Amanda Gawron
Chemistry 101 – Spring 2012

Dr. Orna Kutai

Periodicity: A Study of the Periodic Patterns

Introduction

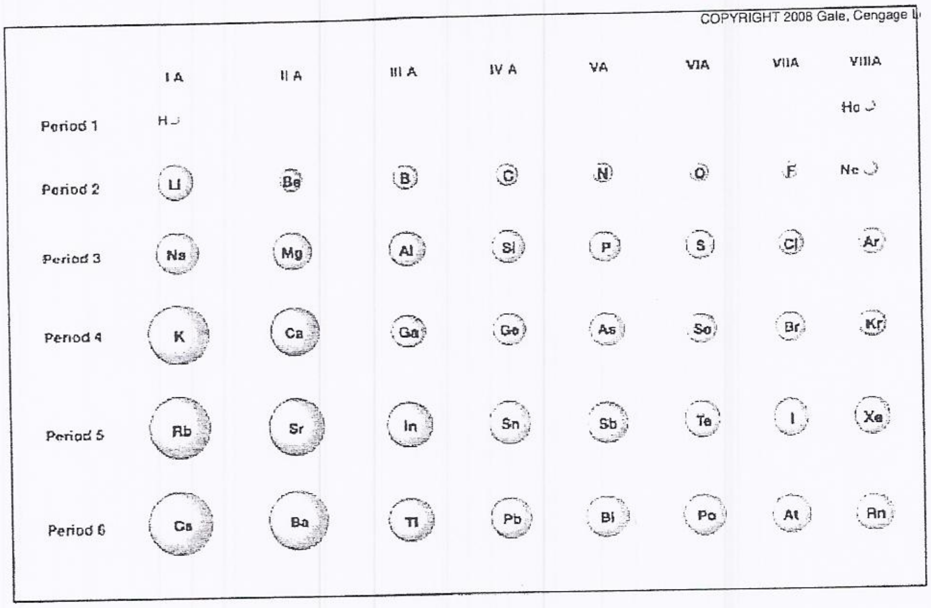
Throughout time, chemists have unveiled many new properties and patterns concerning the elements of the periodic table. The arrangement of the modern periodic table is owed to Russian chemist, Dmitri Mendeleev (1834-1907), who based his arrangement on periodic law. Periodic law states that elements are arranged in order of increasing mass from left to right across a period and that elements with similar properties are arranged from top to bottom down a group (1). The organization of the periodic table, however, is much more than simply atomic mass and similar elemental characteristics. A great deal of knowledge can be gained about the elements when trends of the periodic table are analyzed. With this paper a number of periodic trends, including atomic radius, ionic radius, ionization energy, electron affinity, metallic character, and metal and nonmetal reactivity, will be analyzed.

Atomic Radius

The first periodic trend to be examined is atomic radius, or atomic size. Atomic radius can be defined as the set of average bonding radii determined from measurements on a large number of elements and compounds (1). Atomic radii are incredibly small. For example, the atomic radius of elemental gold is approximately 144 picometers (pm) and that of oxygen is only 73 pm (2). One picometer is 10^{-12} meters or one trillionth of a meter (1).

There are two general periodic trends observed when examining atomic radius. Proceeding vertically down the group, the atomic radius increases. However, when moving across the period from left to right, atomic radius decreases (1). Study Figure 1.1 and notice the differences in atomic size as they correlate to the trends previously mentioned.

Figure 1.1 – Atomic radii of the main group elements, (3)



COPYRIGHT 2008 Gale, Cengage L

The reason that atoms increase in size proceeding down the group is because the principal energy level increases. A higher principal energy level means that there are a greater number of orbitals to occupy space. This means that moving down the group, there is a higher probability that the valence electrons are located in the principal energy level farthest away from the nucleus (2). The more energy levels in an atom, the larger it is.

Atomic radius decreases when moving across the period from left to right. The reason behind this phenomenon is largely due to shielding and effective nuclear charge. Shielding occurs when core electrons shield, or screen, electrons in the outermost principal energy level from the positively charged protons in the nucleus (1). Proceeding across a period, protons are added, thereby increasing the amount of pull that the nucleus has on its negatively charged electrons. This pull is called the actual nuclear charge. While shielding effectively protects the outermost electrons to some degree, the effective nuclear charge increases as more protons are added going left to right across a period (1). When the effective nuclear charge is greater, it causes electrons to be pulled in closer to the nucleus, thereby decreasing atomic radius.

Ionic Radius

The general periodic trends for ionic size are the same as the trends for atomic size. Moving down the group, ions get larger and moving across the period from left to right, ions get smaller. The difference lies in the analysis of cations versus anions. Study Figure 1.2 and notice how the cations and anions of elements differ in size from their corresponding atoms.

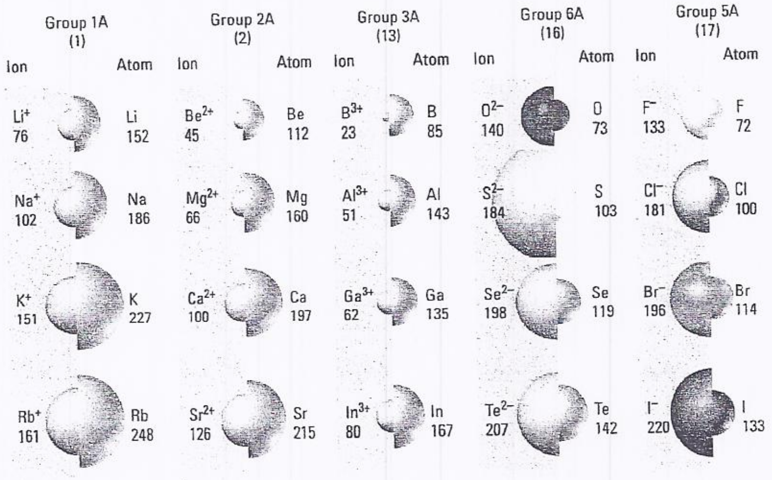


Figure 1.2 – Ionic radii as compared to their corresponding atoms (4)

Cations are much smaller than their corresponding atoms. The reason for this behavior is because when an atom becomes a cation, it loses the electrons in its outermost energy level. When the electrons are lost, so is the outermost principal energy level. The loss of the outermost principal energy level results in a smaller ionic radius as compared to the corresponding atom (1).

Anions are much larger than their corresponding atoms (1). The reason for the increase in size is evident when considering shielding and effective nuclear charge. When an atom becomes an anion, it gains electrons in its outermost principal energy level. However, it does not gain any corresponding protons in the nucleus. Therefore, the effective nuclear charge for an atom's anion is less than the effective nuclear charge for the atom. Since the pull on the outermost electrons is reduced in the anion, they are held more loosely and therefore the anionic radius is greater than the atomic radius of the corresponding atoms.

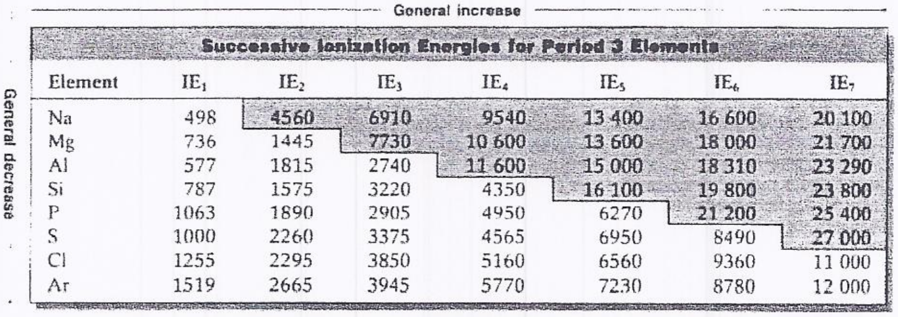
Ionization Energy

Ionization energy is defined as the amount of energy required to completely remove an electron from an atom in its gaseous state (5). Ionization energy is always positive because it always takes energy to remove an electron (1). The energy required to remove the first electron is called the first ionization energy, the energy required to remove the second electron is called the second ionization energy and the energy required to remove electrons beyond the second electron is called successive ionization energy. The trends in the first, second and successive ionization energies must be examined separately (1).

The first electron to be removed from an atom will be an electron in the outermost principal energy level. Ionization energy decreases proceeding vertically down the group because electrons in the outermost energy level are increasingly farther away from the nucleus. The inner energy levels contain electrons that shield the outermost electrons from the charge, or pull, of the nucleus. Therefore the valence electrons in the outermost principal energy level are held more loosely as energy levels are added. It is because of this increased shielding that it will take less energy to remove these electrons from the atom (1).

Ionization energy increases proceeding from left to right across a period because the effective nuclear charge is increased, causing electrons to be held tighter by the nucleus (1). It will take more energy to remove an electron if it is being more tightly held.

The second, and successive, ionization energies are much higher than the first ionization energy. Because in first ionization, an electron in the outermost energy level is being removed it is shielded somewhat from the nuclear charge and therefore is held more loosely than the core electrons (1). In second ionization and beyond, a core electron is being removed which takes a greater amount of energy. Examine Figure 1.3 and take note of the successive values for ionization energies for elements in period 3. The orange box represents core electrons being removed. Notice how the ionization energies increase drastically as core electrons are removed in the second ionizations and beyond.



General increase

General decrease

Successive Ionization Energies for Period 3 Elements

Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7
Na	498	4560	6910	9540	13 400	16 600	20 100
Mg	736	1445	7730	10 600	13 600	18 000	21 700
Al	577	1815	2740	11 600	15 000	18 310	23 290
Si	787	1575	3220	4350	16 100	19 800	23 800
P	1063	1890	2905	4950	6270	21 200	25 400
S	1000	2260	3375	4565	6950	8490	27 000
Cl	1255	2295	3850	5160	6560	9360	11 000
Ar	1519	2665	3945	5770	7230	8780	12 000

Figure 1.3 – Successive ionization energies for period 3 elements (6)

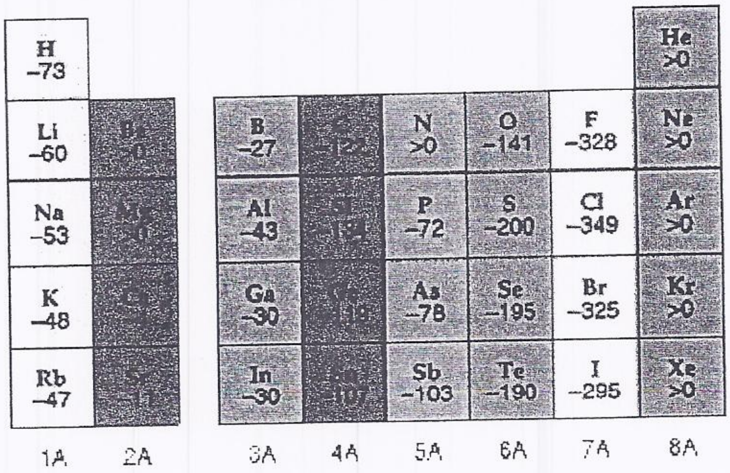
Electron Affinity

Electron affinity is the energy change associated with addition of an electron by an atom in the gaseous state. The value for electron affinity is usually negative, because like exothermic reactions, an atom or ion releases energy when it gains an electron (1). The trends for electron affinity are not as regular as those observed for atomic radius, ionic radius and ionization energy; however there are several worth analyzing.

The first noticeable trend when examining the values for electron affinity is that, while most groups on the periodic table do not exhibit any definite trend going from top to bottom, group 1A metals do. Electron affinity becomes more positive (the atom releases less energy) moving down the group from hydrogen to francium. Proceeding down group 1A, electrons enter orbitals with increasingly higher principal energy levels and will therefore be farther away from the nucleus (1).

Another periodic trend is that electron affinity generally becomes more negative proceeding from left to right across a period. This is because adding an electron makes an atom become more exothermic, or release more energy (1). Take a look at Figure 1.4 and note the electron affinities for sodium and chlorine.

Figure 1.4 – Electron affinities (kJ/mol) (7)



1A	2A	3A	4A	5A	6A	7A	8A
H −73							He >0
Li −60	Be >0	B −27	C −122	N >0	O −141	F −328	Ne >0
Na −53	Mg >0	Al −43	Si −134	P −72	S −200	Cl −349	Ar >0
K −48	[illegible]	Ga −30	Ge −119	As −78	Se −195	Br −325	Kr >0
Rb −47	[illegible]	In −30	Sn −107	Sb −103	Te −190	I −295	Xe >0

The electron affinities of sodium and chlorine are consistent with the aforementioned trend; chlorine, which is farther to the right of sodium, has a much more negative electron affinity value than sodium. In fact, the halogens have the most negative electron affinities of all of the groups. The reason behind this trend is because adding an electron to chlorine, and to the other halogens, will give it a highly desirable and stable noble gas configuration. Additionally, the effective nuclear charge on chlorine is much greater than that of sodium. Because of its potential to achieve noble gas status and its great nuclear charge, the process of adding an electron to chlorine is much more exothermic, and therefore more negative (1).

There is one very important exception to the trend observed in moving across the period. As evidenced from Figure 1.4, the electron affinities for the group 5A elements present a hiccup in the pattern of values becoming more negative proceeding laterally across the period. This is because they already have half-filled *p* orbitals in their outermost energy level. This property makes them very stable in their own right. When an electron is added to nitrogen, for example, it must pair with another electron in an already occupied *p* orbital. Electron-electron repulsion between the two electrons occupying the same orbital causes the electron affinity to be more positive than that element in the preceding group (1).

Metallic Character

There are certain characteristics associated with metals. They are good conductors of electricity, have excellent malleability, ductility, a shiny surface, and most notably, they have a tendency to lose electrons in chemical reactions (1).

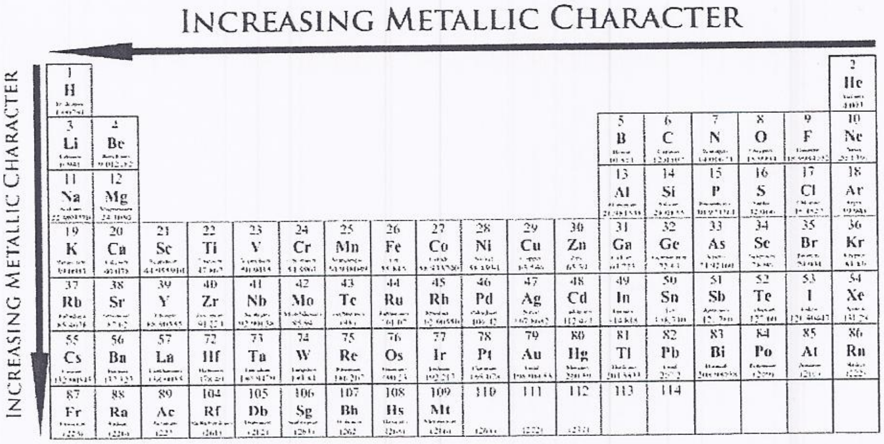


Figure 1.5 – Periodic table of metallic character trend (8)

Take a look at Figure 1.5 and notice the trend for metallic character; it increases proceeding down a group and decreases proceeding from left to right across the period. Moving to the right, ionization energy increases and electron affinity becomes more negative; therefore, elements on the left side of the periodic table are more likely to lose electrons than those on the right (1). Since the ionization energy increases and the electron affinity becomes more negative, there is a greater hold on the electrons in the outermost energy levels of the nonmetals on the right. Since the very definition of a metal is the ability to lose electrons easily it thereby follows that if it is more difficult to remove an electron from an atom, the element is less metallic in character.

Proceeding down a group on the periodic table, metallic character increases. The reason behind this trend is a decrease in ionization energy. As to be expected, the decrease in ionization energy makes it is easier to lose electrons in chemical reactions (1).

These trends explain the distribution of metals and nonmetals in the periodic table. Following the trends, it then makes sense that metals are located on the lower left side and middle of the periodic table and nonmetals are located on the upper right (1).

Metal and Non-Metal Reactivity

Reactivity is defined as the degree to which a particular atom or molecule is active and able to combine chemically with another molecule resulting in a chemical or physical change (9). There are several periodic trends associated with the reactivity of metals and nonmetals. The trends are most easily observed when a few specific groups are examined: the alkali metals (group 1A), the halogens (group 7A) and the noble gases (group 8A). These three groups exemplify the connection between chemical behavior and electron configuration (1).

The most reactive metals are the alkali metals of group 1A. Their high reactivity is due to the fact that their outermost electron configurations are ns^1. Therefore, they only have to lose one electron in order to achieve noble gas configuration, and this single valence electron is easily removed (1). It is for these reasons that alkali metals are very easily oxidized and that they are most often seen in nature as compounds rather than as neat metals, like gold or silver. Additionally, because the ionization energy for alkali metals decreases proceeding down the group, the reactivity increases. Alkali metals react with water to form basic solutions and they react with non-metals to produce vigorous reactions that emit a great deal of heat. The reactions become even more volatile moving down the group (1). There are other trends within the group, as well. For example, densities increase, (with the exception of potassium), masses increase and melting points decrease proceeding down the column (1).

The halogens of group 7A are located on the right side of the periodic table. Like the alkali metals, the atomic radii and densities of the halogens increase proceeding down the group. Also similarly to the alkali metals, the electron configurations for the halogens are one electron away from achieving noble gas configuration; they have electron configurations ns^2np^5 (1). The halogens only need to gain one electron in order to become like the noble gases in terms of electron configuration. Since they have highly negative electron affinities the halogens are among the most reactive non-metals on the periodic table (1). Halogens are easily reduced and will react with metals to form metal halides, hydrogen to form hydrogen halides, and with each other to form interhalogen compounds (1).

Finally, the noble gases of group 8A are an example of chemical inertness. As expected from general periodic trends, the atomic radii increase, densities increase and ionization energies decrease proceeding down the group. What makes the noble gases very special is that they have a full complement of electrons in their outermost principal energy level: ns^2np^6. The filled outermost energy levels and extremely high ionization energies make the noble gases exceptionally unreactive (1).

Conclusion

To conclude, a great deal can be learned from studying periodic trends. There are some general trends that are evident throughout. For example, shielding increases proceeding down the groups and nuclear charge increases moving from left to right across the periods. The shielding effect and nuclear charge are responsible for many of the periodic trends studied in this paper. There are some exceptions to the trends, and it is very important to take note of these as well. The trends of the periodic table may continue to evolve with technology and new discoveries.

Amanda Gawron
Chemistry 101 – Spring 2012
Dr. Orna Kutai

References

(1) Tro, Nivaldo J., Chemistry: A Molecular Approach. 2nd ed; Pearson Education, Inc. Prentice Hall: New Jersey, 2008. pp 314-357.

(2) Atomic radius. *World of Chemistry* [Online]; Gale Science In Context, Posted July 31, 2006. http://mchoudini.montgomerycollege.edu:3186/ic/scic/ReferenceDetailsPage/ReferenceDetailsWindow?displayGroupName=Reference&disableHighlighting=false&prodId=SCIC&action=e&windowstate=normal&catId=&documentId=GALE%7CCV2432500069&mode=view&userGroupName=rock77357&jsid=7e1ac230afad0f0bc3887531fa660bed (accessed Apr 3, 2012).

(3) Atomic radii of the main-group elements. *The Gale Encyclopedia of Science* [Online]; Gale, Cengage Learning, Posted 2010. http://mchoudini.montgomerycollege.edu:3186/ic/scic/StatisticsDetailsPage/StatisticsDetailsWindow?displayGroupName=Statistics&currPage=1&prodId=SCIC&action=e&windowstate=normal&catId=&view=statisticsDocDisplay&documentId=GALE%7CCV2210046216&mode=view&userGroupName=rock77357&jsid=612b905934ce5a6d0347c076b94bebf3 (accessed Apr 3, 2012).

(4) General Periodic Trends, from Chemistry, Moore/Stanitski/Jurs, Thomson, 2006, pg 311. http://schools.birdville.k12.tx.us/cms/lib2/TX01000797/Centricity/Domain/912/ChemLessons/Lessons/Periodic Table/General Periodic Trends.htm (accessed 04/04/12).

(5) Ion and ionization. *The Gale Encyclopedia of Science* [Online]; Gale, Cengage Learning, Posted 2008. http://mchoudini.montgomerycollege.edu:3186/ic/scic/ReferenceDetailsPage/ReferenceDetailsWindow?displayGroupName=Reference&disableHighlighting=false&prodId=SCIC&action=e&windowstate=normal&catId=&documentId=GALE%7CCV2644031226&mode=view&userGroupName=rock77357&jsid=ced030a69186fc6acb9b875e7d3acf24 (accessed Apr 3, 2012).

(6) Periodic Table. http://www.avon-chemistry.com/p_table_lecture.html, (accessed 04/04/12), part of Avon Chemistry.

(7) Dr. Garces, Fred, 3.06 Periodic Table and Periodic Trends, part of Chemistry 100, Miramar College. http://faculty.sdmiramar.edu/fgarces/zCourse/All_Year/Ch100_MM/aMy_FileLec/04MM_LecNotes_Ch100/03_AtomsElements/306_PeriodicTable/306_PeriodicTrends.htm (accessed 04/04/12).

(8) Periodic Table of Metallic Character Trend, part of Periodic Trends, UC Davis. http://chemwiki.ucdavis.edu/Inorganic_Chemistry/Descriptive_Chemistry/Periodic_Table_of_the_Elements/Periodic_Trends (accessed 04/04/12).

(9) Reactivity. *World of Chemistry* [Online]; Gale Science In Context, Posted July 31, 2006. http://mchoudini.montgomerycollege.edu:3186/ic/scic/ReferenceDetailsPage/ReferenceDetailsWindow?displayGroupName=Reference&disableHighlighting=false&prodId=SCIC&action=e&windowstate=normal&catId=&documentId=GALE%7CCV2432500623&mode=view&userGroupName=rock77357&jsid=da9b28476fa89ef5ef1b63293f33d87 (accessed Apr 4, 2012).

Dr. Kutai

CH101 Periodicity Paper

You will write a paper describing and explaining the following periodic trends:

- atomic radius
- ionic radius
- ionization energy
- electron affinity

Your paper will be evaluated on the clarity and accuracy of the explanations and proper citation of sources (according to ACS Style Guide). Online scientific sources must be vetted, for example, no Wikipedia. Proper English (diction, punctuation, grammar, etc.) must be used. Longer papers are not better! Indeed, often longer papers are weaker because they are characterized by tedious repetition of ideas and are less compelling. The repetition often occurs because the initial sentence is not well crafted. Be thorough but economical with your words. Make use of the Writing Center in HU002. For hours see http://cms.montgomerycollege.edu/edu/department.aspx?id=17510.

The assignment is divided into four phases:

Phase 1: Outline--due W 4/4 (1 pt)
Phase 2: Bibliography----due W 4/4 (2 pts)
Phase 3: Draft (submitted after peer review)--due 4/6 (3 pts)
Phase 4: Final Paper--due 4/9 (14 pts)

The Final version of your paper must include a Reference section (what is called a Bibliography in an English paper). The reference section of your final paper may include more sources than are included in the original "Bibliography" that you submitted on 4/4. *The rubric for the final paper follows at the end of this document.*

Phases 1 and 2: Outline and bibliography

- Outline may include words or phrases.
- Bibliography must follow ACS protocol; for ACS protocol, go to Course Content, go to Reference Materials, scroll to bottom where ACS Guidelines are listed in a link. A bibliography (what is usually called a list of References in scientific papers) must be included in your final paper. The reference section of your final paper may include more sources than are included in the original "Bibliography" that you are now submitting on 4/4.

Rubric:
Outlines submitted on time and display organized thoughts/topics/explanations (1 pt)
References (Bibliography) submitted on time include "vetted" sources (1 pt)
References follow ACS guidelines (1 pt)
No credit will be earned for Outlines or References submitted late.

Phase 3: Rough Draft

Write your initial draft with the spirit of getting it right, rather than getting it done. If you start the paper to late, you will run out of steam, the prose will be hurried, and your short-cuts will be obvious in your lack of clarity.

The initial draft must be peer reviewed before you submit it. When you submit, please indicate who peer edited your paper. You should have more than one individual peer edit your paper. Each student who reads your paper offers a new perspective. Invite students to be critical. Ask them to read your paper to be sure it is understandable.

Once you have student feedback, update your paper, reading it critically yourself. Sleep on it, then update the next day. Then submit the rough draft to me. No later than 4/9, please hand me a signed note from the Writing Center indicating the time, date, and name of person who reviewed your draft (or slip it under my door before the Science Center closes). When you have the final copy, you will submit it through SafeAssign.

CH101 Periodicity Paper

Dr. Kutai

Grading Rubric Periodicity Paper

Student: ____________________

	Quality			
	3	**2**	**1**	**0**
Criteria				
Organization	Information is highly organized with well-constructed paragraphs and subheadings.	Information is organized with well-constructed paragraphs.	Information is organized, but paragraphs are not well-constructed.	The information appears to be disorganized.
Content	All trends are addressed and expertly explained.	Trends are addressed but one is not adequately explained	Trends are addressed two or more are superficially or inaccurately explained	One or more trends are not addressed.
Diagrams / Graphics	Diagrams, tables, and graphs have been included for: (1) atomic size, (2) ionic size, (3) ionization energy, and (4) electron affinity *and* are readable, accurate, explained, support the text, and significantly add to the reader's understanding of the topic.	Diagrams, tables, and graphs have been included for: (1) atomic size, (2) ionic size, (3) ionization energy, and (4) electron affinity *and* are usually explained, and add to the reader's understanding of the topic.	Some diagrams, tables, and graphs have been included for: (1) atomic size, (2) ionic size, (3) ionization energy, and (4) electron affinity *and* are partially explained, and add somewhat to the reader's understanding of the topic.	Diagrams, tables, graphs have not been included OR are not accurate OR do not add to the reader's understanding of the topic.
Spelling / Punctuation / Grammar	No spelling errors. Correct use of Punctuation. No grammatical errors.	1-2 mistakes made in either spelling, punctuation OR grammar.	Numerous mistakes in spelling, punctuation OR grammar	Numerous mistakes in spelling, punctuation AND grammar.
References			All sources (information and graphics) are accurately documented in the ACS format.	Not all sources (information and graphics) are accurately documented or in the ACS format.
Writing Center			A signed note from Writing Center indicating time, date, & name of person who reviewed your draft is provided by 4/9 before Science Center closes	A signed note from Writing Center is not provided by 4/9 before Science Center closes