Advanced Periodic Trends

How do the particles in an atom interact to dictate periodic properties of elements?

Why?

Previously you have learned about Coulombic attraction and how it governs the trends in properties among the elements of the Periodic Table. For example, as you move across a period, atoms have more protons in the nucleus, which pulls the electrons in tighter making smaller, more electronegative atoms. As you proceed down a column of the periodic table, the valence electrons get further away, and the atoms get larger and less electronegative. Inquisitive students will always ask, “If more protons in the nucleus pull electrons in tighter, then why don’t atoms get smaller when going down a column?” That is a good question, and one that you will investigate in this activity.

Model 1 – Period 3 Elements

<table>
<thead>
<tr>
<th></th>
<th>Sodium</th>
<th>Aluminum</th>
<th>Chlorine</th>
</tr>
</thead>
<tbody>
<tr>
<td>Atomic Number</td>
<td>$11$</td>
<td>$13$</td>
<td>$17$</td>
</tr>
<tr>
<td>Number of Electrons</td>
<td>$11$</td>
<td>$13$</td>
<td>$17$</td>
</tr>
<tr>
<td>Number of Valence Electrons</td>
<td>$1$</td>
<td>$3$</td>
<td>$7$</td>
</tr>
<tr>
<td>Core Charge</td>
<td>$+1$</td>
<td>$+3$</td>
<td>$+7$</td>
</tr>
<tr>
<td>Atomic Radius</td>
<td>186 pm</td>
<td>143 pm</td>
<td>99 pm</td>
</tr>
<tr>
<td>1st Ionization Energy</td>
<td>496 kJ/mol</td>
<td>578 kJ/mol</td>
<td>1251 kJ/mol</td>
</tr>
<tr>
<td>Electronegativity</td>
<td>0.9</td>
<td>1.5</td>
<td>3.0</td>
</tr>
</tbody>
</table>

1. Describe the relationship of the three elements in Model 1 with regard to their relative positions on the periodic table.

   *The three elements are in the same row of the periodic table.*

2. Refer to a periodic table to complete the first three rows of the table in Model 1.

   *See Model 1.*
3. Consider the charge and location of the protons in an atom.
   
   a. Indicate on the atomic drawings a total charge provided by the protons in the proper location for each atom.
      
      See Model 1.
      
   b. The number you just added to the diagrams in Model 1 is called the nuclear charge. Explain why this is an appropriate name for this value.
      
      The only charged particles in the nucleus are protons. Therefore, the total charge for the nucleus is equal to the charge on all the protons in the nucleus.
      
4. Circle the valence electron(s) in each of the atoms in Model 1.
      
      See Model 1.
      
5. In Model 1, the shaded circle in each atom indicates the core of the atom—the nucleus and the nonvalence electrons. Discuss with your group how a core charge could be calculated for each of the atoms in Model 1 and write the result in the Model 1 table.
      
      The core charge is calculated by finding the sum of the charges on the protons and the core electrons. The following is an example of a formula students may offer:
      
      Core charge = number of protons − (total number of electrons − number of valence electrons)
      
      See Model 1 for specific answers.
      
Read This!

The valence electrons in an atom are influenced not only by the attractive power of the nucleus but also by the repulsive power of neighboring and core electrons. The shielding effect of the core electrons reduces the attractive power of the nucleus. The pulling force that a valence electron actually feels, the effective nuclear charge \( (Z') \), is much less than the nuclear charge because of this shielding effect of core electrons. Although the effective nuclear charge is difficult to calculate directly due to complex quantum effects within the atom, it is approximately equal to the core charge of an atom.

6. Use arrows like those below to illustrate the relative strength of the effective nuclear charge on the valence electrons in the atoms of Model 1.

   → → →

   Increasing effective nuclear charge = thicker arrow

   See Model 1.

7. According to Model 1, what happens to the effective nuclear charge of atoms as you move from left to right on the periodic table?

   As you move from left to right across a row of the periodic table, the effective nuclear charge increases.
8. Explain the trend as you move across a row of the periodic table for each of the following atomic properties using your understanding of effective nuclear charge.

a. Atomic radius

Atoms in a row have the same number of electron shells. As you move from left to right, the number of protons increases. Because the number of core electrons remains constant, the effective nuclear charge increases and the valence electrons experience a stronger nuclear pull and the atomic radius decreases.

b. Ionization energy

As you move across a row of the periodic table, the effective nuclear charge increases due to a larger number of protons in the nucleus but a constant number of core electrons. This increases the pull on the valence electrons by the core of the atom making it more difficult to remove an electron. Thus, the ionization energy increases.

c. Electronegativity

As you move across a row of the periodic table, the effective nuclear charge increases due to an increasing number of protons in the nucleus and a constant number of core electrons. This increases the attractive power of the core of the atom. Thus, the atoms to the right in a row are able to pull on bonding electrons with more force than atoms to the left in a row.

Model 2 – The Alkali Metals

<table>
<thead>
<tr>
<th>Atomic Number</th>
<th>Core Charge</th>
<th>Atomic Radius</th>
<th>1st Ionization Energy</th>
<th>Electronegativity</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lithium</td>
<td>3</td>
<td>+1</td>
<td>152 pm</td>
<td>520 kJ/mole</td>
</tr>
<tr>
<td>Sodium</td>
<td>11</td>
<td>+1</td>
<td>186 pm</td>
<td>496 kJ/mole</td>
</tr>
<tr>
<td>Potassium</td>
<td>19</td>
<td>+1</td>
<td>227 pm</td>
<td>419 kJ/mole</td>
</tr>
</tbody>
</table>

9. Describe the relationship of the three elements in Model 2 with regards to their relative positions on the periodic table.

The three elements of Model 2 are found in the same column of the periodic table, and appear in the order shown.

10. Draw a circle and lightly shade the area representing the core of each atom in Model 2.

See Model 2.

11. Refer to a periodic table to complete the table in Model 2.

See Model 2.
12. According to Model 2, what happens to the effective nuclear charge of atoms as you move from the top to the bottom of a column on the periodic table?

*The effective nuclear charges for all atoms in a column of the periodic table are identical.*

13. Is effective nuclear charge a factor in how periodic properties change within a column of the periodic table? If no, propose another factor that would influence the attractive forces between the nucleus and valence electrons in an atom.

*No, since the effective nuclear charge for all elements in a column is the same, it cannot explain the gradual change in properties observed as you move down a column. Instead, the distance between the valence electrons and the nucleus causes property changes.*

14. Use arrows like those in Model 1 to illustrate the relative strength of attraction between the nucleus and the valence electrons in the atoms of Model 2.

*See Model 2.*

15. Explain the trend as you move down a column of the periodic table for each of the following atomic properties using your understanding of effective nuclear charge and any other factors that might govern periodic trends.

   a. Atomic radius

   *Although the number of protons increases as you go down a column of the periodic table, the effective nuclear charge does not because the number of core electrons also increases. Therefore, it is the distance between the valence electrons and the nucleus that governs the trend. Atoms with more electrons are larger because electrons are located farther from the nucleus as the number of shells increases.*

   b. Ionization energy

   *Although the number of protons increases as you go down a column of the periodic table, the effective nuclear charge does not because the number of core electrons also increases. Therefore, it is the distance between the valence electrons and the nucleus that governs the trend. As the valence electron gets farther from the nucleus, the attractive power of the nucleus decreases and it takes less energy to remove an electron.*

   c. Electronegativity

   *Although the number of protons increases as you go down a column of the periodic table, the effective nuclear charge does not because the number of core electrons also increases. Therefore, it is the distance between the valence electrons and the nucleus that governs the trend. As the atom gets larger, the nuclear pull exerted on bonding electrons decreases.*

16. Answer the question posed in the *Why?* box at the start of this activity: “If more protons in the nucleus pull electrons in tighter, then why don’t atoms get smaller when going down a column?”

*Though there are more protons as you move down a column, there are also more core electrons that shield the valence electrons from feeling the pull of those protons. Therefore, the effective nuclear charge is fairly constant within a column.*
17. Refer to the graph in Model 3.
   a. What property of atoms does the graph illustrate?
      
      The graph illustrates the ionization energies of atoms of different elements.
   
   b. There are four lines on the graph. What do the lines represent?
      
      The four lines on the graph represent the elements in the first four periods of the periodic table.
   
   c. Two elements, hydrogen and helium, have been labeled on the graph. Label the remaining elements on the graph. You should finish with krypton.
      
      See Model 3.

18. How does the graph in Model 3 illustrate the periodic trend for ionization energy as you move across a row of the periodic table?

   Each line on the graph represents a period. The lines, in general, have positive slopes showing the trend that ionization energy increases as you move across the table.

19. How does the graph in Model 3 illustrate the periodic trend for ionization energy as you move down a column of the periodic table?

   Each of the four lines starts and remains lower than the ones above it. This shows that as you go down the periodic table, ionization energies decrease.
20. Are either of the trends you described in Questions 18 and 19 perfect? If not, where are the "blips" or discrepancies in the trends? Do they always occur in the same place?
   
   There are regular discrepancies in the trends for the third element in the row and the fifth element in the row.

21. Consider the elements represented in Model 3 and how their discrepancies may relate to electron configuration. At what points would discrepancies occur? Explain.
   
   When you get to the third element in the row, you are placing the last electron in a new subshell—the p subshell. When you get to the fifth element in the row, the p subshell is half-full and you are beginning to double up electrons in the orbitals.

22. Consider everything you know about the attractive and repulsive forces in the atom and propose an explanation for the discrepancies you see in Model 3.
   
   The discrepancy at the third element in the row occurs because the p subshell is slightly larger than the s subshell on average, so the electron going into that shell is further from the nucleus. This makes that electron easier to remove. The discrepancy at the fifth element is attributable to repulsive forces between electrons. The fifth electron is joining another electron in a p orbital. The electrons are crowded and repel one another, making it easier to remove one of the paired electrons.

23. Explain why aluminum and gallium have almost identical 1st ionization energies even though they are in different periods. (Based on the spacing of the other lines, gallium should be much lower.)
   
   Aluminum is in the third period of the table, which does not have transition metals. The difference between aluminum and magnesium is one proton and a new p subshell in the atomic structure. Gallium, however, is in the fourth period of the table, which does include transition metals. The difference between calcium and gallium is 11 protons, a filled d subshell and a new p subshell. All of those additional protons in the nucleus make it more difficult to remove the electron in the p subshell, so gallium’s ionization energy is much higher than would be predicted by the previous periods.