there are more full electron shells)
 - because the valence electrons are far away from, and less strongly attracted to, the nucleus, there is less attraction for the electrons in a bond, so electronegativity decreases
- b) Electronegativity increases from left to right (\rightarrow) across a Period (except Noble gases):
- as you move from left to right (\rightarrow), net nuclear attraction increases
 - higher net nuclear attraction means that valence and bonding electrons are more strongly attracted to the nucleus, so electronegativity increases
4. **Electron Affinity** is an indication of an atom's attraction for a new electron. It is defined as the change in energy that occurs when an electron is added to a neutral atom in its gaseous state. If an atom will be more stable when an electron is added (closer to a stable octet electron arrangement) then energy will be released when the new electron is added, and the measured energy value will be negative. The convention in science is that when energy is released, it is assigned a negative value. If an atom is less stable when an electron is added (it is further from a stable octet electron arrangement), then energy will have to be added to make the new electron "stick", so the measured energy value will be positive.

That is, a high electron affinity means a strong attraction for a new electron. This is indicated by a large, negative number because energy will be released when the electron is added. The atom is more stable (lower energy) with the added electron.



And, a low electron affinity means a weak attraction for a new electron. This is indicated by a large, positive number because energy increases when the electron is added. The atom is now less stable (higher energy) with the added electron.



4. Electron Affinity (continued)

The trends in electron affinity are quite variable. You are required only to know the level of detail below:

- a) Electron Affinity generally decreases moving down (\downarrow) a Group:
- as you move down a group, the shielding effect increases (there are more full electron shells)
 - because the valence electrons are far away from, and less strongly attracted to the nucleus, electron affinity moves closer to zero (becomes a smaller negative or positive value). This is because there is less attraction or repulsion for a new electron, so the energy changes will be smaller when an electron is added.
- b) Electron Affinity for *non-metals* is larger and more negative moving from left to right (\rightarrow) across a Period (except Noble gases):
- as you move from left to right, net nuclear attraction increases
 - as you move from left to right, non-metals (except Noble gases) are closer to a stable octet, so they release increasing amounts of energy as they gain an electron. This results in increasingly large, negative electron affinities

It may help to think about electron affinity like this:

Suppose a person really likes chocolate. The greater their “chocolate affinity”, the more they will be willing to pay for chocolate. Their bank account will go increasingly negative. So, high “chocolate affinity” is associated with large, negative bank balances.

If someone really does not like to eat chocolate, then they have low “chocolate affinity”. The lower their “chocolate affinity”, the more someone will have to pay them to make them eat chocolate. The more they have to be paid to eat chocolate, the more their bank account will increase. So, low “chocolate affinity” is associated with large, positive bank balances.

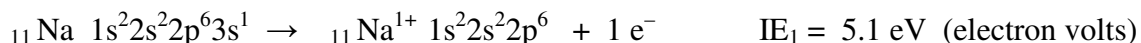
5. Ionic Radius

- a) Metal ions
- metals lose electrons to form positive ions (cations)
 - because they have fewer electrons than the neutral atom, but the same Z_{eff} , their radius is smaller than the corresponding neutral atom's
- b) Non-metal ions
- non-metals gain electrons to form negative ions (anions)
 - because they have more electrons than the neutral atom, but the same Z_{eff} , their radius is larger than the corresponding neutral atom's

In general, the atom or ion within in a period with the largest number of protons and the smallest number of electrons will have the smallest radius.

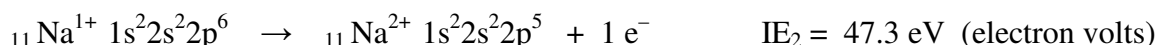
Ionization Energies (IE)

An element's **first ionization energy** (IE_1) measures the amount of energy required to remove the outermost electron from a ground state, neutral, gaseous atom. The first ionization reaction for sodium would look like this:



It requires only 5.1 eV of energy to remove the outermost electron from a neutral sodium atom, a $3s^1$ electron. By removing this electron, the sodium atom achieves a stable octet electron arrangement.

We can remove a second electron from the sodium ion (Na^{1+}), but it won't be so easy this time. The next electron to be removed is a $2p^6$ electron. Removing this electron takes a great deal more energy than removing the first electron, because we are destroying the ion's stable octet electron arrangement. Also, we are trying to remove an electron from a positively charged ion.

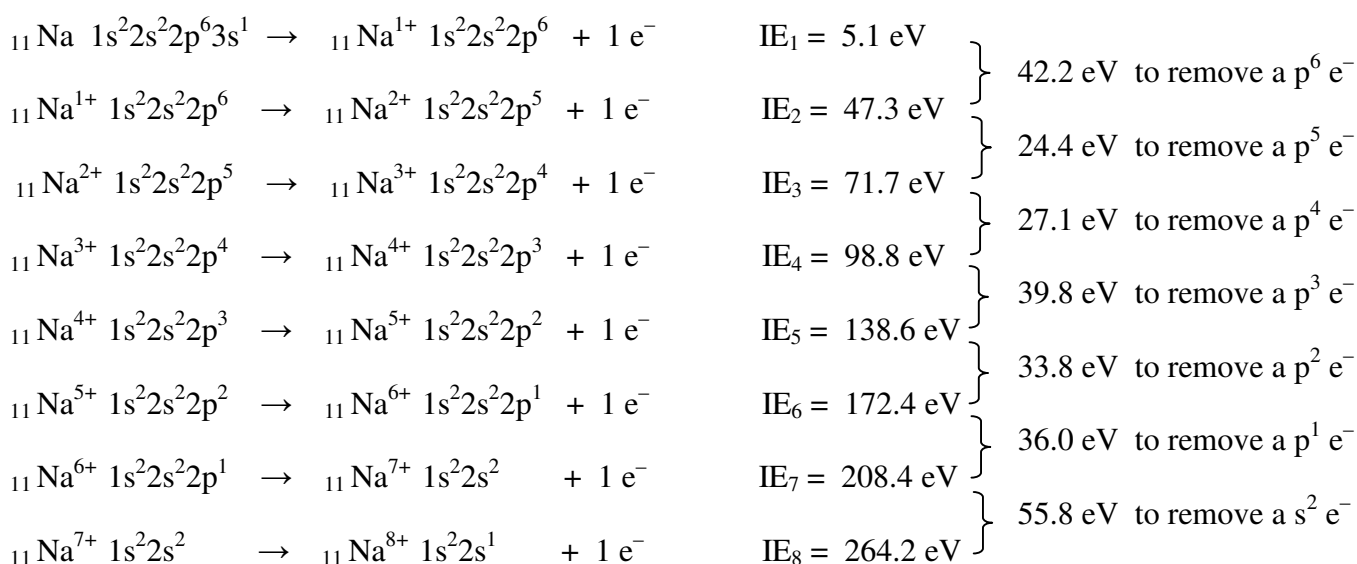


It takes almost 10 times more energy to remove the second electron from a sodium ion than it did to remove the first!!

The amount of energy required to remove a second electron from an atom is called its **second ionization energy** (IE_2). The second ionization energy is always higher than the first because it is harder to remove an electron from a positive ion. If the second ionization involves the destruction of a stable octet, then it will be dramatically higher than IE_1 .

Multi-electron atoms can have second, third, fourth etc ionization energies until all of their electrons have been removed. Each successive ionization requires increasing amounts of energy because you are removing an electron from increasingly positively charged ions.

The magnitude of the IE also depends on which electron is being removed.



From the example above, you can see that:

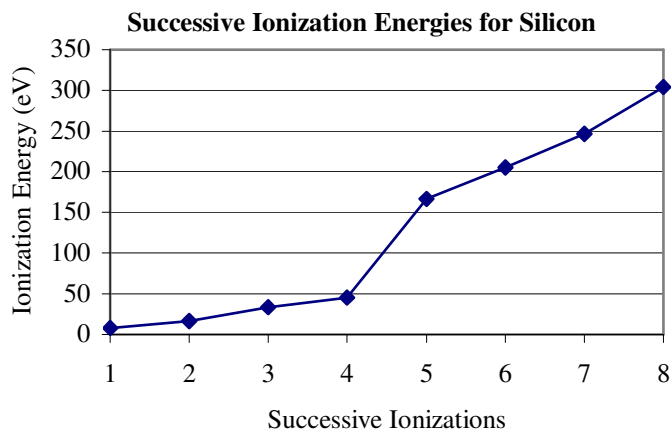
- it takes a great deal of energy to remove an electron from a stable octet ($s^2 p^6$)
- it takes more energy to remove a p^3 electron than a p^4 electron

This is because of the special stability associated with full and half-full orbitals.

Similarly:

- it takes more energy to remove a d^5 electron than a d^6 electron
- it takes more energy to remove a d^{10} electron than a d^9 electron

If you look at a graph of the pattern of successive ionization energies for silicon, you can identify which electrons are being removed.



Notice two things:

1. each subsequent ionization energy is higher (you are removing an electron from more and more positively charged ions) and
2. there is a huge jump in ionization energies from IE_4 to IE_5 . Removing the 5th electron is significantly more difficult than removing the 4th electron, so removing the 5th electron must be breaking a stable octet electron arrangement.

The pattern of ionization energies for each element tells us a great deal about the stability of the various arrangements of electrons for that element.