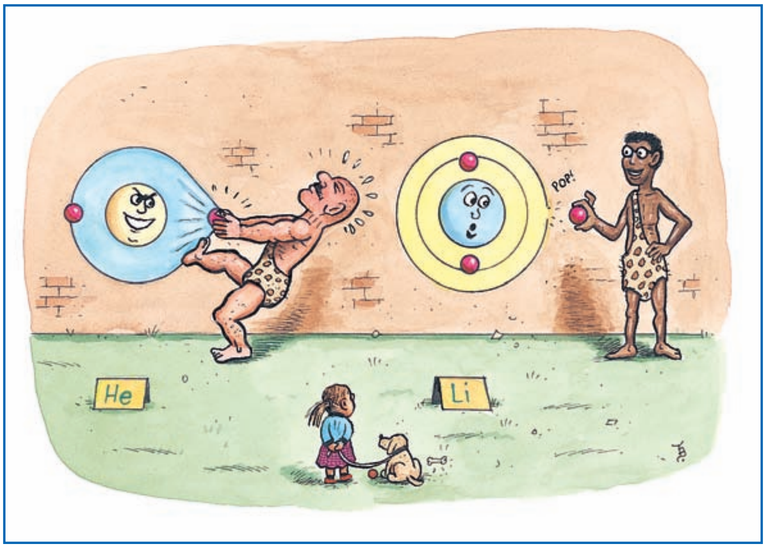
Activity 6 Atoms with More than One Electron



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?

In **Activity 5** you learned that Niels Bohr was able to explain the spectrum of light emitted by hydrogen using a model that assigned the electron to specific energy levels. Bohr was awarded a Nobel Prize in 1922 for expanding the understanding of atomic structure. He worked with hydrogen, the simplest atom, which contains only one electron. However, 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?**
- **What modifications in Bohr's model would need to take place to accommodate the extra electrons?**

Record your ideas about these questions 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 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 as many samples as your teacher demonstrates.

2. Although the spectra of such elements as helium and neon are very beautiful, 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

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

atoms containing more than one electron.

When multiple electrons are present, some are easier (i.e., require less energy) to remove from the atoms than others. The chart on the left provides information about the amount of energy required to remove the electrons in the two highest energy levels. These are the electrons that are easiest to remove. These energies are called the 1st and 2nd ionization energies, and 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 122, 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.

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?

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?

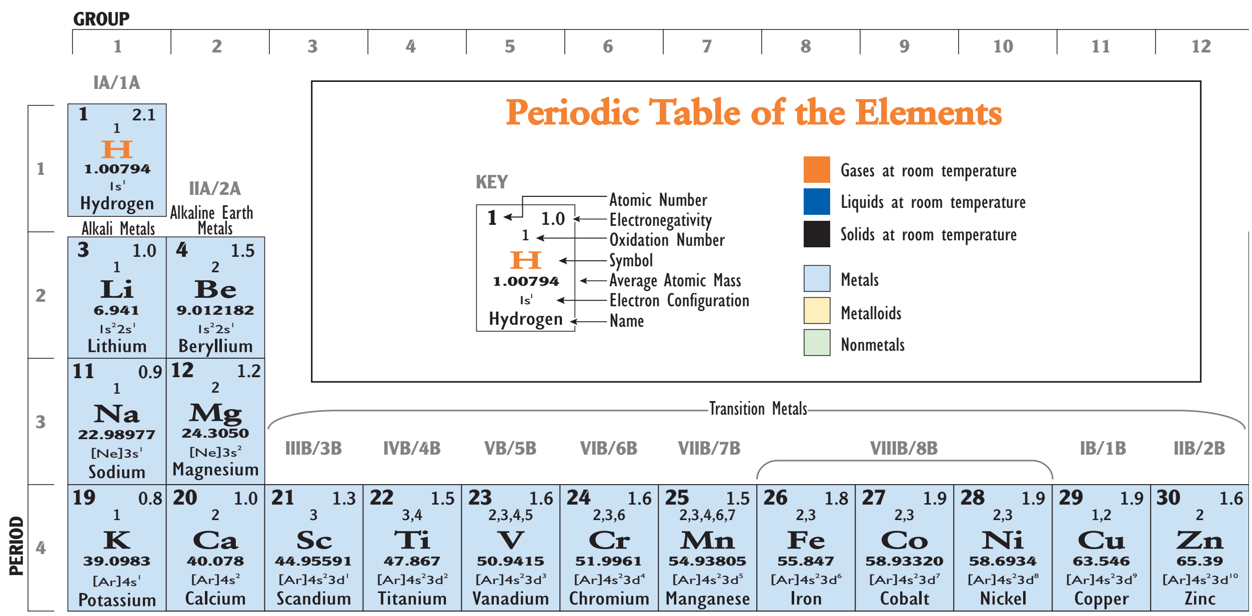
c) Which elements in the third period (atomic numbers 11 through 18) of the periodic table have the most stable arrangements of electrons in their atoms?

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. 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 what sublevel (include number and letter) are the electrons in hydrogen (1 electron) and helium (2 electrons) found?



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 reside — 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 what 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 what region of the periodic table are electrons added in a *p* sublevel? What is the greatest number of electrons found in any *p* sublevel?

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

e) In what region of the periodic table are electrons added to an *f* sublevel? What is the greatest number of electrons found in any *f* sublevel?

f) 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?

g) Mendeleev assigned elements to the same column of the periodic table because the elements had similar properties, both physical and chemical. How, then, does the number and location of the electrons in the outermost sublevel relate to chemical properties? We can now acknowledge that electrons (as opposed to the nucleus) are the key to the chemical properties of elements.

13	14	15	16	17	18
					VIIIA/8A or 0 Noble Gases
IIIA/3A	IVA/4A	VA/5A	VIA/6A Chalcogens	VIIA/7A Halides	2 He 4.002602 $1s^2$ Helium
5 2.0 3 B 10.811 $1s^22s^22p^1$ Boron	6 2.5 -4,2,4 C 12.011 $1s^22s^22p^2$ Carbon	7 3.0 -3,2,3,4,5 N 14.00674 $1s^22s^22p^3$ Nitrogen	8 3.5 -2 O 15.9994 $1s^22s^22p^4$ Oxygen	9 4.0 -1 F 18.998403 $1s^22s^22p^5$ Fluorine	10 Ne 20.1797 $1s^22s^22p^6$ Neon
13 1.5 3 Al 26.981539 $[Ne]3s^23p^1$ Aluminum	14 1.8 2,4 Si 28.0855 $[Ne]3s^23p^2$ Silicon	15 2.1 -3,3,4,5 P 30.973762 $[Ne]3s^23p^3$ Phosphorus	16 2.5 -2,2,4,6 S 32.066 $[Ne]3s^23p^4$ Sulfur	17 3.0 -1,1,3,5,7 Cl 35.4527 $[Ne]3s^23p^5$ Chlorine	18 Ar 39.948 $[Ne]3s^23p^6$ Argon
31 1.6 3 Ga 69.723 $[Ar]4s^23d^{10}4p^1$ Gallium	32 1.8 4 Ge 72.61 $[Ar]4s^23d^{10}4p^2$ Germanium	33 2.0 -3,3,5 As 74.92159 $[Ar]4s^23d^{10}4p^3$ Arsenic	34 2.4 -2,4,6 Se 78.96 $[Ar]4s^23d^{10}4p^4$ Selenium	35 2.8 -1,1,5,7 Br 79.904 $[Ar]4s^23d^{10}4p^5$ Bromine	36 Kr 83.80 $[Ar]4s^23d^{10}4p^6$ Krypton

Chem Talk

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 equivalent 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. The energy required to remove a single electron from the highest occupied energy level is called the first ionization energy, and 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.

Chem Words

ionization energy: the energy required to remove an electron from a gaseous atom at ground state.

ion: an electrically charged atom or group of atoms that has acquired a net charge, either negative or positive.

electron configuration: the distribution of electrons in an atom's energy levels.

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, *sharp*, *principal*, *diffuse*, and *fundamental*. The early scientists used these words to describe some of the observed features of the line spectra. They are governed by the following rules:

(i) The first energy level (corresponding to E_1 in **Activity 5**) has only 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 three *p* orbitals) and are labeled as 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, three *p* orbitals, and five *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 maximum number of electrons that can be contained in an orbital is two. Three p orbitals could contain a maximum of six electrons. The number of the type of electrons is indicated by superscript following the orbital designation. For example, $2p^5$ means five electrons in the $2p$ orbitals.

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 been promoted 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 to go to a more stable form. Particles arranged in an unstable way will move to a more stable arrangement.

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.

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.

Checking Up

1. What is an ion?
2. What is ionization energy?
3. What are the horizontal rows of the periodic table called?
4. Explain the term chemical group.
5. Name three elements in a chemical group.

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. How will you incorporate the information about electron configuration in your game to meet the **Chapter Challenge**?

Chemistry to Go

1. Write the complete sequence of electron energy levels, from 1*s* to 4*f*.
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 sequence 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 is the complete electron sequence for zinc? (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 zinc on the periodic table? Support your prediction with your answers from (d).
   f) 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 sequence 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:
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$.

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.