Teacher's Tools® Chemistry
 Atoms and electrons: Periodic Trends: Student Review Notes

The Periodic Table is an organization of elements by chemical similarity.

Chemical properties of the members of a group are similar—that's what defines the groups. You should understand that this similar reactivity comes from the fact that members of a group have the same valence electron configuration and chemical reactivity has a lot to do with valence electrons.

Periodic Law is the idea that physical properties change in a regular fashion as you move across a period. There are a couple of general ideas you can use to reason out trends on the periodic table as you move across a period.

I. Shielding and Effective Nuclear Charge

You can predict some periodic trends from a simple theory called **shielding**. It is based on the fact that **opposite charges attract.** Negatively charged electrons are attracted to the positively charge nucleus.

The force of attraction between two oppositely charged particles is give by the formula:

$$F = \frac{kq_1q_2}{r^2}$$

This means that the strength of the attraction is proportional to the product of the two charges (goes up when the product of q_1 and q_2 gets bigger) and inversely proportional to the square of the distance separating the charges (goes down as the distance, r, between the charges increases).

Shielding is the idea that the **outer electrons "see" a nuclear charge that is a fraction of the actual nuclear charge** and it is that fraction that determines the strength with which the outer electrons are attracted to the nucleus. This is called the **Effective Nuclear Charge or Z_{eff}**. Inner electrons "**shield**" much of the positive charge of the protons from the outermost electrons. You can visualize shielding in terms of a simple planetary arrangement of the sublevels (remember however that the orbitals on sublevels have non-spherical shapes). This will allow you to calculate a value of Z_{eff} upon which to determine the trend of the property in question. For example, let's use a simple shielding model to calculate Z_{eff} for the outmost electrons in argon:

Each of the outer electrons has a charge of -1.
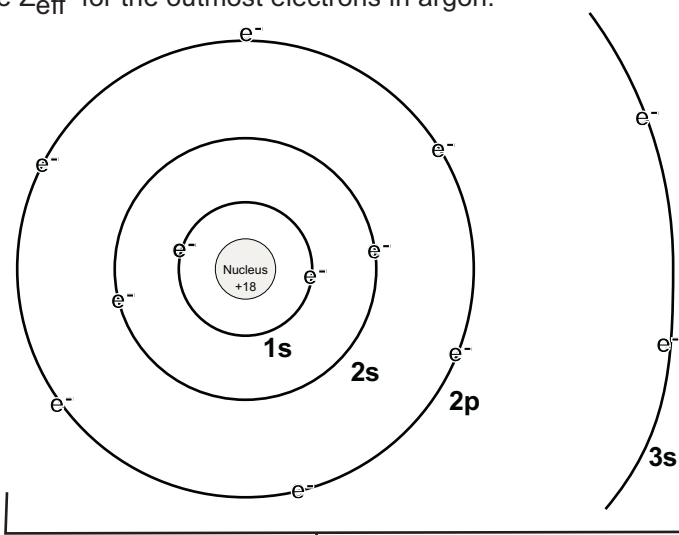
The inner electrons "shield" or cancel out positive charges in the nucleus. In this case, a simple shielding model would estimate 12 positive charges shielded by the 12 inner electrons.
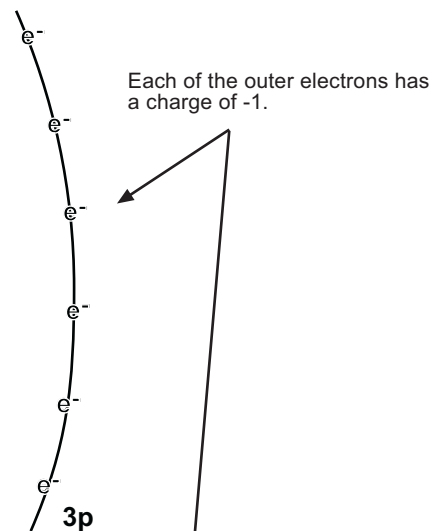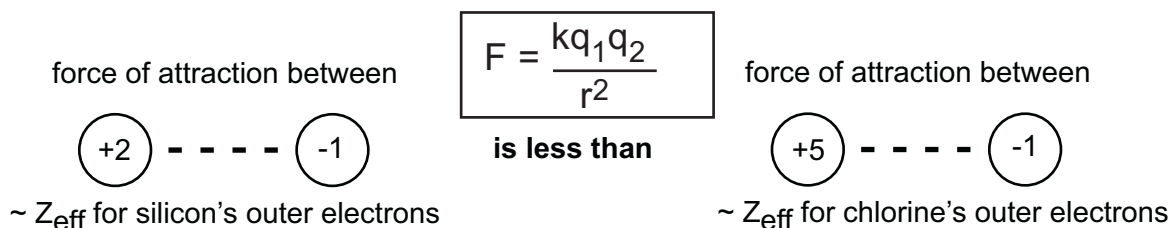
$Z_{eff} \sim$ (+6) · (-1)

Understand that Z_{eff} does not exactly equal the charge of the number of protons minus the number of inner electrons. The true value has to be experimentally determined but, with a simple estimate you can mimic the correct trend in Z_{eff} and therefore predict periodic trends. One other thing, simple shielding explanations break down when you are adding electrons to the d or f sublevels. Orbitals on these sublevels are not as "penetrating," i.e. there is less electron density close to the nucleus than there is in the case of orbitals on s and p sublevels. Subsequently d and f electrons don't shield very much and this is why the radii of the transition metals don't change very much across a period.

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Atomic Radius

Shielding explains this very well. As you move across a period, Z_{eff} increases and the outer electrons are therefore more strongly attracted to the nucleus. For example, take a look at silicon and chlorine. The outer p-electrons of Si ($[Ne]3s^23p^2$) "see" a Z_{eff} of +2 using a simple shielding model while the outer p-electons of Cl ($[Ne]3s^23p^5$) "see" a Z_{eff} of +5.

force of attraction between

$$F = \frac{kq_1q_2}{r^2}$$

is less than

force of attraction between

(+2) – – – – (-1)   (+5) – – – – (-1)

~ Z_{eff} for silicon's outer electrons   ~ Z_{eff} for chlorine's outer electrons

So, the outer electrons of Cl are pulled closer to the nucleus than the outer electrons in silicon because of the greater electrostatic force of attraction. As you move across a period, **Z_{eff} increases** and therefore atomic radius decreases. You don't need to use the idea of shielding to explaing the trend in atomic radius as you go down a group. In this case, energy levels are being added and that increases the size of the atom. Here is the general trend.

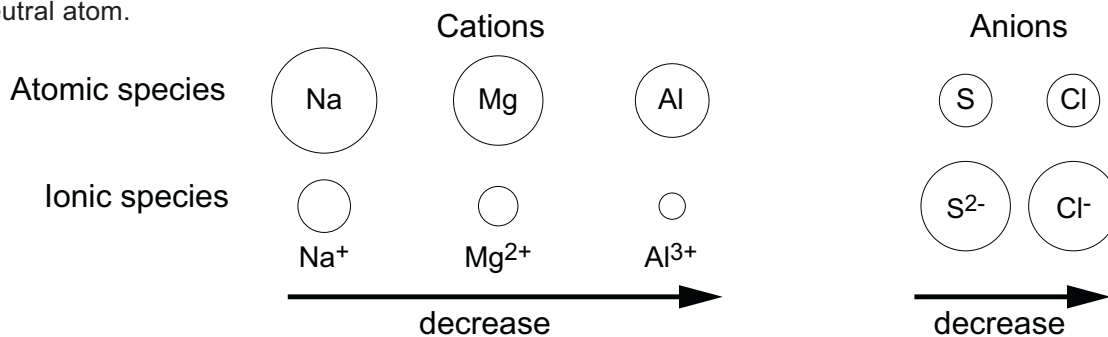decreases

Atomic Radius

increases
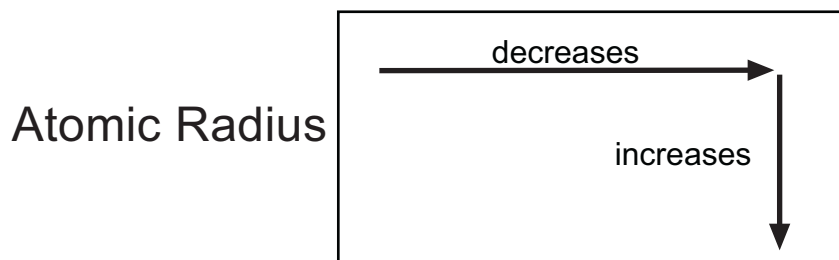
Again, understand that simple shielding explanations break down when you are adding electrons to the d or f sublevels. Orbitals on these sublevels are not as "penetrating," i.e. there is less electron density close to the nucleus than there is in the case of orbitals on s and p sublevels. Subsequently d and f electrons don't shield very much and this is why the radii of the transition metals don't change very much across a period.

Ionic Radius

Shielding doesn't help much here but the basic fact that opposite charges attract and like charges repel can be used to explain trends in ionic radii.

Cations are ions that have lost electrons and are thus left with a net positive charge. During the process of ionization, electrons are lost and this is often all of the electrons on sublevel. The outer electrons that are left see a greater effective nuclear charge than the outer electrons in the neutral atom and thus the ionic radius is smaller than the atomic radius.

Anions are ions that have gained electrons and are thus left with a net negative charge. The effective nuclear charge doesn't change for the outer electrons and electrons don't like each other. The outer energy sublevel therefore has to expand to accomidate the repulsion of the electrons that have been added and therefore the ionic radius is greater than that of the neutral atom.

Cations   Anions

Atomic species   (Na)   (Mg)   (Al)   (S)   (Cl)

Ionic species   ( )   ( )   ( )   (S^{2-})   (Cl^-)

Na^+ Mg^{2+} Al^{3+}

decrease   decrease

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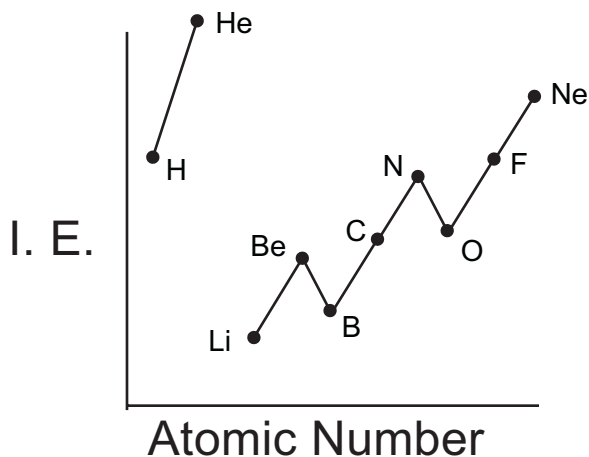
Ionization Energy

Ionization energy is energy necessary to remove the most loosely bound electron from an atom in its gaseous state. The first ionization energy is that required to remove the first electron, the second ionization energy is for the second electron and so on.

Generally speaking, electrostatic attraction and a simple shielding model is a good way to explain the trend of ionization energy on the periodic table. As Z_{eff} increases increases from left to right across a period, the electrostatic charge between the nucleus and the outer electrons increases. Thus ionization energy increases from left to right across a period. Down a group, the outer electrons become further away from the nucleus. Remember that the force of attraction is inversely proportional to the square of the distance separating the charges and therefore I.E decreases down a group.

So the general trend is:

Ionization Energy

increases

decreases
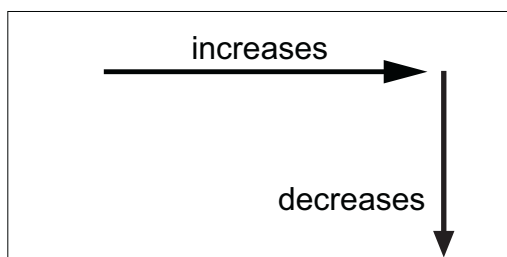
More specifically, there are jumps in ionization energy as you move across a period that are due to the stability of a half-filled sublevel. Take a look.

Half-filled and totally filled sublevels are stable configurations and it takes more energy to remove the outermost valence electron.

Core electrons are the electrons inside of the valence shell. Ionization energy will generally jump by a factor of 10 with the removal of a core electron. Given a table of I.E. values for unknown elements, recognizing this jump and thus the number of valence electrons allows you to assign a group to an element.

Electron Affinity and Electronegativity

Electron affinity is the energy associated with an element gaining an electron in it's gaseous state. The sign convention here can be a little confusing. Elements that want to gain electrons such as the halogens give off energy when they gain an electron so the E.A is negative. Whereas it requires the input of energy to add an electron to elements such as the alkaline earth metals. In this case E.A. will be positive **The trend is that E.A. increases (the more negative the greater) as you go from left to right across a period and decreases as you go down a group.**

Electronegativity is not a physical characteristic of elements. It's a table of relative values developed by Linus Pauling in order to calculate the polar character of chemical bonds. It's similar to the idea of electron affinity, i.e. it has to do with an element's ability to attract electrons except its used in the case of bonds instead of ionization. The trend is the same as for electron affinity. **It increases from left to right across a period and decreases down a group.**