Name:______________________________ Block:______ Date:_____________

Chemistry 11 **Trends Activity** Assignment

Atomic Radius: the distance from the center of the nucleus to the outer most electrons in an atom.

Ionic Radius: the distance from the center of the nucleus to the outer most electrons in an ion.

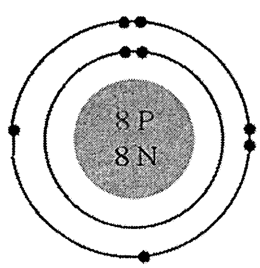
Ionization Energy: the energy needed to remove an electron.
First ionization energy is the energy needed to remove the first electron from an atom. The first ionization energy is always the lowest.

Electronegativity: the ability to attract an electron in a chemical bond.
Most commonly measured with the Pauling Scale, where 0 represents the least ability and 4 is the greatest ability to attract electrons in chemical bonds.

Electron Affinity: the energy change associated with gaining an electron.

Effective Nuclear charge (Z_{eff}): a measure of the pull of the protons in the nucleus on the valence electrons in an atom/ion.
Z_{eff} = atomic number – shielding electrons
Note: shielding electrons for our purposes are the inner electrons that block some of the pull of the nucleus on the electrons that are farther out.

Example: Oxygen



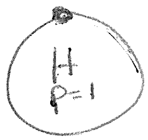
Z_{eff} = atomic number – shielding electrons
$Z_{eff} = 8 - 2$
$Z_{eff} = +6$

Use the provided periodic table to answer the questions below.

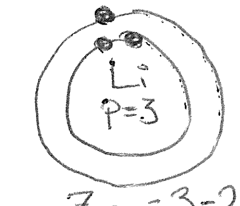
1. a) As you go from hydrogen down a column, what happens to the atomic radius?

   Increases/decreases

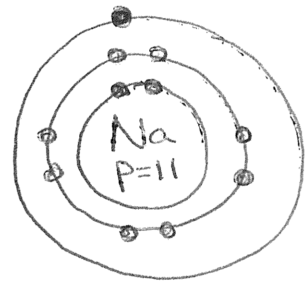
   b) Draw Bohr diagrams for hydrogen, lithium and sodium. Calculate the effective nuclear charge for each atom and use this information to explain the atomic radius trend.



$Z_{eff} = 1 - 0 = +1$



$Z_{eff} = 3 - 2 = +1$



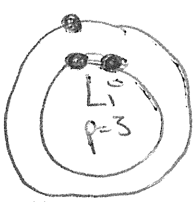
$Z_{eff} = 11 - 10 = +1$

Each atom has one more layer of electrons, which are farther from the nucleus and make the atom larger.

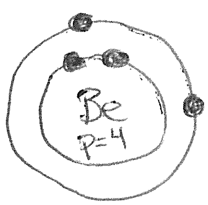
2. a) As you go from lithium across a row what happens to the atomic radius?

   Increases/decreases

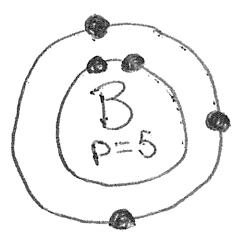
   b) Draw Bohr diagrams for lithium, beryllium and boron. Calculate the effective nuclear charge for each of the three elements and use these values to explain the atomic radius trend.



$Z_{eff} = 3 - 2 = +1$



$Z_{eff} = 4 - 2 = +2$



$Z_{eff} = 5 - 2 = +3$

The effective nuclear charge increases and so the nucleus pulls the outer electrons in closer, making the atom smaller.

3. What unit is ionization energy measured in? kJ (kilojoules)

4. a) As you go from hydrogen down a column, what happens to the ionization energies?

   Increases/decreases

   b) Using your Bohr diagrams from question #1, explain this trend.

   The valence electrons are farther from the nucleus so there is less attraction from the nucleus and so it takes less energy to remove an electron

5. a) As you go from lithium across a row what happens to the ionization energies?

Increases/decreases

b) Using your Bohr diagrams from question #2, explain this trend.

The nuclear charge is increasing, which means the nucleus is more strongly attracting the electrons so it takes more energy to remove the electrons as you go across a row.

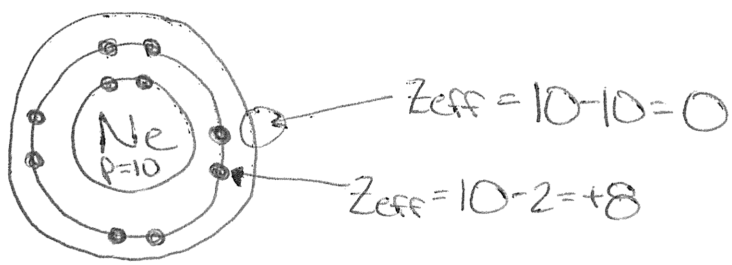
6. Which family of elements has the highest ionization energy? Noble gases

7. Which family of elements has the highest electronegativities? Halogens

8. a) Which family of elements doesn't have electronegativity values listed?

Noble gases

b) Explain why this family wouldn't have electronegativity values (Hint: draw the Bohr diagram for Neon and calculate the Z_{eff} to help you answer this.)



To attract an electron, the new electron would need to go in the next shell where Zeff = 0. This means the nucleus isn't attracting a new electron

9. a) Which property follows the same trend as ionization energy? Electronegativity

b) Explain why these two trends should be connected.

If an atom is good at attracting other electrons (high electronegativity) it would make sense that it is also good at holding onto the electrons it already has (high ionization energy)

10. a) Compare the atomic radii for metals and the ionic radii for their positive ions. Are the radii for the ions or atoms larger?

- positive ions (metal ions) are much smaller (have smaller ionic radii) than their metal atoms

b) Explain the above trend

- the positive ions have lost electrons, so are smaller (loose an entire shell!)

11. a) Compare the atomic radii for non-metals and the ionic radii for their negative ions. Are the radii for the ions or atoms larger?

- negative ions (non-metal ions) have much larger

b) Explain the above trend

- the ions gained electrons so the electrons repel eachother more and spread out making the ion much larger than the atom

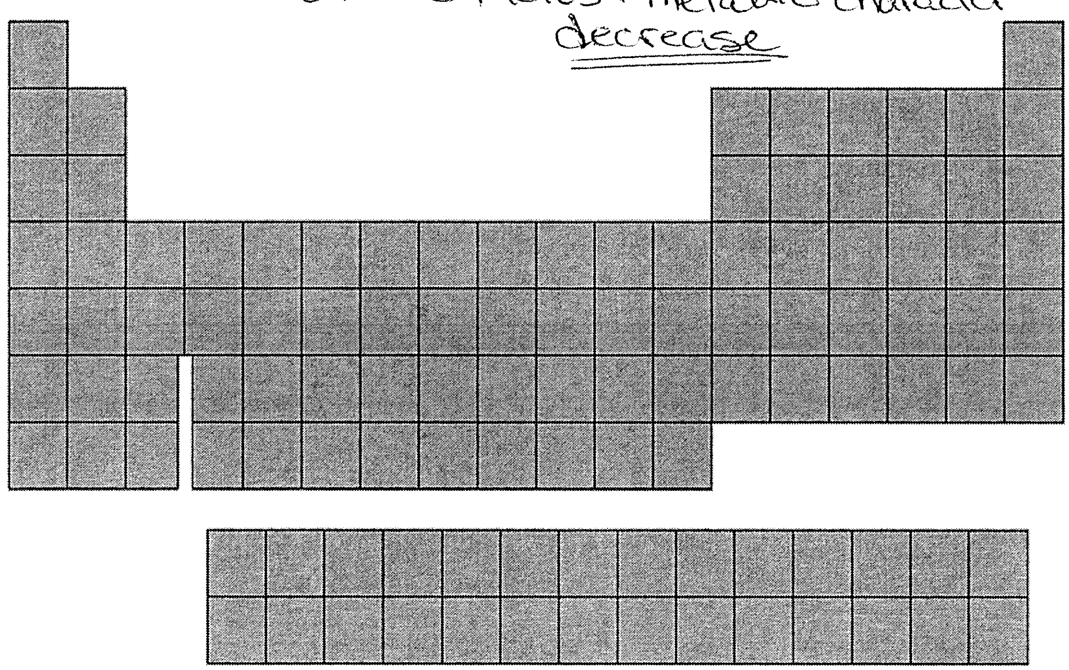
12. On the periodic table below, draw arrows on the sides (going up or down and left or right) to summarize the trends of atomic radius, ionization energy, electronegativity, and metallic character.

ionization energy + electronegativity increase

atomic radius + metallic character decrease

Ionization energy + electro-negativity increase

atomic radius + metallic character decrease



Do the next 3 questions using your own periodic table – based on the trends – not the values listed on the provided periodic table.

13. Indicate which atom in each pair would have larger atomic radius.

a. Li or K
b. Ca or Ni
c. Ga or B
d. O or C
e. Cl or Br
f. Be or Ba
g. Si or S
h. Fe or Au

14. Indicate which ion in each pair would have smaller ionic radius.

a. K^+ or O^{2-}
b. Ba^{2+} or I^-
c. Al^{3+} or P^{3-}
d. K^+ or Cs^+
e. Fe^{2+} or Fe^{3+}
f. F^- or S^{2-}

15. Indicate which atom or ion in each pair would have larger ionization energy.

a. Na or O
b. Be or Ba
c. Ar or F
d. Cu or Ra
e. I or Ne
f. K or V
g. Ca or Fr
h. W or Se

16. Explain why there would be a large jump in ionization energy between the second and third ionization energies for magnesium. (Draw a Bohr Diagram to help you.)

The third ionization energy would refer to removing an electron from a full orbit that is closer to the nucleus, so it would be very difficult