

## Notes#24 Periodic Relationships#2/AP Chemistry

- The general trend for atomic radii holds very true for the representative elements. However, some variation is seen in the d-block and f-block. (look at P.T. Atomic Radius image)

### PRACTICE:

1. Arrange the following atoms in order of increasing atomic radii: P, Si, N
2. Arrange the following atoms in order of increasing atomic radii: Ne, Mg, P

**B. TREND 2: IONIC RADII** - Because atoms and ions of the same element have different numbers of electrons, we should expect atomic radii and ionic radii to have different values....

- **THE TREND** - *The radii of cations are \_\_\_\_\_ than those of the corresponding neutral atom.* Why? Take a look at what happens when you take away an  $e^-$  from Na.

\*\* You lose \_\_\_\_\_. This makes a cation significantly smaller.

\*\* As the (+) charge increases, the cation gets smaller.  $Mg^{2+} < Na^+$ . Why? These two ions have the same  $e^-$  configuration (isoelectronic) but  $Mg^{2+}$  has \_\_\_\_\_  $p+$  pulling on the remaining electrons and  $Na^+$  only has \_\_\_\_\_  $p+$  pulling on the  $e^-$

- *The radii of anions are \_\_\_\_\_ than those of the corresponding neutral atom.* Why? When an  $e^-$  is added, there is an \_\_\_\_\_ in electron-electron repulsion.

### PRACTICE:

1. Circle the smaller ion in each pair and explain your reasoning:

$K^+$  and  $Li^+$

$Au^+$  and  $Au^{3+}$

$N^{3-}$  or  $F^-$

2. Put the following ions in order of increasing atomic radii and EXPLAIN.  $Cl^-$ ,  $Ca^{2+}$ ,  $S^{2-}$ , Ar,  $K^+$

**C. TREND 3: IONIZATION ENERGY** - the amount of energy (in \_\_\_\_\_) it takes to \_\_\_\_\_ one mole of  $e^-$  from the ground state of one mole of atoms in the gas phase.

\*\* In other words, the magnitude of the ionization energy is a measure of how \_\_\_\_\_ the outermost electron is held by an atom. The \_\_\_\_\_ an  $e^-$  is held, the more energy is required to remove it (a.k.a....a \_\_\_\_\_ IONIZATION ENERGY).

\*\* Ionization energy is always an \_\_\_\_\_ process and a ( ) energy value.

-**THE TREND** - *As you move across the periodic table from left to right in an  $n$ -shell, ionization energy \_\_\_\_\_.* \*\* It becomes harder and harder to remove electrons as you go across because.....

- *As you move down the periodic table from low  $n$  shells to higher  $n$  shells, ionization energy \_\_\_\_\_.* \*\* It becomes easier and easier to remove electrons as you go down because.....

\*\* Notice the BIG correlation between the SIZE of the atom, the attractive forces from the nucleus and the Ionization energy.....

**EXCEPTIONS** in the Ionization Energy trend. Most of the following exceptions are driven by the fact that there is that element of \_\_\_\_\_ to a full or half filled sublevel.

a. **IRREGULARITY** between 2A (alkaline E. metals) and 3A (Boron) families:

- The ionization energy for Be is > than that for B. Why?

Consider the  $e^-$  conf, draw out orbital diagrams

Be: [He]

B: [He]

\*\* Less energy is required to remove a single p  $e^-$  than a paired s electron.

There is some stability associated with a FULL s orbital.

\*\* Is this trend pretty much consistent all the way down the P.T.? \_\_\_\_\_

b. **IRREGULARITY** between 5A (nitrogen) and 6A (oxygen) families.

- The ionization energy of N is > than that for O. Why?

Consider the  $e^-$  conf, draw out orbital diag.

N: [He]

O: [He]

\*\* It's easier to remove a single  $e^-$  from the O atom than from the N atom where the p orbital is  $\frac{1}{2}$  full.

\*\* Is this trend pretty much consistent all the way down the P.T.?? \_\_\_\_\_

**Moving beyond the first Ionization Energy. Second, Third, Fourth, etc. ionization energies.**

\*\* Primarily, we have been taking about the FIRST ionization energy,  $I_1$ , (energy needed to remove the FIRST electron). The SECOND ionization energy,  $I_2$ , would be the energy needed to remove a \_\_\_\_\_ electron and so-on.

1. **THE TREND** -  $I_1 < I_2 < I_3$  ..... Why do you suppose this is?

2. **PRACTICE:**

Which of the following atoms should have a smaller second ionization and WHY? Mg or Na?

D. **TREND 4: ELECTRON AFFINITY** - the energy change (kJ/mol) that occurs when a mole of electrons are accepted by a mole of atoms in the \_\_\_\_\_ to form an anion.

1. Overall, electron affinity is a measure of how stable an atom becomes upon \_\_\_\_\_ an electron.

Ex. Let's look at Fluorine.  $F(g) + e^- \rightarrow F^-$   $\Delta H^\circ = -328 \text{ kJ/mol}$

This should make sense because fluorine is just one  $e^-$  away from having a noble gas configuration and really does "want" an  $e^-$  and resultantly, it becomes much more \_\_\_\_\_ upon accepting one.

\* 2. Just remember, the larger the negative number for electron affinity, the \_\_\_\_\_ an element "wants" an electron. If an element has a positive electron affinity, that means that they are really \_\_\_\_\_ upon accepting an electron.

3. **THE TREND** - Not a really clear-cut trend going across or down the P.T.

Although some trends are summarized below:

(a) The E.A. of groups \_\_\_\_\_ and \_\_\_\_\_ (with \_\_\_\_\_ the greatest) are much GREATER than any other elements.

(b) The E.A. of the noble gases are slightly \_\_\_\_\_. Therefore, the anions of these gases are very \_\_\_\_\_.

(c) The group 1A metals have small negative E.A. And, the group 2A metals have even smaller negative EA values with some positive values. Why do you suppose this is?

(d) The electron affinities of the metals are generally \_\_\_\_\_ than those of the nonmetals.