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Misconceptions Demonstrated in AP Chemistry Regarding Bonding and Periodicity Bonds

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The study of misconceptions has had a long history in science education. Research into their origin covers many different aspects of this issue, including their source and how to combat them. Sources include reasoning ability, belief system, popular culture, and also instruction. Students will come to our class with misconceptions, and they will also leave our class with them. Recognizing how they are connected to a particular topic provides insight into how the topic is understood. Communicating the correct answer is not enough for correction. It is important to allow the individual an opportunity to realize that their understanding of the concept cannot explain an event or observation. The goal of this series of articles is to identify misconceptions that are frequently demonstrated by students taking the AP Chemistry Examination, provide a correct way of thinking about the concept, and provide suggestions and references to instructional approaches that you may find useful to combat them in your own classroom.

Half-filled and completely filled shells are happy.

When explaining second ionization energies for alkali metals, students often say that the atom becomes "happy" or even becomes a noble gas! They say that an atom is more stable or that it does not want to break a filled shell because filled shells are happier, when the issue is really that the inner core electrons are closer to the nucleus.

Many books state that half-filled and completely filled sublevels are especially stable. Though this statement is true, there is a need for an energy explanation.

For example, why does chromium have an electron configuration of $[Ar]4s^1 3d^5$ and not $[Ar]4s^2 3d^4$? The explanation is that the arrangement $[Ar]4s^1 3d^5$ is lower in energy than the arrangement $[Ar]4s^2 3d^4$. A popular animation is available on the Web at http://intro.chem.okstate.edu/WorkshopFolder/Electronconfnew.html. This animation, written by Nancy Gettys, John Gelder and Judd Wheeler**.** provides a dynamic view of how the energies of atomic orbitals change when stepping through the elements in the periodic table.

In another example, why does copper have an electron configuration of $[Ar]4s^1 3d^{10}$ and not $[Ar]4s^2 3d^9$? In this case, $[Ar]4s^1 3d^{10}$ is a lower energy arrangement than $[Ar]4s^2 3d^9$.

In both of these cases, the happiness of the electrons is not the issue; the observed arrangement is argued in terms of the energy of the arrangement.

Why is the first ionization energy for oxygen lower than the first ionization energy for nitrogen? Student answers might argue that nitrogen's half-filled subshell is happy and that the ionization of the electron in oxygen is easier because the resulting arrangement is a half-filled subshell. The preferred explanation focuses on what happens when going from nitrogen to oxygen. A proton is added to the nucleus, but more importantly an electron is added to the 2*p* subshell, which has all three orbitals singly occupied. Adding another electron means the electron in one of the orbitals will be paired. Since electrons have a negative charge, some energy is required to pair two electrons in the same orbital. This is called the spin-pairing energy. Adding a second electron to an orbital already singly occupied increases the energy of the new arrangement. This extra energy means that the new electron in oxygen is easier to remove.

Why is the second ionization energy for a sodium atom considerably higher than the first ionization energy? Student answers might argue that the second electron removed from sodium must come from a filled shell and that filled shells are happy; therefore it is more difficult to remove an electron from a filled shell. In this case, the best argument to use is based on an effective nuclear charge along with shielding effects. The electron configuration for sodium is $1s^2 2s^2 2p^6 3s^1$. The first electron that is ionized comes from the 3*s* orbital, since that electron is the furthest from the nucleus and experiences the smallest effective nuclear charge (ENC). In this case, the ENC for an electron in the 3*s* orbital is +1. The second electron must be removed from the second level; likewise, the ENC experienced by all of the electrons in this level is +9 since the only electrons shielding the electrons in the second level are the 1*s* electrons.
Another way to look at the response to this question is in the form of a table:

Ionization Energy	Element	Complete Electron Configuration	Nuclear Charge (Z)	Level of the Electron Being Ionized	Number of Shielding Inner Core (IC) Electrons	Effective Nuclear Charge (Z - IC Electrons)
First	Na	$1s^2 2s^2 2p^6 3s^1$	11	3	10	+1
Second	Na	$1s^2 2s^2 2p^6 3s^1$	11	2	2	+9

Several widely used textbooks provide excellent explanations, including the following:

- Theodore E. Brown, H. Eugene Lemay, Bruce E. Bursten, and Julia R. Burdge. *Chemistry: The Central Science*, 9th ed. Prentice Hall, 2003. (Pages 246-249.)
- Steven Zumdahl and Susan Zumdahl. *Chemistry*, 6th ed. Houghton Mifflin, 2003. (Pages 324-326.)
- Kenneth W. Whitten, Raymond E. Davis, Larry Peck, and George G. Stanley. *General Chemistry*, 7th ed. Brooks/Cole, 2004. (Pages 240-241.)
- John C. Kotz and Paul M. Treichel. *Chemistry and Chemical Reactivity*, 5th ed. Brooks/Cole, 2003. (Pages 308-309.)
- Martin Silberberg, *Chemistry: The Molecular Nature of Matter and Change*, 4th ed. McGraw-Hill, 2006. (Pages 294-296.)

Effective nuclear charge can also be used to explain trends in atomic radii and ionic radii. A frequent question asks students to explain why sodium has a larger atomic radius than magnesium. One very common response to this question is that this is a result of the atomic radius decreasing from left to right across a period. While the statement about the trend is true, stating the trend is not an explanation for the trend. The correct response uses effective nuclear charge. Again consider the table of information below.

Element	Complete Electron Configuration	Nuclear Charge (Z)	Level of the Valence Electrons	Number of Shielding Inner Core (IC)	Effective Nuclear Charge (Z - IC

				Electrons	Electrons)
Na	$1s^2 2s^2 2p^6 3s^1$	11	3	10	+1
Mg	$1s^2 2s^2 2p^6 3s^2$	12	3	10	+2

In this case the valence electrons in magnesium experience a greater attraction to the nucleus than the valence electrons in sodium. Since the valence electrons feel a greater attraction, they are pulled closer to the nucleus, and the atomic radius for magnesium is smaller than the atomic radius in sodium.

The following guided inquiry activity is designed to help students invent better explanations for trends in atomic and ionic radii.

Atomic Radii and Ionic Radii

A Inquiry Activity on Periodicity

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Periodicity Laboratory[1]

Name(s) in Group____

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QUESTION: How do we use the model of the electronic structure of the atom to understand periodic trends of the elements?

I. Data Collection

Using a web browser access the following address:
http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/matters/periodicTbl2.html

When the animation starts you will see Figure I below:

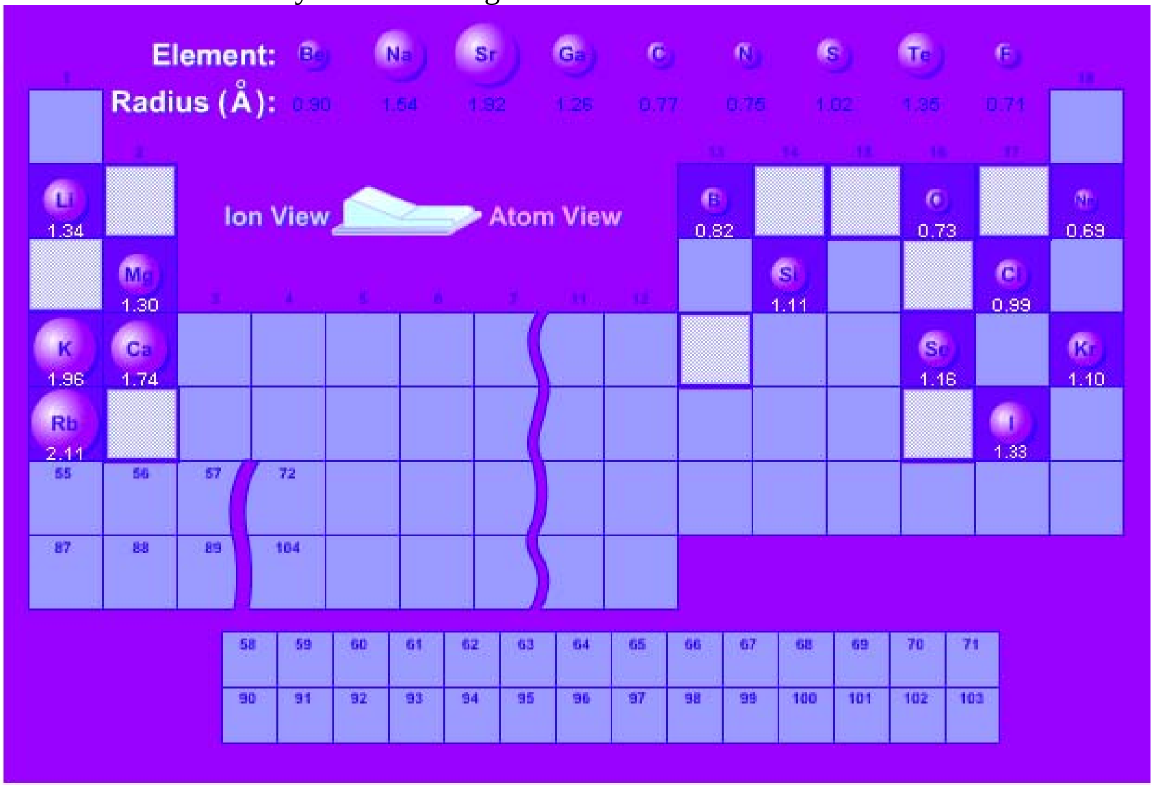

Figure I.

A. Look at the atomic sizes shown in the table and describe all the trends that you see in the space below.

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[1] Based on a tutorial written by Dr. Tom Greenbowe, Department of Chemistry, Iowa state University, Ames, IA 50011 (http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/periodic%20trends.htm)

Using the mouse, click on an atom in the row of atoms above the periodic table and drag them to their respective location in the periodic table.
After moving the atoms into the Periodic Table on your computer, add those atoms and their atomic radii to the Periodic Table in Figure I in this activity.

II. Data Analysis and Interpretation

A. While in Atom View select one of the Groups (1, 2, 16 or 17) and complete all the columns in the Table below: (Note: The empty column on the right will be used later in this activity.)

(a) Element symbol	(b) Atomic Radii	(c) Nuclear Charge (# of protons)	(d) Total # of electrons	(e) Electron configuration	(f) # of inner core electrons	(g) # of valence electrons	(h) Level the valence electron(s) occupy	

Table I.

After completing Table I. work with several other students or participate in a class discussion and answer parts B - D below.

B. Which columns (c – h) in Table I. do not explain the trend in atomic radii observed in your group and why?

C. Which columns (c – h) in Table I. explain the trend in atomic radii observed in your group and why?

D. Which column (c – h) in Table I. best explains the trend in atomic radii observed in your group and why?

III. Conclusion

A. The atomic radius of phosphorus (P) is 1.15 Å and for arsenic (As) the atomic radius is 1.25 Å. Provide an explanation why the atomic radius of arsenic is larger than the atomic radii for phosphorus.

B. Write a general statement that explains the trend in atomic radii that applies to any Group (vertical column) in the Periodic Table.

IV. Data Analysis and Interpretation

A. While in the atom view select one of the periods (vertical row) and complete the table below. (Note: The empty column on the right will be used later in this activity.)

(a) Element symbol	(b) Atomic Radii	(c) Nuclear Charge (# of protons)	(d) Total # of electrons	(e) Electron configuration	(f) # of inner core electrons	(g) # of valence electrons	(h) Level the valence electron(s) occupy	

Table II.

After completing Table II. work with several other students or participate in a class discussion and answer parts B - D below.

B. Which columns (c – h) in Table II. do not explain the trend in atomic radii observed in your period and why?

C. Which columns (c – h) in Table II. explain the trend in atomic radii observed in your period and why?

D. Which column (c – h) in Table II. best explains the trend in atomic radii observed in your period and why?

V. Conclusion

A. The atomic radius of aluminum (Al) is 1.18 Å and for phosphorus (P) the atomic radius is 1.06 Å. Provide an explanation why the atomic radius of aluminum is larger than the atomic radii for phosphorus.

B. Write a general statement that explains the trend in atomic radius that applies to any period in the Periodic Table.

VI. Data Analysis and Interpretation

A. Effective nuclear charge (ENC) is a term used to describe the amount of nuclear charge experienced by a particular electron in an atom. Nuclear charge is the total positive charge in the nucleus. For example the nuclear charge for a carbon atom is +6. The nuclear charge is equivalent to the atomic number for an element. Carbon has six protons so consider the following information;

(a) Element symbol	(b) Nuclear Charge (# of protons)	(c) Total # of electrons	(d) Electron configuration	(e) # of inner core electrons	(f) # of valence electrons	(g) Level the valence electron(s) occupy	(h) Effective Nuclear Charge (ENC)
C	+6	6	$1s^2 2s^2 2p^2$	2	4	2	+4

If we focus on the electron configuration for carbon we see there are 4 electrons in the second level and two electrons in the first level. Electrons in the second level are further away from the nucleus. The 2 electrons in the first level (inner core electrons) partially shield the valence electrons from some of the nuclear charge. To a first approximation we will assume each inner core electron shields one proton from the electrons in the outer most level (valence electrons).

The ENC is calculated by subtracting the number of inner core electrons (IC) from the total nuclear charge (Z): ENC = Z – IC. In the case of carbon, ENC = +6 – 2 = +4.

Return to Table I and Table II and write ENC (Effective Nuclear Charge) in the empty right most column, as shown below.

(a) Group Element symbol	(b) Atomic Radii	(c) Nuclear Charge (# of protons)	(d) Total # of electrons	(e) Electron configuration	(f) # of inner core electrons	(g) # of valence electrons	(h) Level the valence electron(s) occupy	(i) Effective Nuclear Charge (ENC)

Table III.

B. In Table I. And Table II. complete column (i) by determining the effective nuclear charge for each of the elements listed.

C. In the space below show the calculation for ENC for any two of the elements in Table I and from Table II.

After completing the value for the ENC in the right most column in Table I. And Table II. work with several other students or participate in a class discussion and answer parts D - F below.

D. Is there a relationship between ENC and the trend in atomic radius for the elements in the Group from Table I? Explain.

E. Is there a relationship between ENC and the trend in atomic radius for the elements in the period from Table II? Explain.

F. Hypothetical statement made by a student to another student:

> "When I look at Figure I. it appears to me that as the nuclear charge increases the atomic size increases."

Cite evidence that supports and/or refutes this student's statement.

VII. Data Collection

In the animation on the computer, switch to the Ion View.

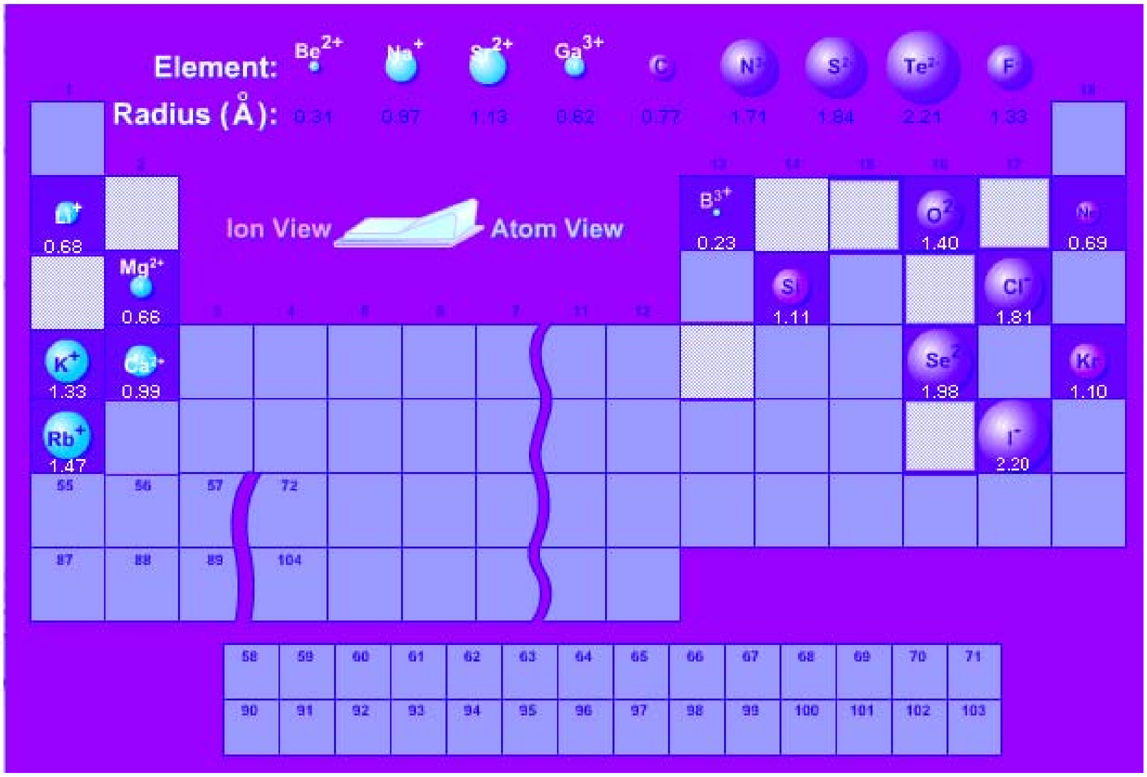

Figure II.

Using the mouse, click on an ion in the row of ions above the periodic table and drag them to their respective location in the periodic table.

After moving the ions into the Periodic Table on your computer, add those ions and their atomic radii to the Periodic Table in Figure II in this activity.

VIII. Data Analysis and Interpretation

A. Using the atomic and ion size information from Figure I and Figure II. select one metal from Group 1 or 2 and select one nonmetal from Group 16 or 17 and complete Table IV. In Table IV. enter the metal you selected in the first row and complete the row of information. In the next row below enter the cation for that metal and complete the information in the row. Do the same for the nonmetal you selected.

(a) Element symbol	(b) Atomic /ion Radii	(c) Nuclear Charge (# of protons)	(d) Total # of electrons	(e) Electron configuration	(f) # of inner core electrons	(g) # of valence electrons	(h) Level the valence electron(s) occupy
metal							
metal ion							
nonmetal							
nonmetal ion							

Table IV.

After completing Table IV. work with several other students or participate in a class discussion and answer parts B - K below.

B. Which columns (c – h) in Table III. cannot explain the trend in radii observed for the metal atom and its cation and why?

C Which columns (c – h) in Table III. explain the trend in radii observed for the metal atom and its cation and why?

D. In general what is the relationship between the radius for the neutral metal atom and its cation?

E. Which column (c – h) in Table III. best explain the trend in radii observed for the metal atom and its cation and why?

F. Write a general statement that explains the trend in radius between a metal atom and its ion that applies to any combination of metal and its cation.

G. Which columns (c – h) in Table III. cannot explain the trend in radii observed for the nonmetal atom and its anion and why?

H. Which columns (c – h) in Table III. explain the trend in radii observed for the nonmetal atom and its anion and why?

I. In general what is the relationship between the radius for the neutral nonmetal atom and its anion?

J. Which column (c – h) in Table III. best explains the trend in radii observed for the nonmetal atom and its anion and why?

K. Write a general statement that explains the trend in radius between a nonmetal atom and its anion that applies to any combination of nonmetal atom and its anion.

Complete the Table below.

Element symbol	Ion Radii	(c) Nuclear Charge (# of protons)	(d) Total # of electrons	(e) Electron configuration	(f) # of inner core electrons	(g) # of valence electrons	(h) Level the valence electron(s) occupy	(i) Effective Nuclear Charge (ENC)
S^{2-}								
Cl^{-}								
K^{+}								
Ca^{2+}								

Table V.

Which columns (c – h) in Table V. cannot explains the trend in ion radii observed in this selection of ions and why?

Which columns (c – h) in Table V. explain the trend in ion radii observed in this selection of ions and why?

Which column(s) (c – h) in Table V. best explains the trend in ion radii observed in this selection of ions and why?

Write a general statement that explains the trend in ion radius that applies to a related (isoelectronic) set of ions in the Periodic Table.