

## CH 7 WORKSHEET

CHAPTER 7: 7.2-7.5

Name: \_\_\_\_\_ Per: \_\_\_\_\_ Date: \_\_\_\_\_

7.2 Effective Nuclear Charge      7.3 Sizes of Atoms and Ions      7.4 Ionization Energy      7.5 Electron Affinities

1. Atomic Size: The diameter of an atom is measured in the unit of the \_\_\_\_\_ which has the symbol of \_\_\_\_\_. This unit is equal to \_\_\_\_\_ cm or \_\_\_\_\_ meters which is \_\_\_\_\_ nm or \_\_\_\_\_ pm. The current practice is to express atomic and molecular dimensions in the unit of \_\_\_\_\_.

**Effective nuclear charge:** The key to understanding the trends in the properties in this section is the concept of effective nuclear charge. For example, the electron configuration of Li is  $[\text{He}] 2s^1$ . The single electron of Li is shielded from the full effect of the nuclear charge of  $3^+$  by the  $2^-$  charge of the helium core which lies between the nucleus and the valence shell. If the shielding were perfect, this electron in the valence shell would experience a *net* charge - the **effective nuclear** - of only  $1^+$ . In other words, if the core shielded the valence electron completely, the effective nuclear charge felt by the valence electron would be the difference between the nuclear charge and the charge on the core.

2. General trends in atomic radius in relation to atomic number shows that as we move across a period on the Periodic Table the sizes of the atoms \_\_\_\_\_ (increase/decrease) and as we move down a group or family the size of atoms \_\_\_\_\_ (increase/decrease). The overall general trend in respect to the periodic table shows that as you move from the \_\_\_\_\_ hand side of the table to the \_\_\_\_\_ hand side of the table atomic size decreases.
3. Explain this trend relating it to the 1) charge on the nucleus, 2) distance from the nucleus and 3) effective nuclear charge.

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**Lanthanide Contraction:** Recalling that the Lanthanide elements, the first row of the inner transition metals fall between La (Group IIIB) and Hf (Group IVB) in Period 6 we must consider the fact that each successive element in this series adds a proton and an electron. The successive addition creates an increase in effective nuclear charge which causes the atoms to become smaller and more dense. Due to this additional decrease in size Hf is approximately the same size as Zr which is above it on the periodic table and many atoms of the 6th period transition elements are about the same size as the atoms above them but have greater mass. This phenomenon is called the Lanthanide Contraction.

4. The atom which is most dense is: a)Zr b)W c)Pd d)Pt and it is most similar in size to e)Hf f)Mo g)Ni h)Pd

**Sizes of Ions:** Particles become smaller when electrons are removed from them and larger when electrons are added. This may be explained as the positive nuclear charge remains constant. For example if Sodium has a positive charge on the nucleus of +11 and in the neutral atom has 11 electrons when it loses one from the outermost shell then it has one more positive charge than negative charge. The nucleus can then pull the electrons closer causing a decrease in the size of the +1 ion. Conversely, in the case of Fluorine with its +9 core when the neutral atom gains one electron then the nuclear pull of the positive charge isn't quite strong enough to hold those electrons quite so tightly. Also, due to the increase in electron(s) in the outermost shell there is an increase in electron-electron repulsions therefore the atomic radius expands slightly.

5. In each pair choose the one with the larger radius:  
 a) Na or  $\text{Na}^+$                       b) Cl or  $\text{Cl}^-$                       c)  $\text{Mn}^{2+}$  or  $\text{Mn}^{3+}$

**Ionization Energy:** The energy needed to remove an electron from an isolated **gaseous** atom, ion or molecule in its ground state usually represented in kJ/mol and by the equation:  $\text{Na}(g) \rightarrow \text{Na}^+(g) + e^-$  This is an endothermic process (energy must be put in) because the electron is attracted to the positive nucleus. (aka: it costs energy to remove an electron.) The first ionization energy is the energy needed to remove the first electron from a neutral atom. Hydrogen only has one electron therefore it is the one element that only has one ionization energy. Periodic trends for ionization energy may be generalized that as size decreases, ionization energy increases. (Diagonally from lower left, to upper right hand corner of the periodic table ionization energy increases.)

6. If you remove an electron then what happens to the attraction of the remaining electrons to the nucleus?

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7. If the attraction increases, then will it require more or less energy to remove the next electron?

As an element loses or gains an electron and becomes closer to the configuration of a noble gas, it becomes more stable and will require more energy to remove another electron from its outermost shell.

8. Among the following, which element has the lowest ionization energy? Na, Cl, Cs or I? \_\_\_\_\_

Now look at trends as you go down the periodic table, recall that as successive shells of electrons are added to the atom, a sort of shielding occurs which diminishes the magnetic attraction of the nucleus to the electron and the radius of the atom increases.

9. Would it take more or less energy to pluck off an electron from the outermost shell of an atom of large radius? \_\_\_\_\_

10. Which element on the periodic table has the highest first ionization energy? \_\_\_\_\_

**Irregularities in Trends Across a Period:** In looking closer at the ionization energies as we move across a period you will notice some irregularities. These may be explained due to the following: 1) It takes more energy to remove an electron from a completed subshell, thus beryllium's ionization energy is higher than boron as beryllium has a completed 2s subshell which would require more energy to remove an electron than from boron with only one electron in the incomplete 2p subshell. 2) There is some stability in a half-filled subshell. Nitrogen has one electron in each of the three orientations of the 2p subshell (Hund's Rule). Oxygen however, has one additional electron paired up in one of the 2p orientations. Due to the electron-electron repulsion of this additional electron (paired spins or not) the ionization energy of oxygen is slightly higher than that of nitrogen.

11. Which atom in this group has the highest first ionization energy: B, C, Al, Si? \_\_\_\_\_

Why? \_\_\_\_\_

**Electron Affinity:** Electron affinity is the energy that is released or absorbed when an electron is added to a gaseous atom or ion in its ground state. It is **usually** associated with an **exothermic** reaction. This is the payoff of receiving the energy associated with gaining an electron. Remember that if the removal of an electron is doing work it increases Potential Energy and is endothermic. Receiving an electron which decreases Potential Energy is exothermic. You can understand this better if you relate this to the fact that as an atom receives electrons and obtains a noble gas configuration it becomes more stable and PE is lowered (exothermic).

Note also that Electron Affinity energy increase parallel's that of Ionization Energy. As we move closer to  $F^-$  we get back more energy as the atom receives an electron. Conversely, the electrons of larger atoms such as Fr feel so little attraction to the nucleus that they have little interest in gaining an electron due to the lack of attraction to the nucleus therefore they lack affinity for the gain of an electron.

Recall also that the variations in the trend are affected by completion of an s subshell where these atoms are content to not gain another electron and also to the half-filled subshell stability of p orbitals. Notice Nitrogen endothermic quantity due to the electron stability of a half-filled p subshell not wanting to receive another electron due to electron repulsion.

**Remember that there are exceptions due completion and half-completions of subshells (more stable)** Also, the small sizes of the atoms in period 2 and the crowding of the electrons in their valence shells causes the electron affinities of these atoms to be less exothermic than the elements below them in period 3.

Based on this information and the fact that chemical activity involves valence shell electrons, which elements are the most active nonmetals? *upper right-hand region of the Periodic Table*

12. Which atom has the least exothermic electron affinity? Si, Al, Ga, Ge \_\_\_\_\_

13. Indicate whether the following reactions are endothermic or exothermic: (See Figure 7.12)



14. What can you summarize about the electron affinities of the noble gases?

Explain: \_\_\_\_\_

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