1.7 Periodic Trends in Atomic Properties

The periodic table contains a great deal of information and presents numerous trends and patterns. In this section, you will explore atoms more closely. You will look at additional atomic properties and see how they are all connected. You will also look for trends among the properties.

atomic radius a measurement of the size of an atom, usually expressed in picometres (pm); the distance from the centre of an atom to the outermost electrons

Atomic Radius

An individual atom is incredibly small. Its radius is expressed in picometres (one picometre = 1 pm = 1 × 10^{-12} m).

Measuring the radius of a tennis ball would be a fairly straightforward task. We could easily measure the diameter with a ruler and divide by 2 to obtain the radius. The ball has definite boundaries and our measuring device is effective for the task.

How might we define and measure the size of something as tiny as an atom? This is a difficult task for a few reasons. The boundaries of an atom are not as definite and it is difficult to determine where the electron orbits end for an atom. Also, it is so incredibly small that we would need a very tiny ruler! In fact, a direct measurement is impossible, but scientists are able to determine the atomic radius indirectly.

Chemists define the radius of an atom in a few different ways. Simply, the **atomic radius** of an atom is defined as the distance from the nucleus to just beyond the outermost electrons (valence electrons). In a diatomic molecule (such as nitrogen, N_2, or oxygen, O_2), the atomic radius is defined as the distance between the two nuclei, divided by 2 (**Figure 1**).

Figure 2 shows the atomic radii of the elements in the periodic table. Look at the relative sizes of the atoms and examine the periodic trends associated with this property.

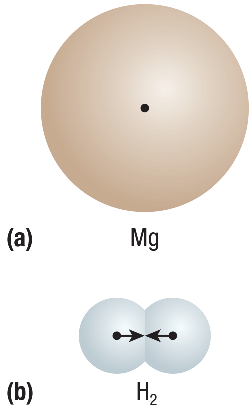
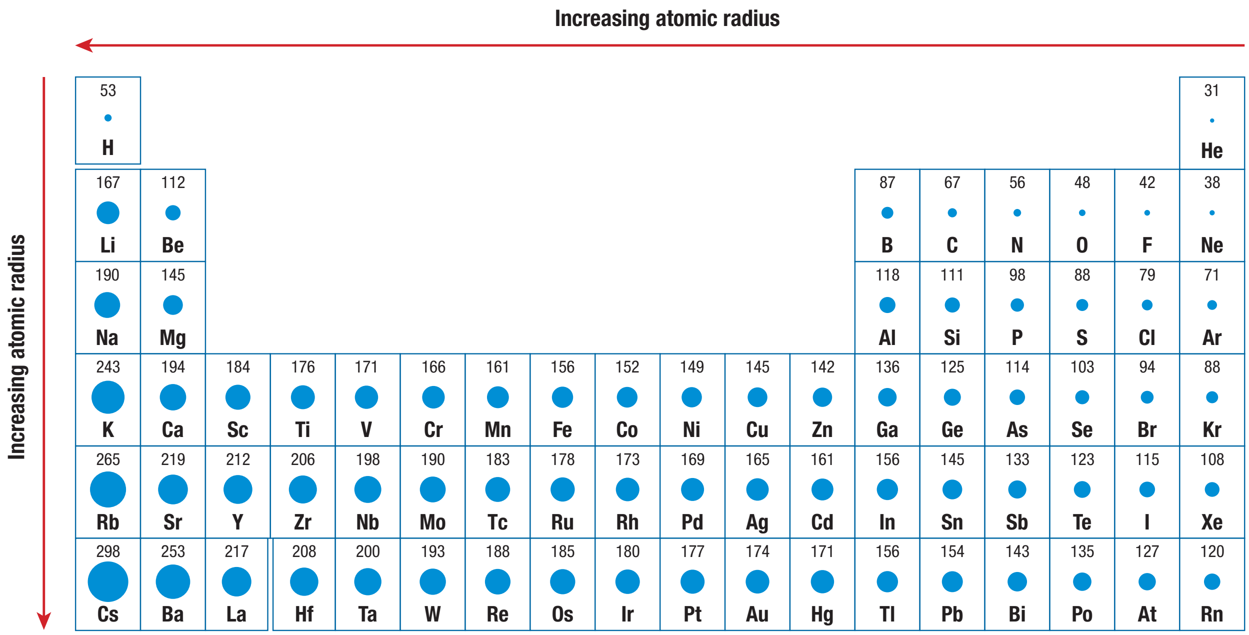


Figure 1 The atomic radii of (a) magnesium, an unbonded atom, and (b) hydrogen, a pair of bonded atoms



53 H																	31 He
167 Li	112 Be											87 B	67 C	56 N	48 O	42 F	38 Ne
190 Na	145 Mg											118 Al	111 Si	98 P	88 S	79 Cl	71 Ar
243 K	194 Ca	184 Sc	176 Ti	171 V	166 Cr	161 Mn	156 Fe	152 Co	149 Ni	145 Cu	142 Zn	136 Ga	125 Ge	114 As	103 Se	94 Br	88 Kr
265 Rb	219 Sr	212 Y	206 Zr	198 Nb	190 Mo	183 Tc	178 Ru	173 Rh	169 Pd	165 Ag	161 Cd	156 In	145 Sn	133 Sb	123 Te	115 I	108 Xe
298 Cs	253 Ba	217 La	208 Hf	200 Ta	193 W	188 Re	185 Os	180 Ir	177 Pt	174 Au	171 Hg	156 Tl	154 Pb	143 Bi	135 Po	127 At	120 Rn

Figure 2 Atomic radius is a periodic trend. What trends can you observe and explain in the first six periods?

Every electron in an atom experiences a force of attraction toward the nucleus, known as the **effective nuclear charge.** The ideas of effective nuclear charge and energy levels help us to explain trends in atomic radius. As you move down a group, notice that the atomic radius increases. This is a result of adding another energy level as you progress from one period to the next. Each additional energy level is a greater distance from the nucleus. This occurs because the electrons of the inner levels shield the outer electrons from the full charge of the nucleus. As a result, the outer electrons are not as strongly attracted by the nucleus, resulting in a larger atomic radius.

Atomic radii decrease as you move from left to right across a period. This trend is a result of the increasing positive charge of the nucleus. As you move across a period, each element has one more proton and one more electron than the element before it. Since the additional electrons are being added to the same energy level, no screening of the nucleus occurs. Therefore, as the nuclear charge increases moving across a period, the attraction for electrons also increases. The increased attraction pulls the electrons closer to the nucleus resulting in a smaller atomic radius. As result of this trend, the noble gases are the smallest atoms in their respective periods.

effective nuclear charge the net force experienced by an electron in an atom due to the positively charged nucleus

Ionic Radius

When an atom gains or loses one or more electrons, it forms an ion. When sodium's one valence electron is removed, the remaining entity is a positively charged sodium ion, Na^+. When chlorine gains one more valence electron, it becomes a negatively charged chloride ion, Cl^-.

The size of an ion is described by the **ionic radius**. Like atomic radius, it is measured in picometres (pm). Is the ionic radius larger or smaller than the atomic radius of the same element? To answer this, we once again need to consider the forces acting on the valence electrons.

ionic radius a measurement of the size of an ion, usually expressed in picometres (pm); the distance from the centre of an ion to the outermost electrons

First, let us consider what happens when an atom of sodium becomes a sodium ion. With the removal of the only valence electron, there is one fewer electron orbit. A sodium ion has an ionic radius of 95 pm compared to the atomic radius of 190 pm for a neutral sodium atom (**Figure 3(a)**). Positive ions (cations) are always smaller than their original neutral atoms. We can explain this by referring to the force of attraction exerted by the nucleus. The force is now shared among fewer electrons, so is slightly stronger on each one. The net result is that the nucleus pulls each electron a little closer.

Negative ions (anions), on the other hand, are always bigger than their original neutral atoms. For example, a chloride ion has an ionic radius of 181 pm, whereas a chlorine atom has an atomic radius of 79 pm (**Figure 3(b)**). When an atom gains an electron, repulsion among the electrons increases while the nuclear charge remains the same. This results in a larger ionic radius. The theoretical explanation for this is that the effective nuclear charge is now shared among more electrons. The force of attraction exerted by the nucleus is therefore slightly weaker on each electron. The nucleus cannot hold each electron quite so close to itself. 🌐

WEB LINK

To see a comparison of atomic radii and ionic radii for all the elements,

GO TO NELSON SCIENCE

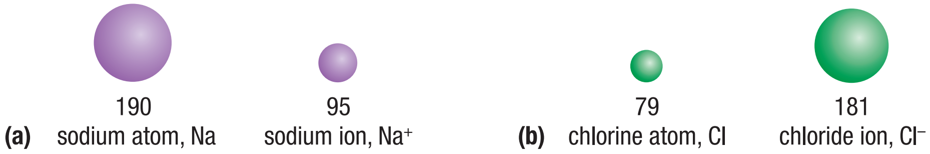


Figure 3 Notice how a positive ion is larger than its atom, whereas a negative ion is smaller than its atom. Suggest a reason for this observation.

Mini **Investigation**

Organizing Aliens

Skills: Performing, Observing, Analyzing, Communicating

Equipment and Materials: a set of cards with drawings of aliens with various characteristics

1. Your teacher will distribute an almost-complete set of cards to each group of students. These cards have drawings of aliens with various features (**Figure 4**).
2. In your group, organize the aliens in a meaningful arrangement or table.
3. Predict what the missing alien(s) would look like, and where it/they would fit in your table.
4. Obtain the missing card(s) from your teacher and compare your prediction with the actual card(s).

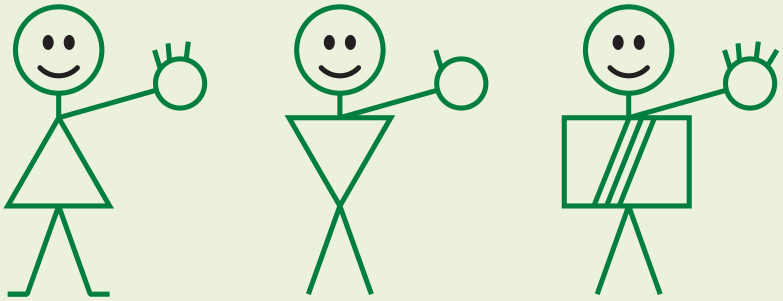


Figure 4 Classify the aliens according to trends in their characteristics.

A. While you were arranging the aliens, what patterns did you notice first? T/I

B. How many groups did your alien table have? T/I

C. How many periods did your alien table have? T/I

D. How did your prediction compare with the missing card(s)? T/I

E. How is your alien table analogous to the formation and usefulness of the periodic table? T/I

F. Think of another example in which items are organized by their properties into some type of a table. A

Ionization Energy and Electron Affinity

Valence electrons are bound to an atom by their attractive force to the nucleus. Removing electrons requires energy. Gaining electrons often releases energy.

Ionization Energy

Have you ever accidentally placed a metal spoon or a sheet of aluminum foil in a microwave oven? The resulting sparks are dramatic and potentially dangerous. Yet you can heat water molecules in a plastic or ceramic cup without incident. The sparks that we see are a result of electrons leaving the metal atoms. This is evidence that metal has a weaker hold on its electrons than do the hydrogen and oxygen atoms in water. The hold that an atom has on its electrons is an important periodic property known as ionization energy.

ionization energy the quantity of energy required to remove an electron from an atom or ion in the gaseous state

Ionization energy is the quantity of energy required to remove a single valence electron from an atom or ion in the gaseous state. If an electron is removed from any metal atom, a positive ion (cation) is formed. For example, it takes 520 kJ/mol of energy to ionize lithium, Li.

General ionization equation:

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

Specific example:

$Li(g) + 520\ \text{kJ/mol} \rightarrow Li^+(g) + e^-$

The unit for ionization energy is kJ/mol, or kilojoules per mole. The kilojoule (kJ) is a unit of energy (1 kJ = 1000 J) and the mole (mol) is a standard quantity of a substance. You will learn much more about the mole in Unit 3.

As there may be more than one electron that can possibly be removed, we usually specify which electron we are dealing with. The first ionization energy is the energy required to remove the most loosely held electron from an atom or ion (**Figure 5**). The second ionization energy is the energy required to remove the next most loosely bound electron, and so on.

Investigation **1.7.1**

Graphing Periodic Trends (p. 45)
In this investigation you will create a graph of first ionization energy and atomic radius against atomic number, using provided data, to look for periodic trends.

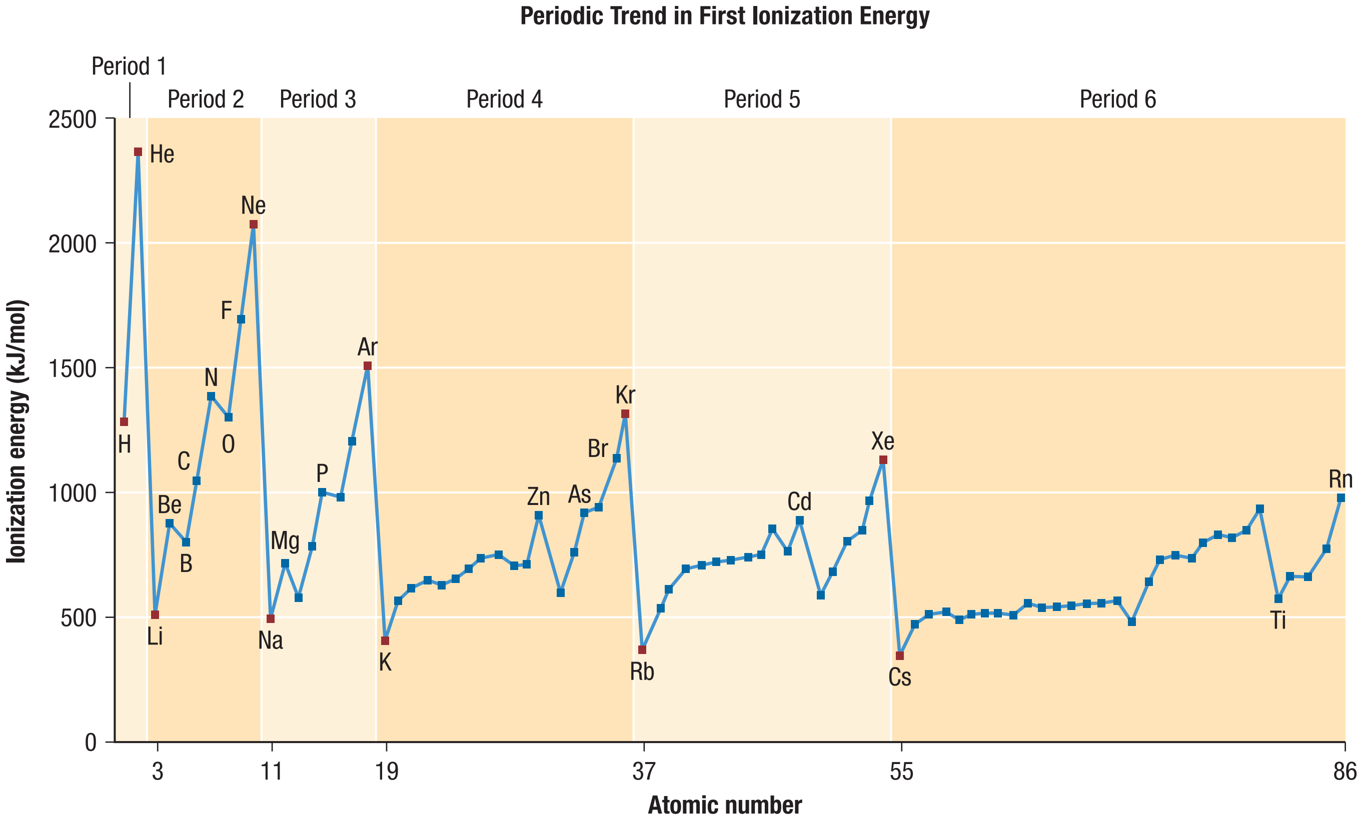


Figure 5 First ionization energy shows periodic trends. What trends can you observe and explain?

Figure 5 shows that, as we move down a group, the ionization energies tend to decrease. Hydrogen, H, has an ionization energy of 1312 kJ/mol; lithium, Li, has an ionization energy of 520 kJ/mol; and sodium, Na, has an ionization energy of 496 kJ/mol. Scientists reason that less energy is required to remove an electron as we move down a group because the force of attraction between the electron and the nucleus decreases as the atomic radius increases. Simply put, the farther away an electron is from the nucleus, the easier it is to remove it.

Look from left to right across a period. Notice that the ionization energy tends to increase. For example, the ionization energy of calcium (atomic number 20) is 590 kJ/mol; the ionization energy of arsenic, As, is 947 kJ/mol; and the ionization energy of bromine, Br, is 1140 kJ/mol. This trend is supported by our explanation that, as atomic radius decreases, the pull on the outermost electrons increases and thus it is harder to remove them. (Remember that atoms are generally smaller on the right side of the periodic table than on the left side.) Following this pattern, the ionization energies for the noble gases are extremely large. It is very difficult to remove any of their electrons.

Electron Affinity

electron affinity the energy change that occurs when an electron is added to a neutral atom in the gaseous state

If a neutral atom gains one or more electrons, it becomes negatively charged. The **electron affinity** for an element is defined as the energy change that occurs when an atom in the gaseous state gains an electron. Electron affinity, like ionization energy, is measured in kJ/mol.

When considering electron affinity, chemists need to be specific about the number of electrons that are acquired. The energy released when an atom acquires one electron is referred to as the first electron affinity. The energy released when a second electron is acquired is the second electron affinity, and so on. For example, 349 kJ/mol of energy is released when an atom of chlorine gas gains an electron.

General electron affinity equation:

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

Specific example:

$Cl(g) + e^- \rightarrow Cl^- (g) + 349\ \text{kJ/mol}$

Table 1 shows the electron affinities for the first 20 elements. What trends can you observe? The electron affinities of helium, beryllium, nitrogen, magnesium, and argon are all less than zero. This means that energy is not released when these elements gain an electron. On the contrary, energy is needed. How can we explain this? When an electron is added to an atom, it is attracted by the nucleus and also repelled by the atom's electrons. If the attractive force is greater, the electron is received, an anion is formed, and energy is released. If, however, the repulsive force is greater, energy is required to add the electron.

Table 1 Electron Affinities and Atomic Radii of the First 20 Elements

Element	Electron affinity (kJ/mol)	Atomic radius (pm)
H	73	53
He	0	31
Li	60	167
Be	0	112
B	27	87
C	154	67
N	7	56
O	141	48
F	328	42
Ne	0	38
Na	53	190
Mg	0	145
Al	43	118
Si	134	111
P	72	98
S	200	88
Cl	349	79
Ar	0	71
K	48	243
Ca	2	194

As you investigate the patterns in the periodic table, you might notice that the electron affinity for elements decreases as you move down a group and it increases as you move across a period. Electron affinity shows the same trends as ionization energy. What is the theoretical explanation for the trend in electron affinity values? As you move to the right across the periodic table, the number of valence electrons increases and the atomic radius decreases. The force of attraction between the nucleus and the valence electrons increases, so more energy is released when a new electron is acquired.

LEARNING **TIP**

Using Analogies in Chemistry

An analogy might help you understand ionization energy and electron affinity better. Imagine paper clips (electrons) attached to a strong magnet (nucleus). You have to use energy to pry the paper clips away. In the same way, it takes energy—ionization energy—to remove the electrons from the atom. When you bring a paper clip close to the magnet, it jumps onto the magnet with a snap. The snap represents electron affinity. The louder the snap, the greater the electron affinity.

Interpreting Data

In this section, data are provided in several formats: as a figure with representative drawings (Figures 2 and 3), as a graph (Figure 5), and in a table (Table 1). You may find one format easier to use than the others. If you have difficulty interpreting data in the future, you could convert it into your preferred format.

1.7 Summary

- Atomic radius, ionic radius, ionization energy, and electron affinity are all atomic properties that exhibit periodic trends (**Figure 6**).
- The trends for atomic radius can be explained using the concept of effective nuclear charge, the attractive force of a positively charged nucleus on the outermost electrons in an atom.
- The periodic trends for ionization energy and electron affinity can be explained using the concept of effective nuclear charge and the trends in atomic radius.

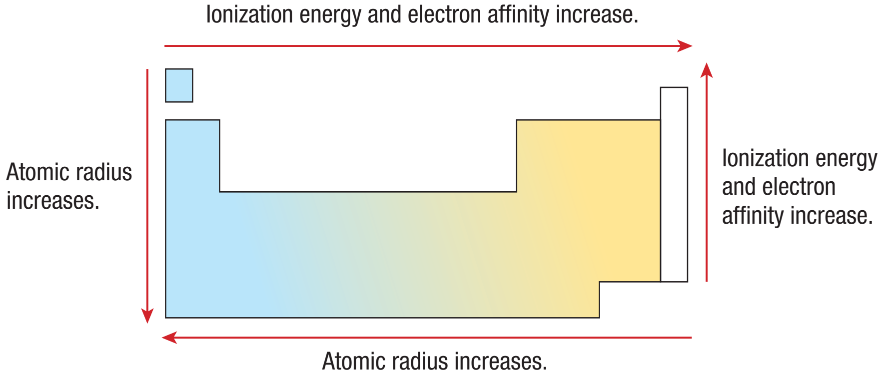


Figure 6 Periodic trends

1.7 Questions

1. Refer to Figure 2 to identify which element in each of the following groups has the largest atomic radius: K/U
   (a) halogens
   (b) alkali metals
   (c) noble gases
2. Arrange the following atoms in order of increasing atomic radius (smallest to largest): K, Rb, Cs, F, Li, C T/I
3. Predict which has the larger radius: a magnesium atom or a magnesium ion. Explain. T/I
4. Compare and contrast ionization energy and electron affinity. K/U
5. For each of the following pairs of atoms, state which would have
   (a) a larger size
   (b) a higher ionization energy
   (c) a lower electron affinity

   Support each of your predictions with an explanation. T/I
   (i) Ca, K (ii) Al, P (iii) I, F (iv) Li, Ra
6. For each of the following pairs of entities, state which entity is larger. Support each of your predictions with an explanation. T/I
   (a) S, S^{2-}
   (b) Ca^{2+}, Ca
   (c) F^{-}, Li^{+}
   (d) Cl^{-}, Br^{-}
7. Consider two isotopes of chlorine, Cl-35 and Cl-37. Would you expect the first ionization energies for these two isotopes to be the same or different? Explain your prediction. A
8. An anomaly is a deviation or departure from the normal. Using the data in this section, find an anomaly for
   (a) the periodic trend for atomic radius
   (b) the periodic trend for ionization energy

   Propose why such an anomaly might exist. T/I
9. The periodic table in Figure 2 shows the atomic radii for the first six periods. Write a hypothesis predicting the atomic radii for the following elements in Period 7: T/I A
   (a) francium
   (b) ununseptium
10. A ceramic plate has a shiny rim. The ceramic plate is mostly composed of silicon and oxygen atoms. The shiny rim is mostly silver. When the plate is heated in a microwave oven, sparks are observed. Compare the ionization energy values of silver to those of the silicon and oxygen atoms. Explain what is happening in the oven. Why are microwaveable containers usually non-metallic? Why are most pots and pans metallic? K/U A
11. Scientists are trying to synthesize element 119. Based on your knowledge of the trends in the periodic table, predict three physical properties and one chemical property for element 119. T/I A
12. Predict the atomic radii of the first three elements in Period 7. Explain your predictions. T/I A