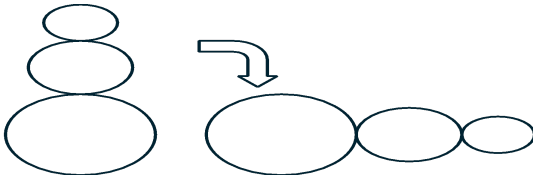
Name:_____ H. Chem, Period _____

Periodic Trends Worksheet

ATOMIC RADIUS

Draw the trend for ATOMIC RADIUS



1. Define Atomic Radius:
   Distance from the nucleus to the outermost orbital of the electron.

2. Define Electron Shielding:
   The reduction in the nuclear charge experienced by the outer electron due to distance and the electron-electron repulsions between the electron of interest and the nucleus.

3. Define Effective Nuclear Charge:
   The actual amount of positive nuclear charge experienced by an electron.

4. What causes the trend in atomic radius you observe as you go down a group/family on the periodic table?
   The atomic radius gets progressively larger as you go down a group because with each period, an energy level or shell is added to the atom.

5. What causes the trend in atomic radius you observe as you go across a period on the periodic table?
   As you go across a period, the atoms get smaller in radius despite having more atomic mass. As the number of protons increase, the nuclear charge increases effectively pulling in the orbitals of electrons closer to the nucleus.

6. Rank each of the following in order of INCREASING atomic radius
   a. Br, F, At _____F, Br, At_____ Explain why:
      The atomic radius gets progressively larger as you go down a group because with each period, an energy level or shell is added to the atom. F has 2 energy levels, Br has 4 energy levels and At has 6 energy levels.

   b. Li, O, F _____F, O, Li_____ Explain why:
      As the number of protons increase, the nuclear charge increases effectively pulling in the orbitals of electrons closer to the nucleus. F would be the smallest with the greatest number of positive protons.

c. K, Na, P, Cl _____Cl, P, Na, K_____ Explain why for each trend if there are multiple trends.
Both trends are present in this example. As you go across a period, from Na to P to Cl, the radius will decrease as the number of protons increase. As you go down a group,

7. Rank each of the following in order of DECREASING atomic radius

a. Cl, Br, Ga _____Ga, Br, Cl_____

b. Ca, Rb, C _____Rb, Ca, C_____

8. Explain what happens to the size of a sodium atom when it loses an electron. Why?
Cations are always smaller than their respective neutral atom. When an atom loses an electron, there are now more protons than electrons giving a positive cation. With the nuclear charge having one more proton, the remaining electrons are pulled in towards the nucleus creating a smaller cation.

9. Explain what happens to the size of a chlorine atom when it gains an electron. Why?
Anions are always bigger than their respective neutral atom. Although, there is research still looking into this, when a negative electron is added into the energy levels with existing electrons, there is electron-electron repulsion, and those electrons want to move away from each other as much as possible. If the energy levels move away from the nucleus to provide more room, the resulting anion is larger than the neutral atom.

10. Place in order from smallest to largest atomic/ ion radius: **Fe Fe^{2+} Fe^{+}** **Fe^{2+}, Fe^{+}, Fe**

11. Place in order from smallest to largest atomic/ ion radius: **Cl^{-}, S^{2-}, Mg^{2+}** **Mg^{2+}, Cl^{-}, S^{2-}**

12. Place in order from smallest to largest atomic/ ion radius: **Na Na^{+} Cl** **Na^{+}, Na, Cl^{-}**

13. Put in order of increasing size (smallest to largest): **K Cl^{-} Ar S^{2-}** **Ar, K, Cl^{-}, S^{2-}**

IONIZATION ENERGY (IE):

Draw the trend for IONIZATION ENERGY



14. Define Ionization Energy:
The minimum amount of energy required to remove the most loosely bound electron in the gaseous state.

15. What causes the trend in IE that is observed as you go down a group or family?
As you go down a group, the atoms get bigger due to increased number of energy levels. As the atom gets bigger, the valence electron gets further away from the nucleus and the effective nuclear charge or attraction the electron feels to the nucleus is reduced. If the nucleus holds onto the electron tightly (in a smaller atom), it will require more energy to remove the electron (high IE). If the electron is held loosely (in a bigger atom), the amount of energy required to remove it will be less (low IE).

16. What causes the trend in IE that is observed as you go across a period?
As you move across a period, the number of protons in the nucleus (nuclear charge) increases while the electron shielding remains relatively constant, resulting in a stronger attraction between the nucleus and the outermost electrons, making it harder to remove them. This is why metals with lower IE readily lose electrons and nonmetals with higher IE prefer not to lose and end up gaining electrons instead. However, this trend is oversimplified and the orbital filling diagrams should be used to determine IE across a period.

17. Rank each of the following in order of INCREASING ionization energy

a. N, O, S, Ge _____Ge, S, O, N_____

b. Mg, Ba, Al, Cl _____Ba, Mg, Al, Cl_____

18. Rank each of the following in order of DECREASING ionization energy

a. Cl, Cu, Au _____Cl, Cu, Au_____

b. Te, Sb, Xe _____Xe, Te, Sb_____

19. The table below gives the ionization energies for **potassium, calcium and bromine**. Identify which element is which from the data given. Explain your answer in the space provided.

ELEMENT 1			ELEMENT 2			ELEMENT 3	
Ionization energy number	**Enthalpy kJ/mol**		**Ionization energy number**	**Enthalpy kJ/ mol**		**Ionization energy number**	**Enthalpy kJ/mol**
1st	418.8		**1st**	1139.9		**1st**	589.8
2nd	3052		**2nd**	2103		**2nd**	1145.4
3rd	4420		**3rd**	3470		**3rd**	4912.4
4th	5877		**4th**	4560		**4th**	6491

ELEMENT 1 is ___K_____. ELEMENT 2 is ____Br_____. ELEMENT 3 is ____Ca___.

Explanation: The ionization energies for each electron removed is shown in the chart. For element 1, there is a large jump in IE from first IE to the second IE. This means that substantially more energy was required to remove the second electron indicating that it was being removed from a stable electron configuration.

K – $1s^2\ 2s^2 2p^6 3s^2 3p^6\ 4s^1$ IE required to remove the valence electron = 418.8 kJ/mol giving K^{+1}

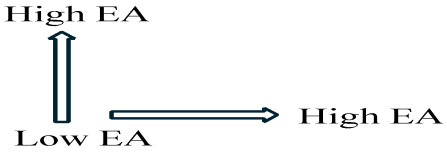
K^{+1} - $1s^2\ 2s^2 2p^6 3s^2 3p^6$ IE required to remove the 2nd electron = 3052 kJ/mol giving K^{+2} which we know doesn't occur. The huge IE is due to trying to remove an electron from a stable configuration like Ar.

Element 2 has the largest IE to start with indicating that even the first electron required a lot of energy to remove so it is more than likely the nonmetal Br.

Element number 3 shows the larges jump after the first and second IE. That would correspond to the 2 valence electron in Ca followed by a large jump in IE from the stable configuration like Ar.

ELECTRON AFFINITY

Draw the trend for ELECTRON AFFINITY



20. Define Electron Affinity:

    Electron affinity is the energy released when an electron is added to the outermost shell of an atom in the gaseous phase.

21. An atom has a very negative electron affinity. Circle all the statements that might apply to this atom.
    a. It may be a noble gas.
    b. It becomes less stable when electron is added.
    c. It becomes more stable when electron is added.
    d. It is probably a metal.
    e. It is probably a non-metal.
    f. There is a release of energy when electron is added.
    g. Energy is absorbed when electron is added.

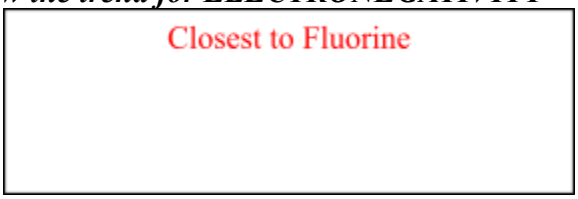
22. The first electron affinity of oxygen is -142 kJ mol-1 .
    a. What does the negative sign in front of the 142 kJ show?
    The negative sign indicates that energy is released as an electron is added to the outermost shell of an atom.

    b. Write the equation which shows the change taking place in the reaction for the first electron affinity for oxygen.

$$O + 1\ e^- \rightarrow O^{-1}$$

ELECTRONEGATIVITY (EN):

Draw the trend for ELECTRONEGATIVITY



23. Define Electronegativity:

    The ability of an atom to attract BONDED electrons to itself in a covalent bond.

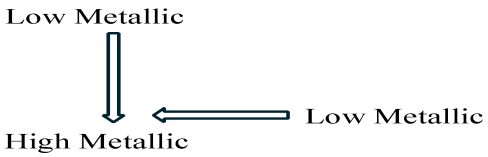
24. Rank each of the following in order of INCREASING electronegativity.
    a. Na, K, Cl _____ K, Na, Cl _____
    b. Fr, Ca, Co _____Fr, Ca, Co_____

25. Rank each of the following in order of DECREASING electronegativity.
    a. As, S, Sn _____S, As, Sn_____
    b. Rb, Br, I _____I, Br, Rb_____

26. Why do most noble gases not have electronegativity values? Noble gases are inert or stable and don't tend to react to form covalent bonds. Therefore, there are no BONDED electrons to attract.

METALLIC CHARACTER:

Draw the trend for METALLIC CHARACTER



Define Metallic Character:

Metallic character refers to the level of reactivity of a metal. Metals tend to lose electrons in chemical reactions, as indicated by their low ionization energies.

General Questions:

27. Based on the concept of periodic trends, answer the following questions for these atoms: Li, Be, Na, Mg. Be able to defend your answers.
    a. Which element has the lowest electronegativity (EN)? _____Na_____
    b. Which element has the least metallic character? _____Li_____
    c. Which element is the largest atom? _____Na_____

28. Based on the concept of periodic trends, answer the following questions for these atoms: P, S, Cl, F. Be prepared to defend your answer.
    a. Which element has the lowest electronegativity (EN)? _____P_____
    b. Which element has the least metallic character? _____P_____
    c. Which element is the largest atom? _____P_____