Simulation: Periodic Trends - Answer Key

Background
In this investigation you will examine several periodic trends, including atomic radius, ionization energy and ionic radius. You will be asked to interact with select atoms as you investigate these concepts.

1. Draw a picture to support a written definition of the word “radius.”

   Students should draw a picture of a circle, indicating the distance from the center of the circle to the edge of the circle.

2. Assuming atoms are shaped like spheres, what subatomic particles would be found in the center? What subatomic particles would be found around the perimeter?

   Center (nucleus) – protons & neutrons. Perimeter - electrons

3. Keeping in mind your answers to questions 1 & 2, in your own words describe the meaning of "atomic radius"

   Distance from the nucleus to the valence electrons.

4. What is an ion? What is a valence electron? How is an ion formed?

   An ion is an atom that has lost or gained electrons, causing it to have an electrical charge. Valence electrons are located in the outermost shell of an atom, and are the electrons involved in chemical bonding. An ion is formed by the removal of valence electron(s) or addition of electron(s) to the outermost shell of an atom.

5. What do you think ionization energy means? Think about this in relation to your answer to question #4.

   Ionization energy is the amount of energy required to remove a valence electron from the atom. Ionization energy is measured in KJ/mole.

6. Keeping in mind all of your answers thus far, attempt to define the term ionic radius.

   This is the distance from the nucleus to the valence electron in an ion. Generally speaking, a cation’s ionic radius will be smaller than its original atomic radius, while an anion’s ionic radius will be larger than its original atomic radius.

*Check your answers before moving on to the next portion of the activity.
Procedure
Using your computer, tablet or mobile device, navigate to the website: http://www.teachchemistry.org/periodic-trends. You should see the picture below on your screen.

![Periodic Table](image)

**Atomic Radius**
1. Choose any element shown in green from **group 1** on periodic table clicking the on the element symbol. You should see details about the element that you chose appear at the bottom of the screen. An example is shown below.

![Atomic Radius](image)

a. Select another element from **group 1** clicking on its symbol. Write the symbols and atomic number for each of the elements that you chose below:

   **Answers will vary.**

b. Which element appears larger in the side-by-side comparison?

   **Answers will vary here, but the element chosen from group 1 that has a larger atomic number will also have a larger atomic radius.**
c. What is the value in picometers (pm) for the radius of each atom? Do these values support your answer for part b?

   Answers will vary, but yes the data values from the simulation will support the visual size comparison.

   Reset the selected data using the reset symbol.

   d. Next, choose an element from a different group by clicking on its symbol. Again choose a second element to compare from the same group. Write the symbols and atomic number for each of the elements that you chose below:

   Answers will vary.

   e. Which element appears larger in the side-by-side comparison?

   Answers will vary here, but since the elements are chosen from the same group, whichever atom has a larger atomic number will also have a larger atomic radius.

   f. What is the value in picometers (pm) for the radius of each atom? Do these values support your answer for part e?

   Answers will vary, but yes the data values from the simulation will support the visual size comparison.

   g. Based on your answers in question 1 parts a-f, what is the general trend in the atomic radius of atoms within the same group? Give suggestions for why you think this trend exists based on your interaction with the elements.

   The atomic radius of atoms in the same group will increase from top to bottom of the group. Atoms with larger atomic numbers will have a larger atomic radius when compared to atoms in the same group.

2. Choose any element from period 2 on the periodic table by clicking on the element symbol. You should see details about the element that you chose appear at the bottom of the screen.

   a. Select another element from the period 2 by clicking on its symbol. Write the symbols and atomic number for each of the elements that you chose below:

   Answers will vary.

   b. Which element appears larger in the side-by-side comparison?

   Answers will vary here depending on the selection, but since the elements are chosen from the same period, whichever atom has a larger atomic number will have a smaller atomic radius due to effective nuclear charge experienced by the valence electrons in the atom.
c. What is the value in picometers (pm) for the radius of each atom? Do these values support your answer for part b?

Answers will vary, but yes the data values from the simulation will support the visual size comparison.

d. Do your answers in part b & c surprise you? Explain.

Answers will vary here. Many students will predict that as the atomic mass, number of protons and number of electrons of an atom increases within a period, the atomic radius will also be larger, when the opposite trend is actually observed. This would be a good opportunity to discuss the meaning of effective nuclear charge with students and dispel any misconceptions.

Reset the selected data using the reset symbol.

e. Choose an element from a different period by clicking on its symbol. Again choose a second element to compare from the same period. Write the symbols and atomic number for each of the elements that you chose below:

Answers will vary.

f. Which element appears larger in the side-by-side comparison?

Answers will vary here depending on the selection, but since the elements are chosen from the same period, whichever atom has a larger atomic number will have a smaller atomic radius due to effective nuclear charge experienced by the valence electrons in the atom.

g. What is the value in picometers (pm) for the radius of each atom? Do these values support your answer for part e?

Answers will vary, but yes the data values from the simulation will support the visual size comparison.

h. Based on your answers in question 2 parts a-g, what is the general trend in the atomic radius of atoms within the same period?

The atomic radius of atoms will decrease from left to right within a period. Atoms with larger atomic numbers will have a smaller atomic radius when compared to atoms in the same period.

i. Think about the possible contributing factors to the atomic radius trend within a period, specifically considering the protons in the nucleus, the electrons and the electron shells. List them below:

Effective Nuclear Charge will impact the atomic radius, as the number of protons in the nucleus increases from left to right in a period the strength of the net positive force felt by the valence electrons is also increased, causing the valence shell of electrons to be pulled closer, reducing the size of the
atom. Also, due to the atoms being in the same period, the number of electron shells will be the same for each of the atoms compared, so each atom in the period will experience the same amount of electron shielding.

Reset the selected data using the reset symbol.

3. Based on what you have learned, and without the assistance of the periodic trends simulation, predict which element is larger in the following pairs of atoms:
   
   a. Be or Sr  
   b. P or Ar   
   c. Rb or S   
   d. F or He   
   e. Br or Ca  
   f. Xe or Ba

Using the simulation, check your predicted answers to see if you are correct!

Ionization energy

4. Choose an element from the Alkali Metal family (group 1) by clicking on the element symbol. You should see details about the element that you chose appear at the bottom of the screen. An example is shown below.

a. Using your cursor attempt to ionize the atom that you chose by pulling a valence electron from the electron shell. Describe what happened. (Were you successful? Was it “easy” to remove the electron? Did the atom seem to have a strong hold on the electron?)  
   Alkali metals should be ionized easily, with little pull needed by the cursor.  

b. What was the ionization energy value for the atom that you chose?  
   Answers will vary.

c. Did any other information about the atom change after your attempt to
ionize the atom?

The atom will reduce in size, and the number of electrons will reduce by 1. The atom will now indicate that it has been ionized, and has a +1 charge.

d. Now choose the Noble Gas element that is in the same period as the Alkali metal chosen in part a. Attempt to ionize this atom by pulling a valence electron from the electron shell. Describe what happened. (Were you successful? Was it “easy” to remove the electron? Did the atom seem to have a strong hold on the electron?)

The valence electrons cannot be removed from the noble gas elements. The atom does have a strong hold on its valence electrons, making them not very reactive.

e. What was the ionization energy value for the noble gas atom that you chose?

Answers will vary.

f. Make a comparison statement about the two elements that you interacted with in terms of why they require different amounts of ionization energy.

Answers will vary. The noble gas is a stable element with a full outer shell of electrons, and therefore not very reactive whereas the alkali metal will be reactive, needing low amounts of ionization energy to lose its 2 valence electrons and become a stable ion.

g. Next, with the two elements still selected click on the “Go to Graphs” button:

A graph should appear, you will need to ensure the “First Ionization Energy” filter is selected in the top right hand location of the graph, for example:

What trend in ionization energy do you observe for elements in the same period based on the data in the graph?

The ionization value increases from left to right for atoms in the same period, with the noble gas requiring the highest amount of energy in the period to be ionized.

h. While still analyzing the graph, make a prediction about the trend in ionization energy between atoms in the same group on the periodic table.
For example, do atoms with larger atomic numbers have greater ionization energy than atoms with small atomic numbers in the same group?

Larger atoms will have smaller ionization energy values when compared to smaller atoms in the same group. This is due to the reduction in the effective nuclear charge on the valence electrons from the shielding effect.

**Navigate back to the main page, and reset the data using the reset symbol.**

i. Now choose two elements that are in the same group. How do their ionization energy values compare? Does this data support your prediction from part h?

The larger atom, located closer to the bottom of the periodic table will have a smaller ionization energy value when compared to a smaller atom located above it in the same group. The data will support the trend.

5. Based on what you have learned, and without the assistance of the periodic trends simulation, organize the following lists of atoms from lowest ionization energy to highest ionization energy:

   a. S, Na, Al, Ar = Na, Al, S, Ar
   b. I, F, Br, Cl = I, Br, Cl, F
   c. Rb, O, Si, Mg, He = Rb, Mg, Si, O, He

**Using the simulation, check your predicted answers to see if you are correct!**

6. Reflecting on what you have learned about both atomic radius and ionization energy at this point, which of the following statements best describe these trends?

   - Atoms that have large atomic radii also have large values of ionization energy.
   - Atoms that have small atomic radii will have large values of ionization energy.

   Explain your choice referencing both the atomic model and subatomic particles:

Small atoms have less electron shells, and do not experience as much electron shielding as large atoms do. Therefore, the positive nuclear charge in the nucleus of a smaller atom can pull its valence electrons closer, resulting in a smaller radius, and a need for a larger amount of energy to ionize the atom.

**Reset the selected data using the reset symbol.**

**Ionic Radius**

7. Choose an element from the Alkali Metal family (group 1) by clicking on the element symbol. You should see details about the element that you chose appear at the bottom of the screen. An example is shown below.
a. What is the atomic radius value for this element?

Answers will vary.

b. Using your cursor ionize the atom that you chose by pulling a valence electron from the electron shell until it is fully removed. What happened to the electron shell where this valence electron was located? How is the change in subatomic particles related to the size of the ion?

The electron shell will lighten in color to indicate that it no longer holds electrons, and is no longer used in the measurement of ionic radius. There are now more protons than electrons in the ion, so the positive charge of the nucleus pulls the surrounding electrons closer, reducing the ionic radius.

c. Since this atom is now ionized, you should see a value for the ionic radius. What is the value? Is this value larger or smaller than the value for the atomic radius in part a?

Answers will vary, however the ionic radius for the cation will be smaller than the original radius of the neutral atom from part a.

_Reset the selected data using the reset symbol._

8. Next, choose an element from the _Alkaline Earth Metal family_ (group 2) by clicking on the element symbol. You should see details about the element that you chose appear at the bottom of the screen.

a. What is the atomic radius value for this element?

Answers will vary.

b. Based on the atomic structure of the atom you chose, how many electrons
will need to be removed in order for it to become a stable ion?

2. Remove the necessary valence electrons from this atom, and record the value for ionic radius below:
   Answers will vary.

9. Both of the atoms selected in question #7 and #8 are metals, that form cations (positively charged ions) when they are ionized.
   a. Based on your answers to these questions, is the atomic radius of the neutral atom bigger or smaller than the radius of its cation?
      The atomic radius of the neutral atom is larger than the radius of its resulting cation.
   b. Why does this trend occur?
      This is because there are now more protons than electrons in the ion, so the positive charge of the nucleus pulls the surrounding electrons closer, reducing the ionic radius.

10. Choose an element from the Halogen family (group 17) by clicking on the element symbol. You should see details about the element that you chose appear at the bottom of the screen.
   a. How many valence electrons are present in the atom that you chose? 7.
   b. Using your cursor attempt to ionize the atom that you chose by pulling a valence electron from the electron shell. Describe what happened. (Were you successful? Was it “easy” to remove the electron? Did the atom seem to have a strong hold on the electron?)
      The student will not be able to ionize the halogen atom, the atom has a strong hold on its valence electrons, indicated by a large value for ionization energy.
   c. In order to make the halogen atom stable, by having a complete outer shell of electrons, what would be an easier solution compared to removing all of the valence electrons?
      Adding an electron.
   d. Based on your answer to part c, do you think the ionic radius will be larger or smaller than the atomic radius for this atom? Justify your prediction with scientific reasoning.
      The anion formed from a halogen will have a larger ionic radius than the neutral atom because of the addition of an electron, electrons will outnumber protons in the ion, so the positive charge of the nucleus will not have as strong of a pull on the surrounding electrons increasing the ionic radius.