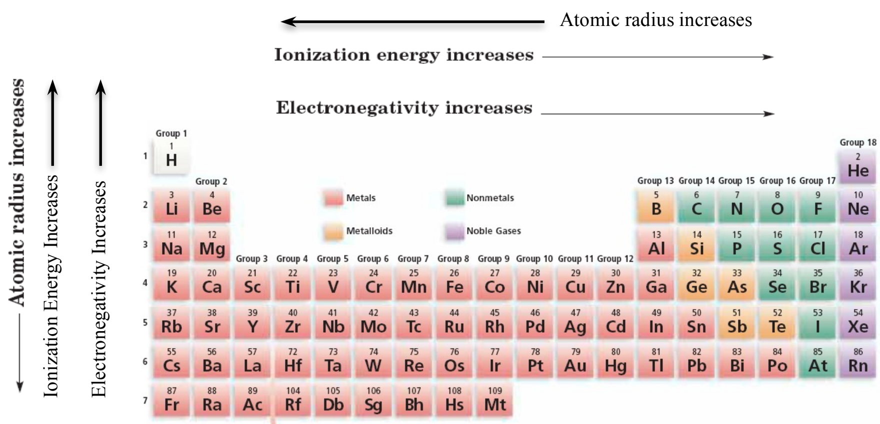
WKS - Honors
Periodic Trends

Name______________________________
Period________ Date___________

TREND Summary:



Answer the following questions:

1) Circle the element in each pair that has the higher ionization energy?
a) K or Li b) He or Rn c) Ne or Li d) Fr or He
(These elements are in the far corners of the table. Use both the across and down trends.)

2) Circle the element of each pair that would have the greater electronegativity:
a) K or Ca b) Li or N c) S or Se d) F or I e) Li or F

3) Circle the atom of each pair that would have the smaller radius.
a) Na or Cs b) Na or Cl c) Na or Al d) F or I

4) As a comparison, a few days ago we discussed how sizes change when one goes from a neutral atom to an ion. Circle the one of each pair that would have a smaller radius:
(Be aware: The number of protons are not changing. Electrons are either lost or gained.)
a) F or F^- b) Na or Na^+ c) Ca or Ca^{2+} d) N or N^{3-}

5) Hydrogen is normally shown in column I with the alkali metals on the periodic table. However, hydrogen's properties are not similar to the properties of metals.
Its radius is too __small__. Its ionization energy is too __high__ Its electronegativity is too __high__

6) In general, which elements tend to have lower electronegativities? **metals or nonmetals?** __metals__

7) In general, which elements tend to have higher electronegativities? **metals or nonmetals?** __nonmetals__

8) Compounds often form between metals and nonmetals. When forming these compounds one element will lose electrons (form positive ions) and the other will gain electrons due to their differences in electronegativity.
   a. The metals __lose__ electrons and form __positive__ ions because they __weakly__ attract their valence electrons.
   b. The nonmetals __gain__ electrons and form __negative__ ions because they __strongly__ attract electrons.

** Forming these compounds is very favorable because the ions formed are very strongly attracted to each other (because of strong electromagnetic attractions.)

9) Within a compound, what is the most common charge for an ion in the alkali family? __+1__
10) Within a compound, what is the most common charge for an ion in the halogen family? __−1__

11) Within a compound, what is the most common charge for an ion in the alkaline earth family? +2
12) Within a compound, what is the most common charge for an ion in group V? −3

13) Sometimes, **two nonmetals** will react to form a compound. Since the electronegativities of both nonmetals are high, the elements both attract electrons strongly and so they must share electrons. However, the elements do not always share electrons equally.
   a. For example, when carbon and oxygen bond together by sharing electrons, which atom would attract the electrons more strongly—C or O ? O
   b. When phosphorus bonds with chlorine, which attracts electrons more strongly—P or Cl? Cl

14) Why don't noble gases have electronegativity values? ns^2np^6 configuration has no attraction for e^-
15) What three elements have the highest electronegativities? N O F

Review of Trends—definitions and explanations

16) The valence electrons in an atom with a high effective nuclear charge are **(weakly, strongly)** attracted to the nucleus.
17) The shielding electrons are the **(inner core, outer valence)** electrons.
18) Calculate the effective nuclear charge for the following elements: (Show calculation.)
   a. Ca $Z_{eff} = 20\ p^+ - 18$ core $e^- = 2+$ b. S $Z_{eff} = 16\ p^+ - 10$ core $e^- = 6+$
19) The larger the atomic radius, the **(larger, smaller)** the size of the atom.
20) The size or radius of an atom mainly depends on the space **(the nucleus takes up, the electrons take up)**
21) The higher the ionization energy, the **(harder, easier)** it is to remove an electron from an atom.
22) The higher the electronegativity of an atom, the **(more, less)** the atom attracts electrons to itself.
23) Why do the radii of atoms increase as one goes down a family?
(Make sure to discuss both the effective nuclear charge and the number of main energy levels.)
Coming down a group/family, while the effective nuclear charge remains constant (Z goes up but so does the number of core electrons). The valence electrons enter orbitals that are in higher main energy levels and, on average, farther from the nucleus. Since the radius is the distance from the nucleus to the "edge" of the valence electrons, the radius must increase.

24) Why do the radii of atoms decrease as one goes across to the right in a period?
(Make sure to discuss both the effective nuclear charge and the number of main energy levels.)
Moving L → R across a period, the number of protons increases while the core shielding electrons remain constant, so Z_{eff} increases. The valence electrons are all in the same main energy level, so their attraction to the nucleus increases. This stronger attraction pulls them closer to the nucleus, decreasing the distance from the nucleus to the "edge" of the valence electrons and, thus, the radius.

25) Why do smaller atoms have higher ionization energies?
Electrons in smaller atoms are closer to the nucleus, thus feel a stronger attraction. Also, smaller atoms tend to have a higher Z_{eff}, also increasing the strength of the attraction of the electrons to the nucleus. The stronger attraction makes it more difficult to remove the electron, thus higher ionization energy is required.

26) Why do smaller atoms have higher electronegativities?
In smaller atoms, electrons are pulled closer to the nucleus and may also experience a higher Z_{eff}. Thus, they are more strongly attracted to the nucleus, increasing the electronegativity.