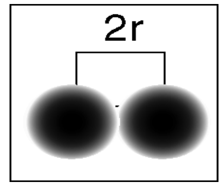
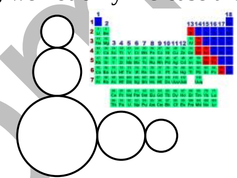
Name ___________________ Honors Chemistry ___/___/___

Periodic Trends



The periodic table is designed such that a great deal of information can be deduced by where elements are positioned. We will look at four trends today that will lay important ground work for the next few chapters.

- **ATOMIC RADIUS** is defined as half the distance between the nuclei of two like atoms located next to one another. We must measure the radius in this way, as it is difficult to pinpoint the outer energy level of an atom. But, since we can locate the nuclei of the two atoms, dividing this distance in half gives us the radius of one of the atoms.



LEFT TO RIGHT ACROSS A PERIOD, ATOMIC RADIUS DECREASES. This occurs because as we go across, we not only increase the number of protons, but we also increase the number of electrons. And, since no energy levels are added, the increased number of protons and electrons creates a **greater attraction** between the two and thus pull them closer to one another.

ATOMIC RADIUS INCREASES AS YOU GO DOWN A GROUP. This is because energy levels are added and the distance between the nucleus and the outer electrons is greater. Also, the added inner electrons reduce the attraction between the outer electrons and the nucleus. This reduction in the attraction between a nucleus and its outer electrons due to the blocking effect of inner(core) electrons is referred to as the **shielding effect.**

- **IONIZATION ENERGY** is defined as the amount of energy needed to remove an electron from an atom or ion in its ground state in the gas phase.

Consider the energy required to remove several electrons in succession from aluminum in the gaseous state:

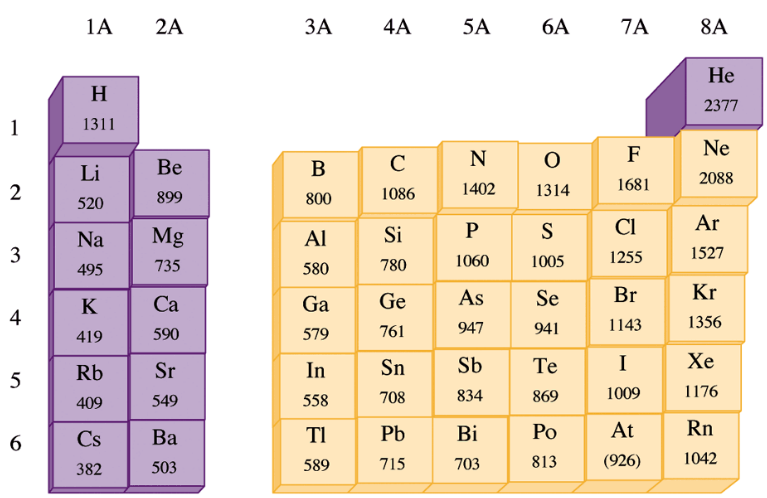
$Al(g) \rightarrow Al^{+}(g) + e^{-}$ $I_1 = 580$ kJ/mol
$Al^{+}(g) \rightarrow Al^{2+}(g) + e^{-}$ $I_2 = 1815$ kJ/mol
$Al^{2+}(g) \rightarrow Al^{3+}(g) + e^{-}$ $I_3 = 2740$ kJ/mol
$Al^{3+}(g) \rightarrow Al^{4+}(g) + e^{-}$ $I_4 = 11{,}600$ kJ/mol

In a stepwise ionization process, it is always the highest-energy electron (the one bound least tightly) that is removed first. The first ionization energy, I_1, is the energy required to remove the highest-energy electron of an atom. Note that as more electrons are removed, the successive ionization energies increase. This is because as the atoms lose an electron they become positively charged. This increase in positive charge binds the remaining electrons more firmly. **The big increase in aluminum's ionization energy between I_3 and I_4 is because in the fourth step a core electron is being removed. Core electrons are bound more tightly than valence electrons.**

LEFT TO RIGHT ACROSS A PERIOD, FIRST IONIZATION ENERGY INCREASES. Increasing the number of protons and electrons in the same principle energy level increases the attraction between the protons and the electrons, making it more difficult to remove an electron.

FIRST IONIZATION ENERGY DECREASES AS YOU GO DOWN A GROUP,. Because atoms gain energy levels of electrons as we move down a group, and inner electrons cause a **shielding effect**, it is easier to remove an electron from an atom with seven energy levels than it would be to remove an electron from an atom with only 2 energy levels.

The diagram to the right shows the trends in ionization energies (kJ/mol) for the representative elements. Note the discontinuities in ionization energy in going across a period. **In period two, discontinuities exist between beryllium and boron and nitrogen and oxygen**. The decrease in ionization energy going from Be to B is due to the fact that **the s subshell electrons shield the nucleus from the p subshell electron in boron** making its first ionization energy lower. The reason for the decrease in ionization energy when going from nitrogen to oxygen is due to **the fact that oxygen has one doubly filled orbital. The extra electron repulsions make its first ionization energy lower**.



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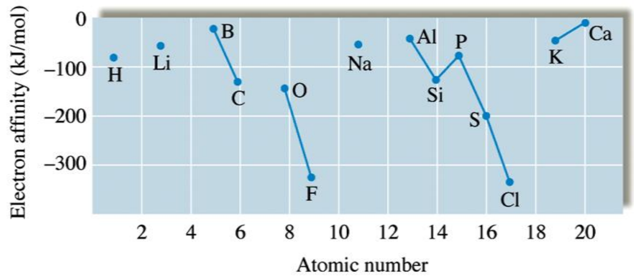
- **ELECTRON AFFINITY is the energy change associated with the addition of an electron to a gaseous atom.**
  - $X(g) + e^- \rightarrow X^-(g)$

The figure to the right shows the electron affinity values for the atoms among the first 20 elements that form stable, isolated negative ions. When the addition of an electron makes the atom more stable, energy is given off. This is true for most of the elements. When the addition of an electron makes the atom less stable, energy must be put in. That's because the added electron must be placed in a higher energy level, making the element less stable. This is the case for elements with full subshells, like the alkaline earth metals and the noble gases.



LEFT TO RIGHT ACROSS A PERIOD, THE ENERGY GIVEN OFF WHEN AN ELECTRON IS ADDED INCREASES.

ELECTRON AFFINITIES DO NOT CHANGE MUCH DOWN A GROUP.

The most reactive metals have the smallest ionization energy (Fr). The most reactive non-metals have the largest ionization energy and are the most exothermic when they gain electrons (F).

- **ELECTRONEGATIVITY** is defined as the tendency for an atom to attract electrons to itself when it is bonded to another atom.

LEFT TO RIGHT ACROSS A PERIOD, ELECTRONEGATIVITY INCREASES. One important note about **electronegativity** is that it **does not apply to the noble gases. This is because the noble gas elements do not readily form compounds**.

ELECTRONEGATIVITY DECREASES AS YOU MOVE DOWN A GROUP. This is again due to the shielding effect of the inner (core) electrons.

At the completion of this assignment you will be prepared to take the following Chapter 3 on-line quizzes:

• alkali metals trends quiz	• alkaline earth metals trends quiz	• atomic radius trends quiz
• chalcogen trends quiz	• electronegativity high low quiz	• electronegativity trends quiz
• halogen trends quiz	• increase decrease trends quiz	• ionization energy trends quiz
• largest smallest atomic radius quiz	• metalloid trends quiz	• mixed trends quiz 1
• mixed trends quiz 2	• period 2 trends quiz	• period 3 trends quiz
• period 4 trends quiz	• period 5 trends quiz	• period 6 trends quiz
• group 13 trend quiz	• group 14 trend quiz	• group 15 trend quiz

Homework: Write the symbol for the element described.

1. **Rb** The largest 5th period metal.
2. **F** The second period element with the largest electron affinity.
3. **Ge** The fourth period metalloid with the smallest atomic mass.
4. **Fr** The element with the lowest electronegativity that reacts with air and water and has to be stored in oil.
5. **Se** The nonmetal with an oxidation number of (2-) that has the greatest atomic mass.
6. **N** The element with 2 protons less than the most electronegative element.
7. **Ar** The least reactive third period element.
8. **Fr** The element with more protons than Sulfur that has the lowest ionization energy.
9. **Cl** The third period element with the largest electron affinity.
10. **Tc** The synthetic element with the smallest atomic number.
11. **Zn** The transition element with the smallest atomic radius.
12. **Se** The largest 4th period non-metal.
13. **Si** The third period metalloid with the largest atomic mass.
14. **Li** The element with the highest electronegativity that reacts with air and water and has to be stored in oil.
15. **F** The nonmetal with an oxidation number of (1-) that has the smallest atomic mass.
16. **Sc** The element with 12 protons more than the most electronegative element.
17. **Sb** The fifth period element with the electron configuration ending in p^3.
18. **Os** The element with the electron configuration $5d^6$.

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19. **Kr** A non-reactive element with the highest ionization energy that also has a higher atomic mass than Argon.
20. **S** The element with 13 more protons than the least electronegative second period element.
21. **K** The largest 4th period metal.
22. **Ga** The Group 13 element that is smaller than Indium and the least electronegative.
23. **As** The fourth period metalloid with the largest atomic mass.
24. **U** The element with 5 protons more than the element with the lowest electronegativity that also reacts with air and water and has to be stored in oil.
25. **P** The nonmetal with an oxidation number of (3-) that has the greatest atomic mass.
26. **Na** The element with 2 protons more than the most electronegative element.
27. **V** The element with the electron configuration $3d^3$.
28. **Na** The least electronegative third period element.
29. **Ar** The element with more protons than Sulfur that has the highest ionization energy.
30. **Sn** The Group 14 metal that has the highest ionization energy.
31. **??** Your favorite element. ☺
32. **Sc** The transition element with the smallest atomic mass.
33. **P** The third period element with the electron configuration ending in p^3.
34. **Ru** The element with the electron configuration $4d^6$.
35. **Ta** An element with 73 protons.
36. **Ar** The element with 7 more protons than the least electronegative third period element.
37. **Zn** The smallest of the Group 12 elements.
38. **Be** The smallest of the Alkaline Earth Metals.
39. **Ba** What do you do with dead people?
40. **At** The heaviest metalloid
41. **K** An alkali in the fourth period.
42. **Fe** A transition element whose d orbital configuration is $3d^6$.
43. **Br** The halogen in the fourth period.
44. **Rn** Noble gas element whose atoms are the heaviest.
45. **Ga** Atom whose electron configuration ends in $4p^1$.
46. **O** A second period element with a 2- oxidation number.
47. **H** An element with the largest atoms in the first period.
48. **Ar** A third period inert gas.
49. **He** Smallest atom of all the elements.
50. **H** Lightest atom of all the elements.
51. **Rn** Sixth period element whose configuration ends with p^6.
52. **Li** The second period element with the lowest electronegativity.
53. **Xe** The fifth period element with the highest ionization energy.
54. **Fr** The element with the lowest electronegativity.
55. **Po** The chalcogen metal with the lowest ionization energy.
56. What is atomic radius?

    Half the distance between the nuclei of 2 atoms of the same element located next to one another.
57. What is the shielding effect?

    Reduction in the attraction between the nucleus and the outer electrons due to the blocking by the inner electrons.
58. What is ionization energy?

    The amount of energy needed to remove an electron from an atom in the ground state and gas phase.
59. What is electron affinity?

    The energy change associated with the addition of an electron to a gaseous atom

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60. What is electronegativity?

The tendency for an atom to attract an electron to it when it is bonded to another atom.

61. When writing the periodic trend for electronegativity, why aren't the noble gases included?

Noble gases are not included because they do not form compounds.

For each of the following statements, determine which term it best describes. Use: **metal, metalloid** or **nonmetal**. You will use some terms more than once.

62. **nonmetal** These elements are brittle or gases.

63. **metal** These elements generally form cations.

64. **nonmetal** This group of elements contains solids, liquids & gases at room temperature.

65. **metal** All d block and f block elements belong to this group.

66. **metal** These elements are malleable and ductile.

67. **nonmetal** These elements generally gain electrons when they form ions.

68. **metal** These elements have high melting points.

69. **metal** These elements are good conductors of heat and electricity.

70. **metal** These elements have a shiny metallic appearance.

71. **metalloid** These elements have properties of metals and nonmetals.

72. **metalloid** These elements are used as semiconductors.

73. **nonmetal** These elements are poor conductors of heat & electricity.

74. **metal** These elements have high densities.

75. **nonmetal** These elements have high ionization energies.

Free Response: Briefly (in one to three sentences) explain each of the following in terms of atomic structure.

(a) In general, there is an increase in the first ionization energy from Li to Ne.
More protons in the nucleus and more electrons in the same energy level which leads to more attraction and a higher ionization energy

(b) The first ionization energy of B is lower than that of Be.
The s orbital electrons shield the nucleus from the p orbital electrons in B making its p orbital electron easier to remove.

(c) The first ionization energy of O is lower than that of N.
In nitrogen there is one electron in each p orbital. Oxygen has one doubly filled orbital which causes repulsion and allows for oxygen to have a lower first ionization energy.

(d) The electron affinities of beryllium and nitrogen are close to zero but the other elements in the period have much higher values.
Adding an electron to beryllium and nitrogen requires energy as it creates an unstable electron arrangement. It mimics the anomalies seen in ionization energy.

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