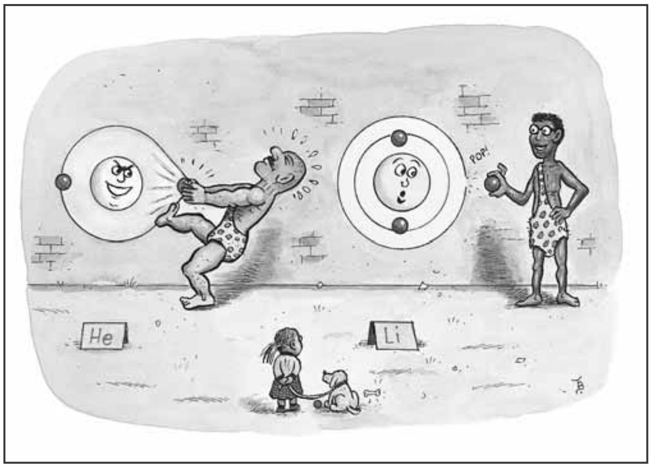
Activity 6 Atoms with More than One Electron



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.

What Do You Think?

Students will probably predict that the line spectrum would be more complex with more lines being produced. They may also think that all of the electrons would have the same energy level as that of the hydrogen electron. If that was true you would probably see the same line spectrum but it would be brighter due to reinforcement. In order to accommodate extra electrons, the atom would need to be larger and there would need to be a corresponding increase in the number of orbitals. The students that have a stronger background may recall that they have seen these electrons orbiting around the nucleus and it would look like the planets orbiting around the Sun.

Student Conceptions

This activity asks students to recognize patterns in graphical data. Students have difficulty with this if they have poor skills in making and reading graphs. Students also tend to mix up rows (periods) and columns (groups) on the periodic table, so while they are able to recognize the patterns, they may not be able to articulate correctly how these patterns are related to the organization of the periodic table. In writing electron configurations, students often have difficulty remembering that the sublevels begin to overlap at higher energy levels (4*s* sublevel is lower energy than 3*d*) due to the energy levels getting increasingly closer together as the atom nears ionization (see **Activity 5**). It is difficult to help students form any intuitive feel for electron configurations since at this level it is merely an exercise in recognizing patterns, as most high school students do not have the mathematical ability to solve the Schrödinger equation. Students may also have difficulty connecting electron configurations with the Bohr model of the previous activity. If they have been introduced to the terminology in supplementary reading, it is difficult for students to reconcile the concepts of orbitals (three-dimensional regions representing the electron density in the quantum mechanical model) and orbits (fixed trajectories in which electrons move about the nucleus in the Bohr model).

Active Chemistry The Periodic Table

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

Investigate

Teaching Suggestions and Sample Answers

1. **a–b), d)** A general description of some of the common elements are:

   Helium: 23 lines with 11 lines well defined,
   (1-light violet; 6-blue; 1-light green; 1-yellow; 1-light red; and 1-red)

   Neon: Numerous lines (75) with violet, blue, green, yellow and red.
   Majority of the lines are blue, yellow, and reds.

   Oxygen: 73 lines, 4 main colors (blue, green, yellow, and red).
   Many lines in the blue region, 3 lines in the green area, 1 or 2 in the yellow and 5 distinct lines in the red section.

   Argon: 159 lines, many lines in the blue region, some light greens and light reds.

   c) In comparison to the hydrogen spectrum we note that the spectra are not the same and in most cases they have more lines than the hydrogen line spectrum. However, the line spectrum of sodium only produces two strong lines in the yellow region.

2. Students provide graph.

Chem Tip:
In the demonstration you are trying to get students to look at the line spectrum of atoms that have more than one electron in their structure. The students may need to review the line spectrum of hydrogen for comparisons.

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.

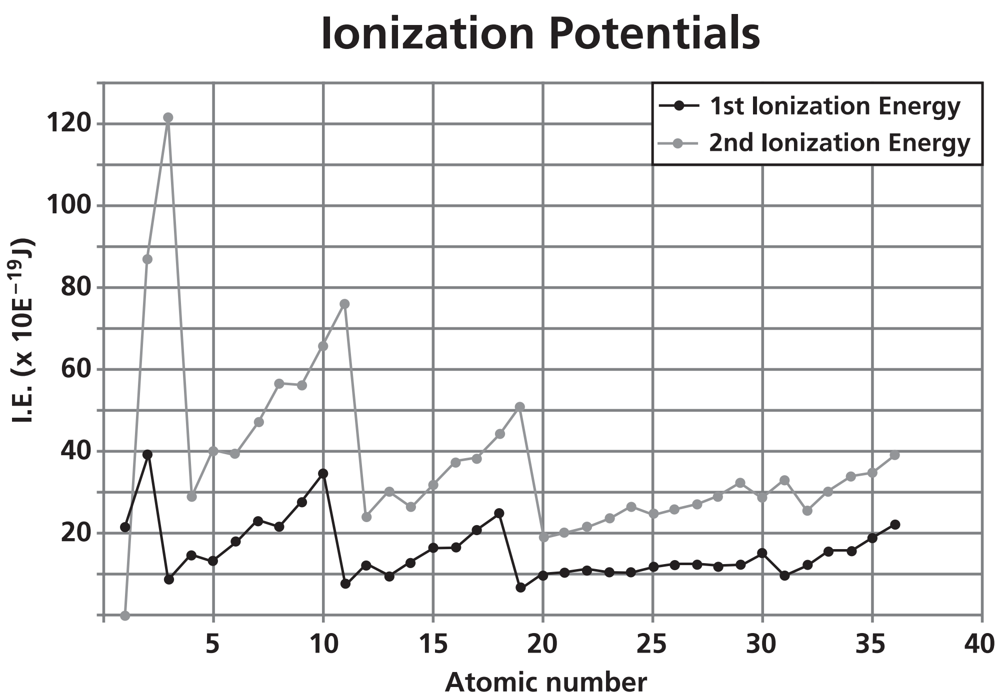
e) Is there any interruption in the general trend of ionization energies as the atomic number increases for a "period"? If so, describe it.

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



3. a) Helium has the greatest value and then the ionization drops sharply for lithium. It continues to generally increase through to neon and then it again drops sharply for sodium similar to lithium. Then the next 7 elements increase as before. Once again we see that potassium value is at the low point. If we were to continue we would assume that the ionization energy would increase again and the drop sharply for the next alkali metal.

b) The high points are noble gases and the low points are alkali metals.

c) In general the ionization energy decreases as it goes through each energy level.

d) Lithium is the lowest, beryllium is a little bit higher, slight drop for boron, then carbon a little higher, nitrogen still higher, oxygen drops a little, chlorine is higher and neon has the highest ionization energy. Period 3 would be similar. 4th period is different in that potassium and calcium are similar to the pattern of sodium and magnesium. However, the next 10 elements (scandium through zinc) have small energy change. We then see that gallium through krypton follow a pattern similar to aluminum through argon.

e) There is an interruption between beryllium and boron, between nitrogen and oxygen. The transition elements in period 4 interrupt the pattern seen in periods 2 and 3.

4. a) The second graphline looks like it is parallel to the first. The high point now is the alkali metals and the alkaline earth elements are now the low points. The energies are much greater for one line.

Active Chemistry The Periodic Table

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?

Periodic Table of the Elements

PERIOD / GROUP	1	2	3	45	5		7	8	9	10	11	12
	IA/1A											
1	1 2.1 1 H 1.00794 $1s^1$ Hydrogen	IIA/2A										
	Alkali Metals	Alkaline Earth Metals										
2	3 1.0 1 Li 6.941 $1s^2 2s^1$ Lithium	4 1.5 2 Be 9.012182 $1s^2 2s^1$ Beryllium										
3	11 0.9 1 Na 22.98977 $[Ne]3s^1$ Sodium	12 1.2 2 Mg 24.3050 $[Ne]3s^2$ Magnesium	Transition Metals									
			IIIB/3B	IVB/4B	VB/5B	VIB/6B	VIIB/7B	VIIIB/8B			IB/1B	IIB/2B
4	19 0.8 1 K 39.0983 $[Ar]4s^1$ Potassium	20 1.0 2 Ca 40.078 $[Ar]4s^2$ Calcium	21 1.3 3 Sc 44.95591 $[Ar]4s^2 3d^1$ Scandium	22 1.5 3,4 Ti 47.867 $[Ar]4s^2 3d^2$ Titanium	23 1.6 2,3,4,5 V 50.9415 $[Ar]4s^2 3d^3$ Vanadium	24 1.6 2,3,6 Cr 51.9961 $[Ar]4s^2 3d^4$ Chromium	25 1.5 2,3,4,6,7 Mn 54.93805 $[Ar]4s^2 3d^5$ Manganese	26 1.8 2,3 Fe 55.847 $[Ar]4s^2 3d^6$ Iron	27 1.9 2,3 Co 58.93320 $[Ar]4s^2 3d^7$ Cobalt	28 1.9 2,3 Ni 58.6934 $[Ar]4s^2 3d^8$ Nickel	29 1.9 1,2 Cu 63.546 $[Ar]4s^2 3d^9$ Copper	30 1.6 2 Zn 65.39 $[Ar]4s^2 3d^{10}$ Zinc

KEY

1	Atomic Number
1.0	Electronegativity
1	Oxidation Number
H	Symbol
1.00794	Average Atomic Mass
$1s^1$	Electron Configuration
Hydrogen	Name

- Gases at room temperature
- Liquids at room temperature
- Solids at room temperature
- Metals
- Metalloids
- Nonmetals

5. **a)** Helium has the highest ionization energy. Some teachers like to call this the tear-away energy.

   b) Neon has the most stable arrangement. Secondary stabilities could be assumed at beryllium and nitrogen.

   c) Argon has the most stable arrangement. Secondary stabilities could be assumed at magnesium and phosphorus.

   d) Krypton has the most stable arrangement. Secondary stabilities are at calcium and arsenic. We note that zinc is quite stable as well.

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

6. **a)** 1*s* (1st energy level with only one type of sublevel "*s*").

   b) Alkali metals and alkaline earth metals (groups IA and IIA). The maximum number of electrons allowed in an *s* sublevel is 2.

   c) Boron through neon are the next 6 elements and their electrons are placed in the *p* sublevel of the 2nd energy level. Aluminum through argon will again place their electrons in the *p* sublevel of the 3rd energy level. The greatest number of electrons found in any *p* sublevel is six.

   d) Scandium through zinc have a new group of electrons that are placed in the *d* sublevel. The *d* sublevel can hold a maximum of 10 electrons.

   e) Cerium through lutetium and thorium through lawrencium have their electrons go into the *f* sublevel. A maximum of 14 electrons can occupy the *f* sublevel.

   f) The first column is called the alkali metals and they all have an s^1 configuration. The second column is called the alkaline earth metals and they have an s^2 configuration. The halogens (group VIIA) have their last electron placed in the p^5 configuration and the noble gases have their electron placed in the p^6 configuration.

   g) Elements of the same family have the same electron configuration (but different energy level) and this is the reason why they have like chemical reactions.

Active Chemistry The Periodic Table

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.

ChemTalk

Chem Tip:
Electron configuration shows you where the last electron is placed. For example, hydrogen's electron is placed in the $1s^1$. The first 1 tells you that you are in the 1st energy level, the *s* tells you that you are placing the electron in *s* sublevel and the second 1 is telling you how many electrons you have in the *s* sublevel.

(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 2*p* 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.

ChemTalk

Checking Up

1. When an atom loses or gains an electron it is called an ion. Examples are:

   Na^0 + energy → Na^+ + e^-; F^0 + e^- → F^- + energy

2. Ionization energy is the amount of energy needed to remove an electron from an atom.

3. The horizontal rows of the periodic table are called periods.

4. A family of chemicals that have like electron configuration are also called a chemical group. An example would be the first column called the alkali family. They all have their last electron occupying the s^1 configuration.

5. The halogen group contains fluorine, chlorine, bromine, and iodine. Students may provide different examples.

Active Chemistry The Periodic Table

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?

Chemistry to Go

1. $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}$

2. a) Boron's atomic number is 5

   b) Boron has 5 electrons

   c) $1s^22s^22p^1$

3. a) Zinc's atomic number is 30

   b) Zinc has 30 electrons

   c) $1s^22s^22p^63s^23p^64s^23d^{10}$

   d) $3d$; 10 electrons

   e) Same family as cadmium and mercury. $3d^{10}$ tells you – period 3; *d* sublevel; 10 is last space in section

   f) Cadmium and mercury have $4d^{10}$ and $5d^{10}$ respectively

4. a) Calcium's atomic number is 20

   b) Calcium has 20 electrons

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:

   $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}6p^67s^25f^{14}6d^8$.

   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:

   $1s^22s^22p^63s^23p^64s^23d^6$.

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

1. 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.

c) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

d) $4s^2$; 2 electrons

e) 2nd column and commonly called the alkaline earth metals

f) Beryllium, magnesium, strontium, barium, and radium and have $2s^2$, $3s^2$, $5s^2$, $6s^2$, and $7s^2$ respectively.

5. a) 110 electrons

b) The atomic number is 110

c) This metal is radioactive because of its high atomic number. Nickel, palladium and platinum are in this group so this metal is fairly resistant to reacting with other chemicals.

6. The element would be iron (atomic number is 26)

Assessment: Graphing Skills

This assessment rubric breaks down skills in graphing data to give you an idea where work is needed or where skills are lacking. For each element, place a check mark in the appropriate box to indicate a score of 0, 1, or 2. The maximum score is 20 points if the rubric is used to score graphs of more than one set of data, or 18 points if only one set of data (when element 10 is not applicable).

Assessment Element		Score	
	0 = Incorrect or not done	1 = Work has errors or critical details are overlooked	2 = Thorough job without errors
1. Graph has a title that is relevant and informative.			
2. The variables plotted on the vertical and horizontal axes are clearly labeled.			
3. Units of measurement are provided in parentheses beside the axis labels.			
4. Independent variable is plotted along the horizontal axis. Dependent variable is plotted along the vertical axis.			
5. Appropriate ranges are chosen for axes.			
6. Appropriate intervals are chosen for axes and numbers are marked along axes.			
7. Data are plotted correctly.			
8. A line is drawn through the data points (either best-fit or connecting line, depending on instructions).			
9. Graph is neat and orderly, and can be read with ease.			
10. If two or more sets of data are plotted on the same graph, they are shown in different colors. A legend is provided to indicate which data are which.			

NOTES