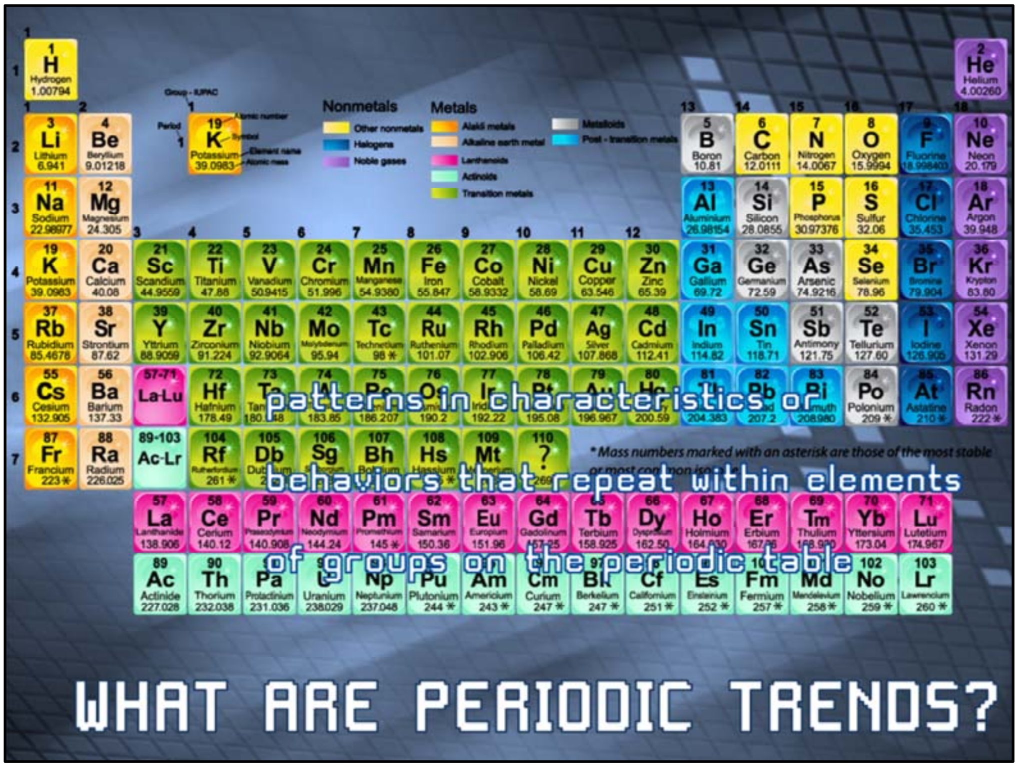
Module 3: Modern Atomic Theory, Electron Structure, and Periodicity
Topic 6 Content: Periodic Trends Presentation Notes

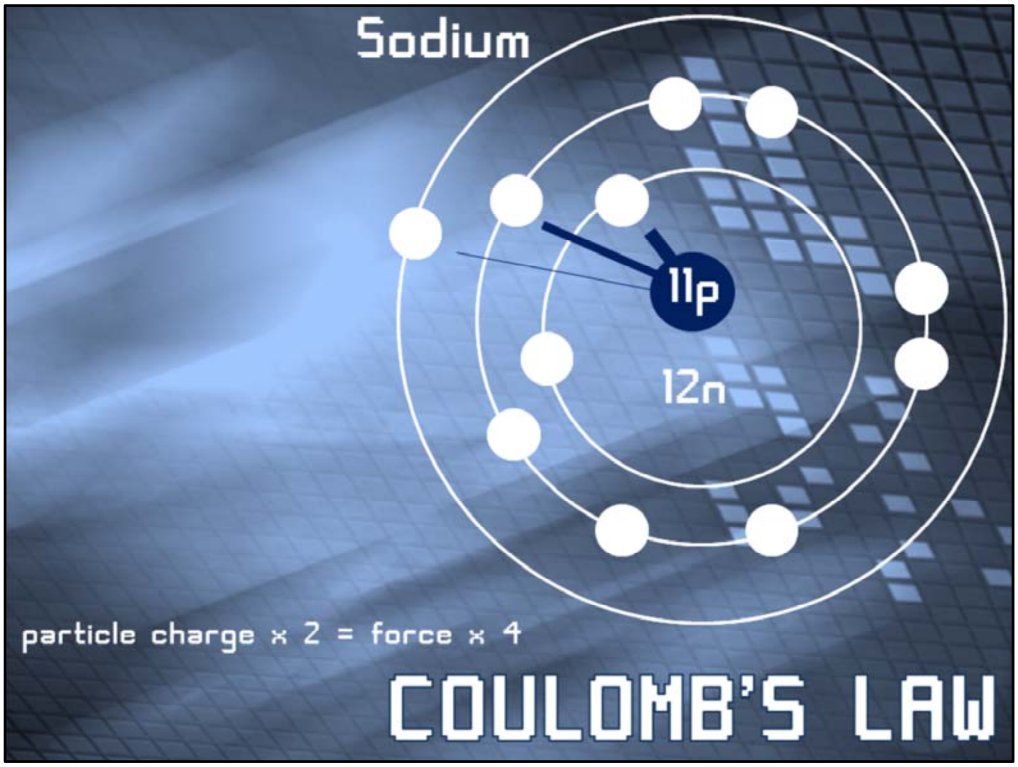
Periodic Trends

Module 3: Modern Atomic Theory, Electron Structure, and Periodicity

Topic 6 Content: Periodic Trends Presentation Notes



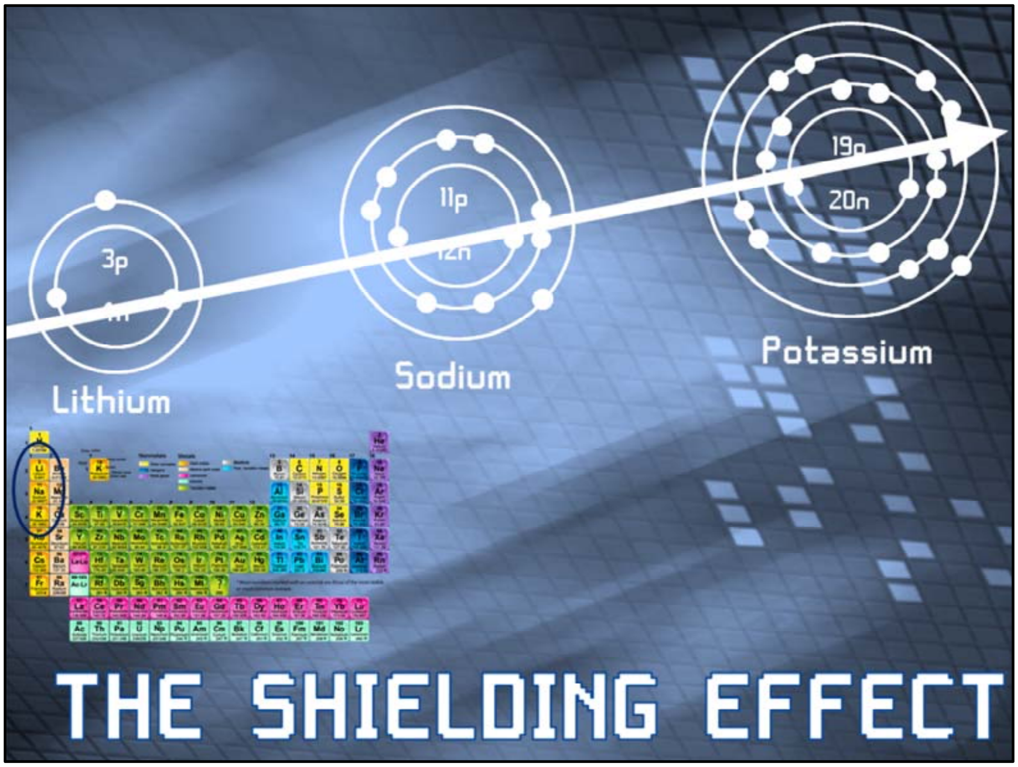
Periodic trends, or periodicity, are patterns in characteristics or behaviors that repeat within elements of groups on the periodic table. When all of the known elements are placed on the periodic table in the order of increasing atomic number, scientists can observe these trends. There are a variety of explanations for these trends, including Coulomb's Law, the shielding effect, and the principle quantum number.



Coulomb's Law relates the distance between charged particles to the forces of attraction or repulsion between them. In the case of the electrons outside of the nucleus, the force is an attractive force since the particles are oppositely charged. According to Coulomb's Law, as the distance between the particles increases, the attractive forces are minimized.

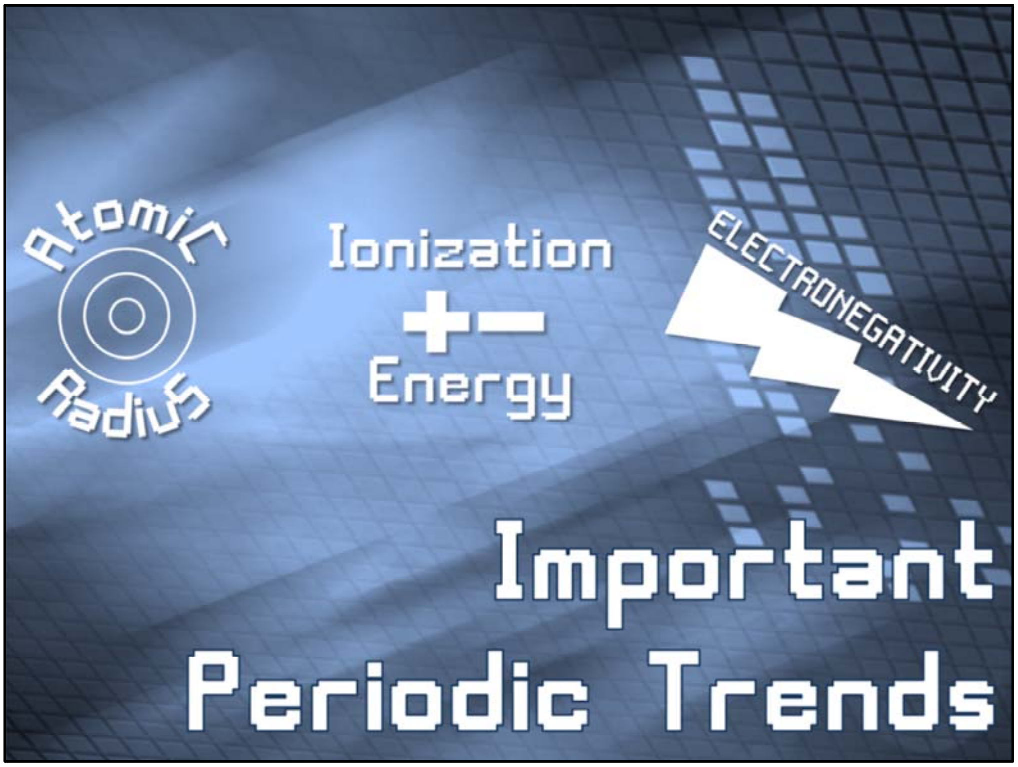
Coulomb's Law also helps to clarify the relationship between nuclear charge and the forces of attraction. According to Coulomb's Law, if the charges of two particles are both doubled, then the force is magnified four times. An increase in the attractive force is expected as more protons are found in the nucleus. The increase in nuclear charge increases the attractive force on the electrons.

Module 3: Modern Atomic Theory, Electron Structure, and Periodicity
Topic 6 Content: Periodic Trends Presentation Notes

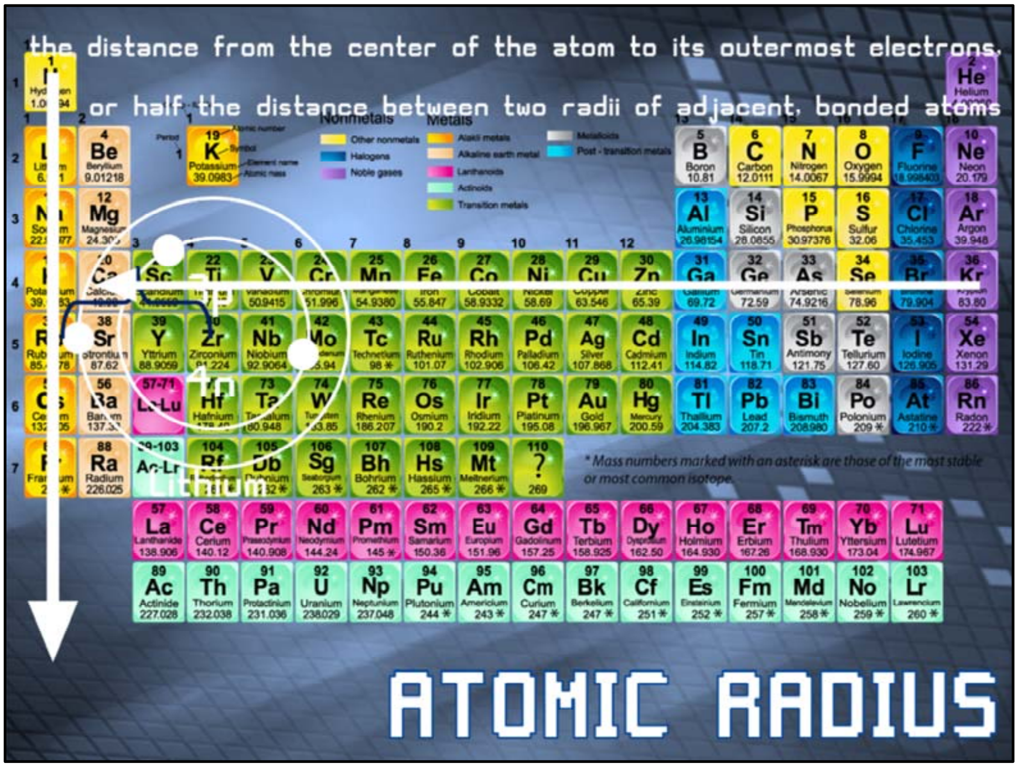


Look at the three different Bohr models of lithium, sodium, and potassium shown here. All three are alkali metals in group one of the periodic table. The core electrons or non-valence electrons are what cause the shielding effect. What is the shielding effect? The shielding effect describes the decrease in attraction between an electron and the nucleus in any atom. This effect will only take place with atoms that have more than one energy level.

Notice that in this case, lithium only has two core electrons and therefore has the lowest shielding effect of the three. Potassium has eighteen core electrons and they present a much higher level of shielding between the protons in the nucleus and the valence electron. Sodium has eleven core electrons. This means that it has a moderate shielding effect when compared to potassium. As the number of core electrons increases, so does the shielding effect.



There are several important periodic trends, including atomic radius, ionization energy, and electronegativity. Click on one of the icons to learn more about a periodic trend. Make sure to click on each of the icons to learn about all of the trends.



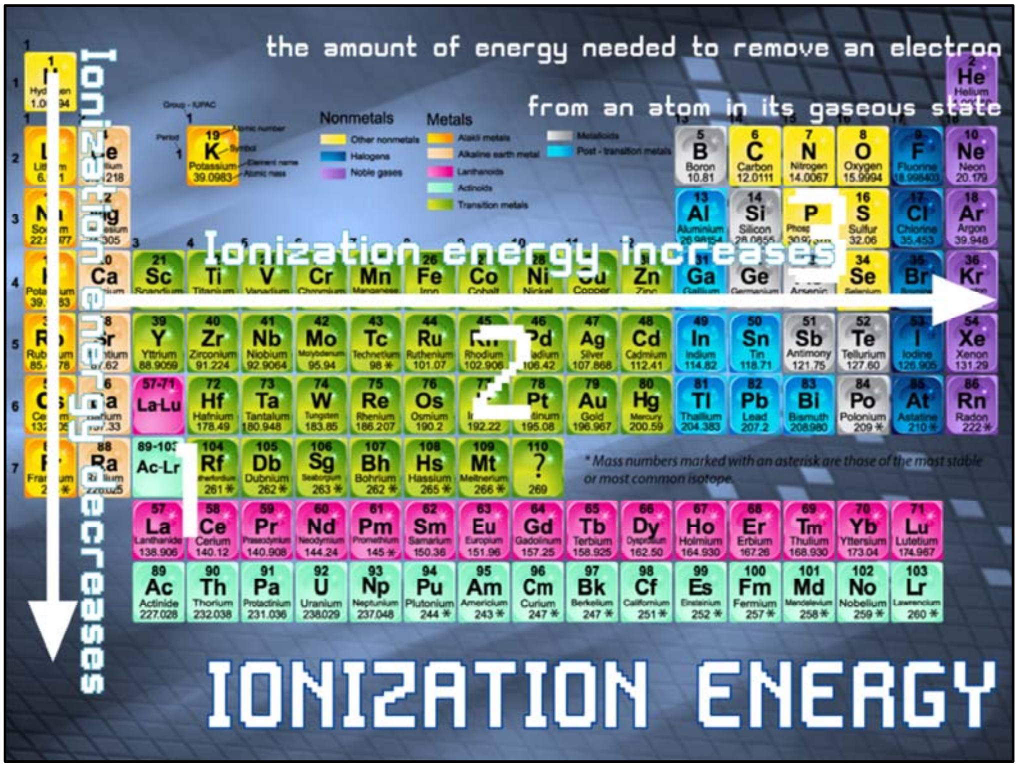
Atomic radius is the distance from the center of the atom to its outermost electrons. It is sometimes described as half the distance between two radii of adjacent, bonded atoms.

Atomic radius is the fundamental trend to which many other characteristics and behaviors are compared. This trend is analyzed in two major directions: the atomic radius increases as you move down a group and decreases as you move across a period. When comparing elements that are in the same group or family, it is easy to visualize why the atomic radius increases. As atoms increase in energy levels, their electrons are a greater distance from the nucleus. It is the period trend that is more difficult. When going across the period, the atoms are gaining protons and electrons, but not gaining a new energy level. This means that the increased nuclear charge is able to pull the outer electrons in closer, causing a decrease in radius.

The image shows the two periodic trends associated with the atomic radii of elements. Notice that as you move down a group, the atomic radius increases. As you move across the period from left to right, the atomic radius decreases.

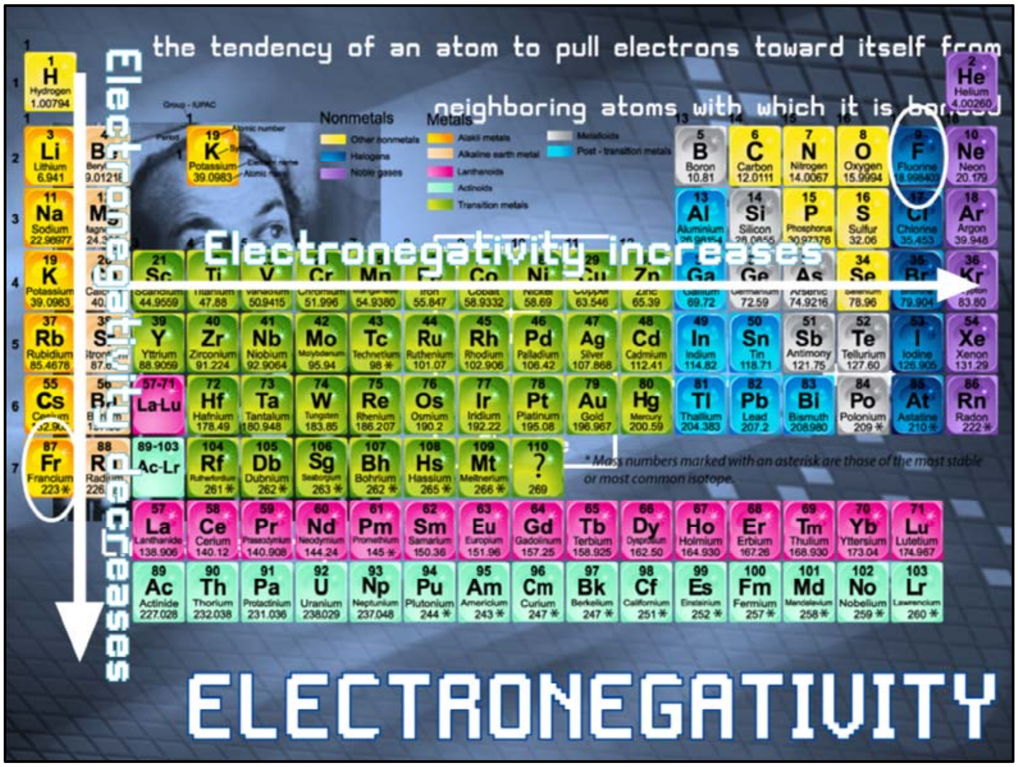
Module 3: Modern Atomic Theory, Electron Structure, and Periodicity

Topic 6 Content: Periodic Trends Presentation Notes



Ionization energy is the amount of energy needed to remove an electron from an atom in its gaseous state. Some elements have multiple valence electrons and there is a unique ionization energy for each of them. The first ionization energy is the amount of energy required to remove the first electron from the valence shell. The second ionization energy is the amount of energy needed to remove the second electron, and the third ionization energy is the amount needed to remove a third electron. Generally speaking, this trend is easiest to understand when related back to the radius. The larger the radius of the atom, the easier it is to remove an electron from the outer energy levels. The smaller the radius of the atom, the more challenging it is to remove an electron. Once an element has lost electrons and has a stable electron configuration, like that of a noble gas, it is extremely difficult to remove any more electrons.

Topic 6 Content: Periodic Trends Presentation Notes



Electronegativity is the tendency of an atom to pull electrons toward itself from neighboring atoms with which it is bonded. In the 1930's, chemist Linus Pauling proposed the phenomena of electronegativity. He assigned a value of 4.0 to fluorine, the most electronegative element. By doing so, Pauling was able to set up relative values for all of the elements. This was when he first noticed the trend that the electronegativity of an atom was determined by its position on the periodic table, and that the electronegativity tended to decrease as you moved down a group and increased as you moved from left to right across a period. The range of values for Pauling's scale of electronegativity ranges from fluorine, the most electronegative, to francium, the least electronegative. Furthermore, if the electronegativity difference between two atoms is very large, then the bond type tends to be more ionic, however if the difference in electronegativity is small then it is a nonpolar covalent bond.

The trend in electronegativity works across the periodic table with the exception of the noble gases. Noble gases have extremely low or no electronegativity. Since noble gases already have stable electron configurations with their octet, they do not attract other electrons toward themselves.