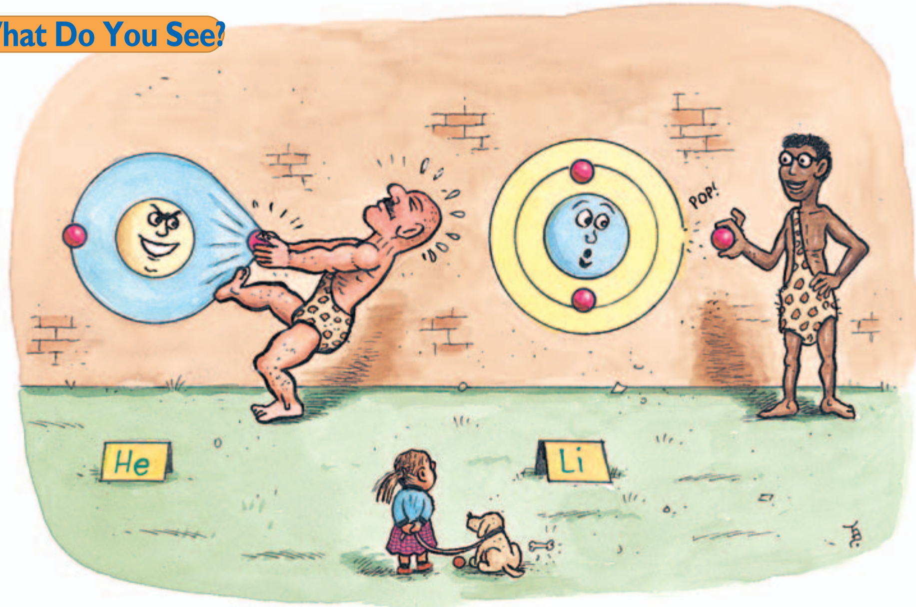
Activity 6 Atoms with More than One Electron

What Do You See?



GOALS

In this activity you will:

- View the spectra of various materials.
- Graphically analyze patterns in the amounts of energy required to remove electrons from different kinds of atoms.
- Compare trends in stability of atoms in the periodic table.
- Compare the structure of the periodic table with the patterns of levels and sublevels to which electrons can be assigned.
- Develop a shorthand notation to describe the configuration of electrons in an atom.

What Do You Think?

Niels Bohr was able to explain the spectrum of light emitted by hydrogen using a model that assigned the electron to specific energy levels. Hydrogen is a simple atom that contains only one electron. The atoms of other elements contain more than one electron.

- **How do you think an increase in the number of electrons would impact the spectrum of an atom?**

Record your ideas about this question in your *Active Chemistry* log. Be prepared to discuss your responses with your small group and the class.

Investigate

1. In *Activity 5* you observed the visible spectrum of hydrogen gas as its electron moved from a higher energy level to a lower energy level. You also explored a model that used Bohr's theory to explain this spectrum. Now it's time to look at the spectra of some other elements.

   a) Your teacher will connect a tube containing an element other than hydrogen to a high-voltage supply. Record the name of the element in your *Active Chemistry* log. Look at the spectrum of light of this element through the spectroscope.

   b) What colors do you see? Make a diagram in your log of the spectrum (pattern of colors) you see inside the spectroscope.

c) Record how this spectrum is similar to and different from the hydrogen spectrum you observed in *Activity 5*.

d) Repeat *Steps (a)*, *(b)*, and *(c)* for other samples of elements as available.

2. Spectra of such elements as helium and neon are very beautiful. However, they cannot be explained by Bohr's simple theory for the single electron in the hydrogen atom. The basic idea is still true. Light is emitted when electrons jump from a higher energy level to a lower energy level. The energy levels, however, are more complex if there are additional electrons. A more elaborate labeling of electron energy levels is necessary. In this activity you will explore the pattern of electron energy levels in atoms containing more than one electron.

When multiple electrons are present, some are easier (i.e., require less energy) to remove from the atom than others. The chart of ionization energies provides information about the amount of energy required to remove the electrons in the two highest energy levels. These are the outermost electrons and are easiest to remove. These energies are called the first and second ionization energies. They are given in units of joules. Notice that all values are multiplied by 10^{-19}.

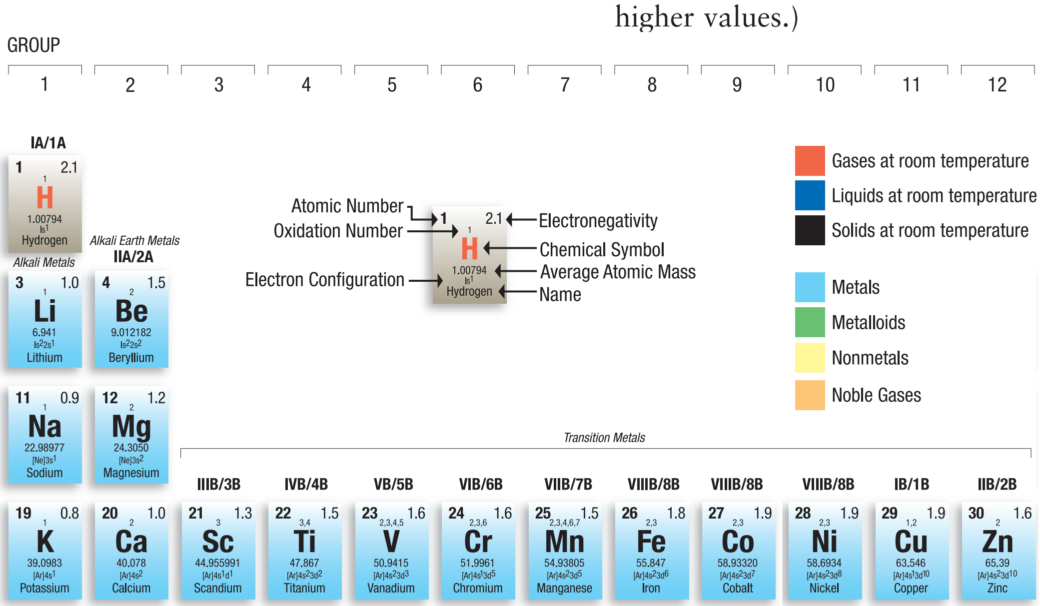
a) Make a graph that shows how the ionization energies vary with atomic number. Since the atomic numbers range from 1 to 36, label the *x*-axis with atomic numbers from 1 to 36. Since the ionization energies range from 7 to 121.2, label the *y*-axis with ionization energies from 0 to 130. Plot the first ionization energy data from the chart in one color, connecting the data points as you go along.

b) Plot the values for the second ionization energies in a different color.

c) Include a title and legend on your graph.

Atomic number	Element (symbol)	1st ionization energy J ($\times 10^{-19}$)	2nd ionization energy J ($\times 10^{-19}$)
1	H	21.8	
2	He	39.4	87.2
3	Li	8.6	121.2
4	Be	14.9	29.2
5	B	13.3	40.3
6	C	18.0	39.1
7	N	23.3	47.4
8	O	21.8	56.3
9	F	27.9	56.0
10	Ne	34.6	65.6
11	Na	8.2	75.8
12	Mg	12.3	24.1
13	Al	9.6	30.2
14	Si	13.1	26.2
15	P	16.8	31.7
16	S	16.6	37.4
17	Cl	20.8	38.2
18	Ar	25.2	44.3
19	K	7.0	50.7
20	Ca	9.8	19.0
21	Sc	10.5	20.5
22	Ti	10.9	21.8
23	V	10.8	23.5
24	Cr	10.8	26.4
25	Mn	11.9	25.1
26	Fe	12.7	25.9
27	Co	12.6	27.3
28	Ni	12.2	29.1
29	Cu	12.4	32.5
30	Zn	15.1	28.8
31	Ga	9.6	32.9
32	Ge	12.7	25.5
33	As	15.7	29.9
34	Se	15.6	34.0
35	Br	18.9	34.9
36	Kr	22.4	39.0

3. Look at the graph of the first ionization energies and answer the following questions:
   a) What kinds of patterns do you see? How could you quickly relate the shape of the graph to someone who had not seen it? If you were given a piece of blank paper and only five seconds, how would you sketch the pattern of ionization energies?
   b) Where are the ionization energies the largest? The smallest?
   c) What happens to the ionization energies as the atomic number increases?
   d) Group the elements by their ionization energies into four consecutive "periods." List the range of atomic numbers in each group.
   e) Is there any interruption in the general trend of ionization energies as the atomic number increases for a "period"? If so, describe it.
4. Look at the second colored graph line you drew.
   a) Describe how the two graphs are alike and/or different. Do you see similarities between the two graphs?
5. If a large amount of energy is needed to remove an electron from an atom, the arrangement of electrons in that atom is considered to be especially stable. Thus, a high first ionization energy means that a lot of energy must be supplied to remove an electron from an atom and that the electron arrangement in that atom is especially stable. Any element that has a larger first ionization energy than its neighboring elements has an electron arrangement in its atoms that is more stable than its neighboring elements.
   a) Which element in the first period (atomic numbers 1 and 2) has the most stable arrangements of electrons in its atoms? (Remember, you are looking for elements that have larger ionization energies than their neighbors. In reality, you are looking for peaks in your graph, not just those elements with higher values.)

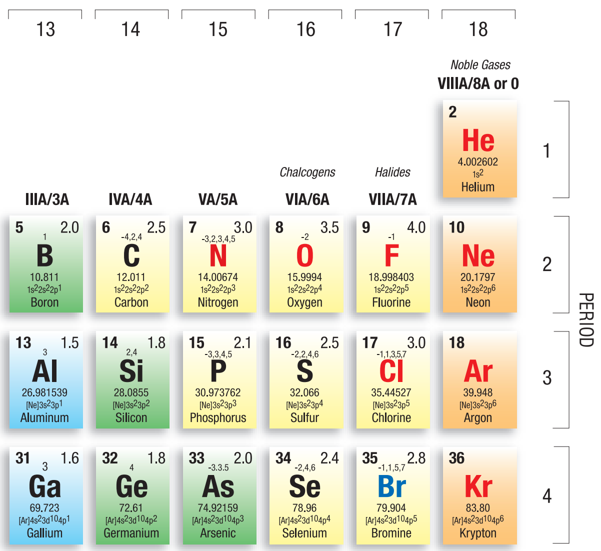


b) Which elements in the second period (atomic numbers 3 through 10) of the periodic table have the most stable arrangements of electrons in their atoms?

c) Which elements in the third period (atomic numbers 11 through 18) of the periodic table have the most stable arrangements of electrons in their atoms?

d) Which elements in the fourth period (atomic numbers 19 through 36) of the periodic table have the most stable arrangements of electrons in their atoms?

6. As mentioned earlier, the Bohr model was not able to account for the spectrum of an element containing more than one electron. A more elaborate model was needed. In this new model, the energy levels are broken down into sublevels. When these sublevels are filled, the atom exhibits a higher degree of stability. In this model, the sublevels are designated by the four letters *s*, *p*, *d*, and *f*.



The periodic table shows the atomic number, the chemical symbol, and how many electrons in an atom of each element are in each sublevel. The total number of electrons is equal to the atomic number of the element. This is because the atoms are neutral and therefore have a number of electrons equivalent to the number of protons. Remember that the atomic number is equal to the number of protons that are positively charged. This arrangement of the electrons in each sublevel will be referred to as the electron assignment or electron configuration of the element. Use this periodic table to answer the following questions:

a) In which sublevels (include number and letter) are the one electron in hydrogen and the two electrons in helium?

As you move to the second period (second row on the periodic table), each new element has one more proton in its nucleus and one more electron. The electrons must find a place to stay—an energy level and a sublevel within that energy level. As you move along in the periodic table to increasing atomic numbers, you see that the additional electrons fill the sublevel. A completed sublevel is one that is holding the maximum number of electrons allowed to it before electrons must be placed in the next higher sublevel.

b) In which region of the periodic table are electrons added in an *s* sublevel? What is the greatest number of electrons found in any *s* sublevel?

c) In which region of the periodic table are electrons added in a *p* sublevel? What is the greatest number of electrons found in the *p* sublevel?

d) In which region of the periodic table are electrons added in a *d* sublevel? What is the greatest number of electrons found in the *d* sublevel?

e) Select a column in the periodic table. (A column of elements on the periodic table is called a family or group.) Look at the electron configuration for each element within the column. Take special note of the last entry, the sublevel to which the last electron in an atom of each element in that column is added. What do all of these sublevels have in common? How many electrons are in these particular sublevels?

f) Mendeleev assigned elements to the same column of the periodic table because the elements had similar properties, both physical and chemical. For example, elements may have had similar electrical conductivity or similar reactions with acid, or were metals. These were the properties that you explored in *Activity 2*. How, then, does the number and location of the electrons in the outermost sublevel relate to chemical properties? You can now acknowledge that electrons (as opposed to the nucleus) are the key to the chemical properties of elements. Write this statement in your log using your own words.

7. At the beginning of this activity, you constructed a graph of the ionization energy versus the atomic number. If you take this graph and rotate it 90°, you will find that the graph reminds you of the periodic table, constructed by Mendeleev because of similar chemical and physical properties of elements.

a) What is the relationship between ionization energies and the rows of the periodic table?

Chem Talk

Chem Words

ionization energy: the amount of energy needed to totally remove an electron from an atom.

ions: atoms that have lost or gained electrons.

A PERIODIC TABLE REVEALED

Ions and Ionization Energy

In the table in the *Investigate* section, the amount of energy required to remove an electron from an atom was called **ionization energy**. Atoms are neutral. That is, the number of electrons is equal to the number of protons. However, atoms can gain or lose electrons. Atoms that have lost or gained electrons are called **ions** and thus the energy used to remove the electrons is known as the ionization energy. A sodium (Na) ion is formed when a sodium atom loses an electron:

$$Na + \text{energy} \longrightarrow Na^{+} + e^{-}$$

The energy required to remove a single electron from the highest occupied energy level is called the first ionization energy. The energy needed to remove a second electron from the same atom, after the first one has already been removed, is called the second ionization energy. For another example, look at the removal of two electrons from a calcium (Ca) atom:

$$Ca + \text{energy} \longrightarrow Ca^{+} + e^{-}$$

This equation represents the 1st ionization energy.

$$Ca^{+} + \text{energy} \longrightarrow Ca^{2+} + e^{-}$$

This equation represents the 2nd ionization energy.

The second electron to be removed from the nucleus is more tightly bound. This is because of a greater electrostatic attraction to the positively charged nucleus. Therefore, it takes more energy to remove this electron. The second ionization energy is always higher than the first.

Measuring the atomic radii of the elements correlates with the ionization energies. As you go across a row of the periodic table, the atomic radius of each element becomes smaller. Consider this in light of the ionization energies. Moving from left to right across a row in the periodic table, the ionization energies increase. The nuclear charge (the number of protons) is increasing so the electrons are held more tightly. As the electrons are held more tightly, you would expect the size of the atoms to decrease.

Also, as you go down a column, the atomic radius for each element becomes larger. This is because the electrons are now being placed in orbitals farther from the nucleus.

Electron Configuration and Energy Levels

As you discovered, the Bohr model was not able to account for the spectrum of an element containing more than one electron. In the new model you investigated, the energy levels are broken down into sublevels. This arrangement of the electrons in each sublevel is called the electron assignment or **electron configuration** of the element. When these sublevels are filled, the atom exhibits a higher degree of stability. The sublevels are designated by the four letters *s, p, d,* and *f.* The letters come from the words that the early scientists used to describe some of the observed features of the line spectra. The sublevels are governed by the following rules:

(i) The first energy level (corresponding to E_1 in *Activity 5*) has only

Chem Words

electron configuration: the arrangement of the electrons of an atom in its different energy sublevel(s).

one type of orbital, labeled 1*s*, where 1 identifies the energy level and *s* identifies the orbital.

(ii) The second energy level (corresponding to E_2 in *Activity 5*) has two types of orbitals (an *s* orbital and *p* orbitals) which are labeled the 2*s* and 2*p* orbitals.

(iii) The third energy level (corresponding to E_3 in *Activity 5*) has three types of orbitals (an *s* orbital, *p* orbitals, and *d* orbitals) and are labeled as the 3*s*, 3*p*, and 3*d* orbitals.

(iv) The number of orbitals corresponds to the energy level you are considering. For example: E_4 has four types of orbitals (*s, p, d,* and *f*); E_5 has five types of orbitals (*s, p, d, f,* and *g*).

(v) The *s* orbital has a maximum of two electrons. The *p* orbitals have a maximum of six electrons. The number of electrons is indicated by a superscript following the orbital designation. For example, $2p^5$ means that in the 2nd energy level and the *p* orbitals there are five electrons. $1s^2$ means that in the 1st energy level and the *s* orbital there are two electrons.

Electron configuration is determined as follows. Electrons are placed into the lowest energy levels first and "built up" from there. For example, a neutral carbon atom has six electrons. The first electron is placed in the 1st energy level and the s orbital. The second electron is placed in the 1st energy level and the s orbital. You can represent these two electrons as $1s^2$. The 1*s* orbital is now filled. There are no more orbitals in the 1*s* level. The next two electrons are placed in the 2nd energy level and the *s* orbital. You can signify these two electrons as $2s^2$. The 2nd energy level also has the *p* orbitals. The final two electrons are placed in the 2nd energy level and the *p* orbitals. You can signify these two electrons as $2p^2$. The resulting configuration is $1s^22s^22p^2$.

The electron configuration for Argon (Ar, with 18 electrons) is:

$1s^22s^22p^63s^23p^6$. To be sure that all 18 electrons are accounted for, add the superscript numbers to see that they do add up to 18 ($2 + 2 + 6 + 2 + 6 = 18$).

The electron configuration for arsenic (As, with 33 electrons) is:

$1s^22s^22p^63s^23p^64s^23d^{10}4p^3$ or a shorthand designation as [Ar] $4s^23d^{10}4p^3$.

The [Ar] simply implies the electron configuration is the same as argon up to the point just before $4s^2$.

Example:

What is the element with an electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$?

You can find the number of electrons by adding the electrons in each orbital (2 in the 1*s* orbital, 2 in the 2*s* orbital, 6 in the 2*p* orbital, etc.) That is, $2 + 2 + 6 + 2 + 6 + 1 = 19$. This element has 19 electrons. Referring to the periodic table, you find that potassium (K) has 19 electrons and is the element with that electronic configuration.

Stability is an important feature for all matter. Remember the excited electron of the hydrogen atom? If the electron were in energy level 3, it would drop down to energy level 2 and give off a specific wavelength of light. Alternatively, the electron in energy level 3 could drop down to energy level 1 and give off a different, specific wavelength of light. The word "excited" is used to describe an electron that has absorbed enough energy to move to a higher energy level, before it falls back down to its original state. The electron in the **excited state** was unstable and lost energy by giving off light in order to get to a more stable form. Particles arranged in an unstable way will move to a more stable arrangement. The most stable arrangement is called the **ground state** and this is where electrons occupy the lowest orbitals possible.

Electrons: Where Are They, Really?

The atomic orbitals designated as *s, p, d,* and *f* are regions in space. Each has its own shape. They are mathematical descriptions of a *probability* of locating an electron at any given time. The quantum mechanical model of modern physics concentrates on the electron's wavelike properties. This is not easily understood in terms of your everyday experiences. The concept of wave/particle duality is simply a model. It is a highly mathematical model. You cannot see atoms and electrons and observe their behavior directly. The best you can do is construct a set of mathematical models that best fit atomic properties that you see experimentally.

Describing the location of an electron in terms of probability resulted from an idea of Werner Heisenberg, a German physicist. He proposed what is now called the Heisenberg uncertainty principle. It basically says that it is impossible to precisely determine the exact position and momentum of an electron at the same time. This is why scientists speak of orbitals instead of orbits (like in Bohr's model). They refer to average distance from the nucleus, not actual distance.

Chem Words

excited state: an electron of an atom that has absorbed enough energy to be raised to a higher energy sublevel.

ground state: the lowest energy sublevel that an electron of an atom can occupy.

Chem Words

period: a horizontal row of elements in the periodic table.

chemical group: a family of elements in the periodic table that have similar electron configurations.

If you know the position of an electron with a high degree of certainty, then you can't know its momentum. This principle is stated mathematically as:

$$(\Delta x)(\Delta p) \geq \frac{h}{2\pi}$$

where Δx is the uncertainty in the electron's position, and

Δp is the uncertainty in the momentum, and

h is Planck's constant.

Here's another way to look at this concept. In order to "see" an electron, it is necessary to shine light (a photon with a certain energy) on it. This energy will be passed on to the electron, increasing its energy and its momentum. So, the very process of looking at an electron causes it to be somewhere else!

Werner Heisenberg

The Periodic Table

In previous activities you tried to organize elements by their properties and then by their atomic number. When elements are arranged according to their atomic numbers a pattern emerges in which similar properties occur regularly. This is the periodic law. The horizontal rows of elements in the periodic table are called **periods**. The set of elements in the same vertical column in the periodic table is called a **chemical group**. As you discovered, elements in a group share similar physical and chemical properties. They also form similar kinds of compounds when they combine with other elements. This behavior is due to the fact that elements in one chemical group have the same number of electrons in their outer energy levels and tend to form ions by gaining or losing the same number of electrons.

Checking Up

1. What is an ion?
2. What is ionization energy?
3. Explain the term chemical group.
4. Name three elements in a chemical group.
5. Provide the complete electron configuration for the atom argon (Ar).

What Do You Think Now?

At the beginning of the activity you were asked:

- **How do you think an increase in the number of electrons would impact the spectrum of an atom?**

Your response might have been along the lines of "it will become more complicated, more complex." Now that you have additional information about line spectra, describe what you would see and why you would see it when you have more than one electron in an atom.

Chem Essential Questions

What does it mean?

Chemistry explains a macroscopic phenomenon (what you observe) with a description of what happens at the nanoscopic level (atoms and molecules) using symbolic structures as a way to communicate. Complete the chart below in your *Active Chemistry* log.

MACRO	NANO	SYMBOLIC
Describe what you see when an atom other than hydrogen emits light.	*Describe how electron behavior is able to account for the specific wavelengths of emitted light.*	*A symbolic structure is used to describe the electron configuration of an atom. Explain what the symbols* $1s^2$, $2s^2$, *and* $2p^5$ *tell you about the electron and the atom.*

How do you know?

Calcium and strontium are close to one another on the periodic table. What do you know about their ionization energies, their electron configurations, and their characteristics based on their position on the periodic table?

Why do you believe?

In viewing a fireworks display, how can you determine the chemicals being used?

Why should you care?

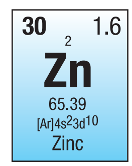
Understanding the ionization potentials of the elements helps you understand the electron configuration of the elements. It also helps you to understand why the atomic size decreases as you go from left to right and also why it increases as you go down a column. How can information about electron configurations and their relation to the periodic table be an interesting part of your game?

Reflecting on the Activity and the Challenge

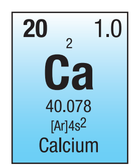
In this activity you learned that electrons in atoms are assigned not only to energy levels but also to sublevels, labeled *s, p, d,* and *f*. You have also learned that the electron configuration of atoms of all elements in the same column of the periodic table end with the same sublevel and number of electrons in that sublevel. Mendeleev organized elements into columns based on similar chemical properties. Thus, electron energy sublevels are clearly associated with chemical properties of elements and their position on the periodic table. You may wish to incorporate the information about electron configuration in your game to meet the *Chapter Challenge*.

Chem to Go

1. Write the complete configuration of electron energy levels, from 1*s* to 4*s*.
2. Consider the element boron (B) as an example.
   a) What is boron's atomic number?
   b) How many electrons does boron have?
   c) What is the complete electron configuration for boron? (Be sure to include the number and letter of the appropriate sublevels, as well as the number of electrons in each sublevel.)
3. Answer the following questions for the element zinc.



   a) What is zinc's atomic number?
   b) How many electrons does zinc have?
   c) What other elements might you expect to have chemical properties similar to zinc? Explain your choices.
4. Answer the following questions for the element calcium.



   a) What is calcium's atomic number?
   b) How many electrons does calcium have?
   c) What is the complete electron configuration for calcium? (Be sure to include the number and letter of the appropriate sublevels, as well as the number of electrons in each sublevel.)

d) What is the last sublevel (number and letter, please) to which electrons are added? How many electrons are in this sublevel?

e) Where would you expect to find calcium on the periodic table? Support your prediction with your answers from (d).

f) What other elements would you expect to have chemical properties similar to calcium? Explain your choices.

5. A chemist has synthesized a heavy element in the laboratory and found that it had an electron configuration:

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^{14} 6d^8$

a) What is the number of electrons in this element?

b) What is the atomic number?

c) What might you predict about this element?

6. If the electron configuration is given, you should be able to determine what element it is. Identify the following element from its electron configuration:

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

7. Which list of elements is arranged in order of increasing atomic radii?

a) Li, Na, Mg, Be

b) Na, Mg, Be, Li

c) Li, Be, Na, Mg

d) Be, Mg, Li, Na

8. Which is smaller, Br or Br^-? Explain your choice.

9. *Preparing for the Chapter Challenge*

Write a sentence or two to explain in words the pattern you noticed between any group and the electron configurations of the elements belonging to that group.

Inquiring Further

Determining electron configuration

In this activity, you were able to look at the electron configuration for a given element provided in the periodic table. Research other ways that the electron configuration can be determined.