CHM 101 Name____________________

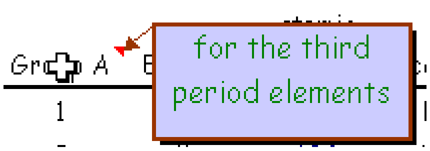
Interactive Periodic Trends: A Graphical Experience

This activity explores a variety of properties of the chemical elements as they vary based on position on the periodic table. You should have completed the Pre-lab Queries handout before starting this. You will need a copy of the periodic table and the interactive Excel spreadsheet available at:

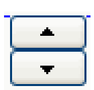
http://academic.pgcc.edu/~ssinex/excelets/PT_interactive. xls

We will examine properties as a function of increasing atomic number in two fashions: (1) across a period and (2) down a group. The tabs as shown in the illustration below will be used to navigate the spreadsheet.

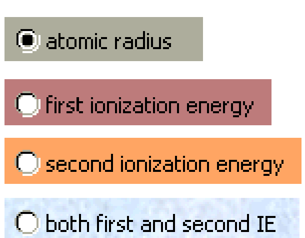
The following interactive Excel features will be found on this spreadsheet.



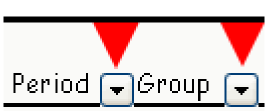
comment boxes – look for red triangles on a cell for information



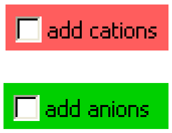
spinner – click on the up or down arrow to move the tracer points on the graphs



option buttons - to select what data are graphed – click on button (you can only select one at a time)



autofilters – click on black arrow to get a dropdown menu, select the period or group you want to plot (select all to get the full data set)

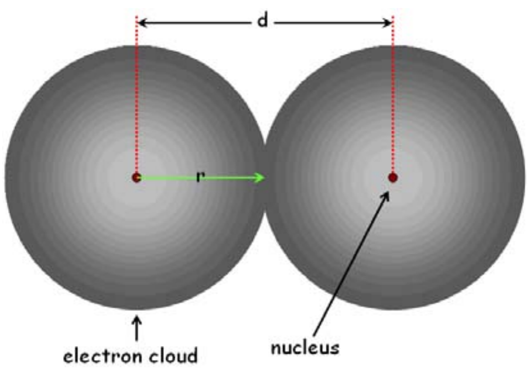


check boxes - click on for action (you may select any number of these)

hyperlink button – click for action

We will explore the atomic radii, ionization energies, electron affinities, and electronegativities for the first 38 elements by atomic number. These are the elements of Periods 1 to 4 plus the first two elements from Period 5.

The atomic radius or "size of an atom" can be found experimentally by dividing the distance between two nuclei of an element as measured by x-ray diffraction of the element in the solid state such as Fe (s) or Ar (s).



The radii are measured in units of picometers, pm, where 1 pm = 10^{-12} m.

The variation in the atomic radii is shown when you click on the atomic properties tab.

1. **Atomic Radius or the Size of an Atom**

Move the tracer to identify the elements and explore their atomic sizes. Locate the beginning of each period on the plot to assist you in the questions below.

Were your predictions correct on Question 1 of the Pre-lab Queries?

Which is larger? Li or K F or Br Explain.

Looking at the plot how would you describe the variation in atomic radius:

across a period?

down a group?

Without consulting a textbook, attempt an explanation for the variation in atomic radius (remember the electron configurations):

down a group

across a period

2. **Ionization Energy (Ionization Potential)**

Electrons are attracted to the nucleus due to opposite charges; hence, energy is required to remove an electron from an atom. The **ionization energy**, IE, is a measure of the energy for the following process:

$$X(g) \rightarrow X^{+}(g) + e^{-}$$

IE_1 is the designation for the first ionization energy (energy to remove the first electron).

Is this an endothermic or exothermic process? Explain.

For hydrogen:

$IE_1 = 2.18 \times 10^{-18}$ J/atom or 1.31 MJ/mole where MJ = 10^6 J

Verify the conversion from J/atom to MJ/mole. Show your calculations.

Click on the first ionization energy option button to display the first ionization energies, IE (in units of MJ/mole). If you place the cursor on any datum point on the graph, it will display the value. The valence electrons or outermost electrons are removed first. Looking at the graph for the first thirty-eight elements, find the element with:

easiest electron to remove ______ hardest electron to remove ______

Locate the points that represent the beginning of a new period. Describe the **general** variation in first ionization energy:

across a period

down a group

Locate the B Group elements (transition elements) on the graph. The B Group elements have partially filled d orbitals. Even though electrons go into a 4s orbital before the 3d, the first electron removed from the transition elements is a 4s electron. How do the first ionization energies for the Group B elements compare to each other? To the Group A elements?

Successive electrons can be removed from an ion. The second ionization energy (click the button), IE_2 would be given by:

$$X^{+}(g) \rightarrow X^{+2}(g) + e^{-} \quad IE_2$$

For the second ionization energy which of the following relationships would hold? Circle your choice and explain why.

$IE_2 < IE_1$ $IE_2 = IE_1$ $IE_2 > IE_1$

Plot the first and second ionization energies as a function of atomic number. In general, what can you say about the values of IE_2 compared to IE_1 for each element?

How do the graphs of IE_1 vs atomic number and IE_2 vs atomic number compare to each other? Where are the peaks in each?

Now, how do you think atomic radius and first ionization energy related? Explain you answer.

Click the "Plot of 1st IE as a function of atomic radius" button to view a graph and explain what you see.

Consider the ionization energy for each successive electron on an atom. You know the IE increases as the number of electrons removed increases; however, does it follow a pattern? How do you think the ionization energy changes as you remove electrons closer and closer to the nucleus? Why?

Click on the electron properties tab to explore a graph of the ionization energy of all calcium electrons as a function of ionization number (# of the electron removed). Which electrons are removed first?

Explain the pattern.

Here are the $1s^1$ or core electron ionization energies for the second period elements.

Element	$1s^1$ IE	Element	$1s^1$ IE
Li	11.81	N	64.36
Be	21.01	O	84.08
B	32.83	F	106.43
C	47.28	Ne	131.43

Source: Huheey et al (1993) *Inorganic Chemistry: Principles of Structure and Reactivity*, 4^{th} edition

Why does the energy to remove the core electron change? Would this be true for the $2s^1$ or $3d^3$ electron on any atom?

How does this difference in energies relate to the unique atomic spectrum for each element?

3. **Electron Affinity**

Electron affinity (EA) is usually defined as the energy released for the addition of an electron to a neutral atom as shown by the reaction below:

$$X(g) \; + \; e^{-} \; \rightarrow \; X^{-}(g) \quad \text{(usually exothermic)}$$

While removing an electron is always an endothermic process, adding an electron may be endothermic or exothermic. Why do you think this is true?

If the above reaction is reversed, then the energy for the process is the energy required to remove an electron from an ion, a definition that resembles the ionization energy.

$$X^{-}(g) \; \rightarrow \; X(g) \; + \; e^{-} \quad \text{(usually endothermic)}$$

Because we are removing an electron to make a neutral atom this reaction might be said to describe the **zeroth ionization energy**.

Click the atom properties. Click the option button to plot the electron affinities for the first 38 elements. These are calculated as if they were ionization energies for the anions as a function of atomic number.

What does a positive value tell you about the energy exchange in the process?

List three elements that would prefer to lose electrons based on their electron affinities.

In general, how does electron affinity change as you go:

down a group

across a period

How does the amount of energy required to remove an electron from an ion compare to removal from a neutral atom? Click the option button to plot EA, 1st IE, and 2nd IE. Explain the reason for the difference.

4. **Electronegativity**

Electronegativity (EN) is the attraction of the nucleus for the valence electrons or outermost electrons on an atom, or for the pair of electrons in a bond. In most textbooks it is classically given by the Pauling scale since Linus Pauling was the chemist who defined and first determined this property.

Click the option button to view the Electronegativity data. How does EN vary across a period?

How does EN vary down a group?

Does this trend make sense based on atomic radius and ionization energy? Explain.

5. **Ionic Radius**

Ions are formed by gaining or losing electrons and the resulting radius is called the **ionic radius**. Consider the electron configuration for the atom and the resulting ion. How would the ionic radius change compared to the atomic radius for the following transitions? Look back at your predictions on Question 2 of the Pre-lab Queries. Click the ion properties tab for some data. Explain the pattern.

atom to anion

atom to cation

Nitrogen can occur in a variety of oxidation states. Rank these ions in different oxidation states from highest to lowest radius and explain your answer:

N^{+3} N^{+5} N N^{-3}

How would the ionic radius of the following ions compare?

Na^{+} Mg^{+2} Al^{+3} Si^{+4} P^{+5} S^{+6}

Which ion is larger? Fe^{+2} Fe^{+3} Explain your choice.

6. **Group B Elements**

Do all periodic trends behave in the manner we have seen previously for the first 38 elements? Let's examine three properties for the Group B or transition metals for the 4th, 5th, and 6th periods on the periodic table. What is the trend for each of the following properties?

Melting point

across the period (by number of d electrons)

down a group

Density

across the period (by number of d electrons)

down a group

Resistivity

across the period (by number of d electrons)

down a group

7. **Summary**

Let's summarize the trends you uncovered in this activity. For each of the atomic properties listed in the table below, indicate whether, in general, the property *increases* or *decreases* across or down the periodic table.

Property	Across the Periodic Table	Down the Periodic Table
Atomic Radius		
Ionization Energy		
Electron Affinity		
Electronegativity		

8. **A Flashback to Spectroscopy**

Ions, like neutral atoms, have characteristic spectra that are generated by the transitions of electrons between energy levels. Since ions have different radii compared to their neutral atoms, explain how the spectra of an ion would compare to that of the corresponding neutral atom.

This is a modified version of the activity Discovering Periodic Trends: A Graphical Approach from *Exploring the Chemical World* by Gage, Sinex, and Basili (2003).