

2. (b)

**Table 2:** Groups According to Number of Valence Electrons

1 Valence Electron	2 Valence Electrons	5 Valence Electrons	6 Valence Electrons	7 Valence Electrons
sodium	magnesium	nitrogen	oxygen	chlorine
cesium	beryllium	phosphorus	sulfur	bromine
lithium	calcium			fluorine
potassium	helium			

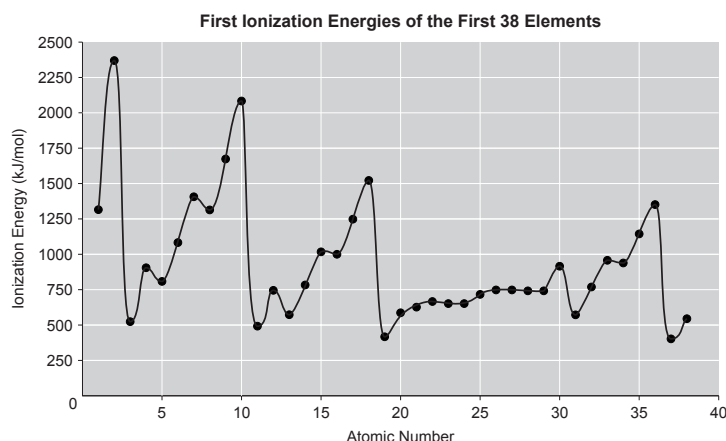
- (c) 1 valence electron: physical and chemical properties common to elements of the alkali metals (soft; metallic; react violently with water to form hydrogen and a basic solution; react strongly with oxygen; react with halogens to produce a crystalline solid)  
 2 valence electrons: physical and chemical properties common to elements of the alkaline earth metals (metallic; light; react with oxygen to form an oxide; react with water to release hydrogen; react with hydrogen to form a hydride)  
 5 valence electrons: chemical properties common to elements of Group 15 (nonmetals)  
 6 valence electrons: chemical properties common to elements of Group 16 (nonmetals)  
 7 valence electrons: physical and chemical properties common to elements of the halogens (nonmetals; extremely reactive with hydrogen and most metals)

## 1.5 TRENDS IN THE PERIODIC TABLE

### ACTIVITY 1.5.1 GRAPHING FIRST IONIZATION ENERGY

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#### Procedure



#### Analysis

- (a) • Group 1 elements have the lowest first ionization energies.  
 • Group 8 elements have the highest first ionization energies.  
 • First ionization energies of metals are lower than those of nonmetals.  
 • First ionization energies generally decrease as you move down a group in the periodic table.  
 • First ionization energies generally increase as you move from left to right across a period.

#### Synthesis

- (b) The general decrease in first ionization energies as you move down a group in the periodic table is related to the greater number of non-valence electrons between the nucleus and the valence electrons (a shielding effect). As a result, the attraction between the negatively charged electrons and the positive nucleus becomes weaker, so less energy is required to remove the first valence electron.

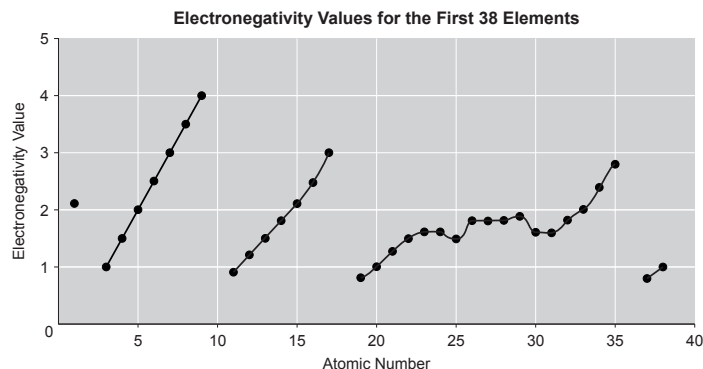
The general increase in first ionization energies across a period is related to the increasing size of the charge in the nucleus. As you move from left to right across a period, the nuclear charge increases while the shielding

effect provided by non-valence electrons remains the same. As a result the valence electrons are more strongly attracted to the nucleus, resulting in a decrease in atomic radius and an increase in the amount of energy required to remove an electron.

## ACTIVITY 1.5.2 GRAPHING ELECTRONEGATIVITY

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### Procedure



### Analysis

- (a) • Electronegativity generally increases as you go along a period, from left to right.
- Electronegativity decreases as you look down a given group on the periodic table.
- The electronegativity values for elements at the bottom of a group are lower than those for elements at the top of a group.

### Synthesis

- (b) As the atomic radii tend to decrease along a period, the attraction between the valence electrons and the nucleus increases as the distance between them decreases. The electronegativity also increases.
- (c) As the atomic radii tend to increase down a group, the attraction between the valence electrons and the nucleus decreases as the distance between them increases. The electronegativity also decreases.

## PRACTICE

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### Understanding Concepts

1. When an atom gains an electron to form a negative ion, its radius increases. In gaining an electron, repulsion among the electrons increases but there is no compensating increase in nuclear charge. The result is an enlarged electron cloud and a greater ionic radius.

When an atom loses an electron to form a positive ion, its radius decreases. In losing an electron, repulsion among the electrons decreases while nuclear charge remains the same. The result is a reduction in size of the electron cloud and a smaller ionic radius.

2.



Li



Li<sup>+</sup>

The atomic radius of the lithium atom is 152 pm. The ionic radius of the lithium ion is 68 pm.

After losing an electron, repulsion among the electrons decreases, and the result is a reduced electron cloud and a smaller ionic radius.



F



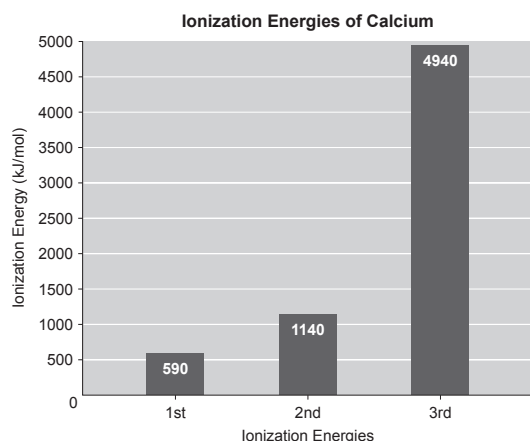
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The atomic radius of the fluorine atom is 64 pm. The ionic radius of the fluoride ion is 136 pm.

After gaining an electron, repulsion among the electrons increases, and the result is an enlarged electron cloud and a greater ionic radius.

3. (a) • First ionization energies of metals are lower than non metals.  
 • First ionization energies generally decrease as you move down a group in the periodic table.  
 • First ionization energies generally increase as you move from left to right across a period.
- (b) The general decrease in first ionization energies as you move down a group in the periodic table is related to the increase in the size of the atoms from top to bottom of a group. As the atomic radius increases, the distance between the valence electrons and the nucleus also increases. The non-valence electrons between the nucleus and the valence electrons produce a shielding effect. As a result, the attraction between the negatively charged electrons and the positive nucleus becomes weaker, so less energy is required to remove the first valence electron.
- The general increase in first ionization energies across a period is related to the decrease in the size of the atoms. As you move from left to right across a period, the nuclear charge increases while the shielding effect provided by the non-valence electrons remains the same. As a result the valence electrons are more strongly attracted to the nucleus, resulting in a decrease in atomic radius, and more energy is required to remove an electron from the atom.
- (c) Ionization energies in kJ/mol of the elements have been observed and measured.
4. (a) The most reactive metal would be cesium (some students will nominate francium, but it has not been discussed). It has the lowest ionization energy.
- (b) The most reactive nonmetal would be fluorine. It has the highest electronegativity value.

5.



6. (a) Reactivity increases with the halogens as you move up the group.
- (b) The prediction is supported — the reactivity increases as you look down this group of metals.
- (c) Halogens: as you move up the group, valence electrons remain the same, nuclear charge decreases, shielding decreases, and distance of valence electrons from the nucleus decreases. Electronegativity values increase as you move up the group, resulting in greater reactivity. Fluorine, at the top of the group, with the highest electronegativity value of all elements, is most reactive.
- Alkali Metals: as you look down the group, valence electrons remain the same, nuclear charge increases, shielding increases, and distance of valence electrons from the nucleus increases. First ionization energy values decrease as you look down the group, resulting in greater reactivity. Cesium, at the bottom of the group, with an extremely low first ionization energy value, is most reactive.
- (d) Cesium and fluorine.

## Applying Inquiry Skills

### 7. Predictions

- (a) The surface of rubidium, if cleaned, should react rapidly with atmospheric oxygen, forming an oxide layer. If rubidium is reacted with water, there will be a quick and vigorous reaction. The reasoning that supports these predictions is that as you move down a group, the first ionization energy values decrease, resulting in greater reactivity.

## Experimental Design

- (b) A small piece of rubidium is placed in water and the reactivity is observed and recorded.

### Materials

safety shield, tongs, tweezers, mineral oil, sharp edge (knife or scoop), paper towel, petri dish, large beaker, distilled water, wire gauze square, sample of rubidium

## Procedure

1. Transfer a small piece of rubidium into a petri dish containing mineral oil.
2. Using a knife or similar sharp edge, remove the surface layer from one side of the piece of rubidium. Observe the fresh surface.
3. Fill the beaker about half full with water.
4. Cut a small piece (about 2 mm cubed) of the element. Remove any oil with paper towel and drop the piece of the element into the beaker. Immediately cover the beaker with the wire gauze. Record observations of any chemical reaction.

### 8.(a) Experimental Design

Small pieces of various element solids from the same group in the periodic table are melted, and the melting-point range of the substance is observed and recorded. The process can be repeated for various element solids from other groups as well, to further investigate the periodic trends of melting points of elements.

## Materials

watch glass, 150-mm test tube, melting-point tube, ruler, rubber band, thermometer, 250-mL beaker, split one-hole stopper for thermometer, buret clamp, iron ring, wire gauze, ring stand, burner, wire loop stirrer, stirring rod, glass tube (1-m length), various samples of element solids from the same group

## Procedure

1. Place a small piece of an element solid on a clean watch glass and pulverize it if possible, by rubbing it with the bottom of a clean test tube. Avoid rubbing so hard that you break the test tube or the watch glass.
2. Carefully force a portion of the solid into a melting-point tube by pushing the open end of the tube into the powdered solid.
3. Allow the melting-point tube to drop (bottom first) through a 1-m length of glass tubing positioned vertically on the floor or laboratory bench. This knocks the solid into the sealed end of the melting-point tube. Repeat this process until the tube contains solid to a depth of 5 mm. Do not add too much element to the tube at any one time.
4. Attach the melting-point tube to the thermometer bulb of a thermometer with a rubber band, and support the thermometer in a 250-mL beaker about two-thirds full of water. The open ends of the melting-point tube must be above the water level.
5. Heat the water very slowly by adjusting the burner flame to a low level and removing the burner under the beaker periodically. Stir the water constantly. Try to obtain a rate of temperature increase of not more than  $3^{\circ}\text{C}/\text{min}$ .
6. Observe the solid in the melting-point tube carefully. Note and record the temperature at which the first crystal begins to melt and the temperature at which the last crystal disappears.
7. Allow the water to cool. Note and record the temperature at which the first crystal forms and the temperature at which all of the element has become solid.
8. Repeat the procedure for each of the other element solids.

## Reflecting

9. Elements that are used for structural support of buildings, roads, vehicles, furniture, and so on should be unreactive — for example, metals that are resistant to the “rusting” reaction. Containers should be made of unreactive materials, especially in a chemistry lab or a kitchen.

A match that does not “light” is useless, as is a fuel that will not burn, or a food that will not “cook”.

## SECTION 1.5 QUESTIONS

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### Understanding Concepts

1. The valence electrons are the ones that usually participate in chemical bonding. The number of valence electrons possessed by an atom determines the number of other atoms with which that atom can combine. The most chemically reactive metals are found in Groups 1, 2, and, to a lesser extent, 13 of the periodic table. These metals have low ionization energies. The most active nonmetals are found in Groups 16 and 17. These nonmetals have relatively large electron affinities.
2. (a) Ionization energy is the amount of energy required to remove an electron from an atom or ion in the gaseous state. Electron affinity is the energy change that occurs when an electron is accepted by an atom in the gaseous state. Both represent energy changes related to either the removal or the gaining of an electron.  
(b) Electronegativity describes the relative ability of an atom to attract electrons, whereas electron affinity is the energy change that occurs when an electron is accepted by an atom in the gaseous state. The electron affinity of

an atom affects the ability of an atom to attract electrons and is one of the components considered in the determination of an atom's electronegativity. Both terms are related to the gaining of an electron by an atom.

3. (a) Order of increasing ionization energy: students might predict K, Sr, Mg, Al, S (see below)  
Order of increasing atomic radius: S, Al, Mg, Sr, K  
Order of increasing electron affinity: K, Sr, Mg, Al, S  
Order of increasing electronegativity: K, Sr, Mg, Al, S
- (b) Ionization energy: Elements that have an electron in the valence shell farthest from the nucleus, where it is most shielded from the attractive forces of the positive nuclear charge, have lower ionization energies. (Ionization energies generally increase as you move left to right across the periodic table. The exception in this case is Al. Although it is to the right of Mg it has a lower ionization energy.)

Atomic radius: The atomic radii decrease from left to right across each period. As you move left to right across a period there is no increase in the electron shielding effect. Yet as a result of an increasing nuclear charge, there is an increase in nuclear attraction, which pulls the electrons closer to the nucleus, thus decreasing the size of the atom. The atomic radii increase from top to bottom in a given group because of the increasing number of energy levels that are filled with electrons. This has the effect of increasing the shielding effect, decreasing nuclear attraction, and increasing the size of the atom.

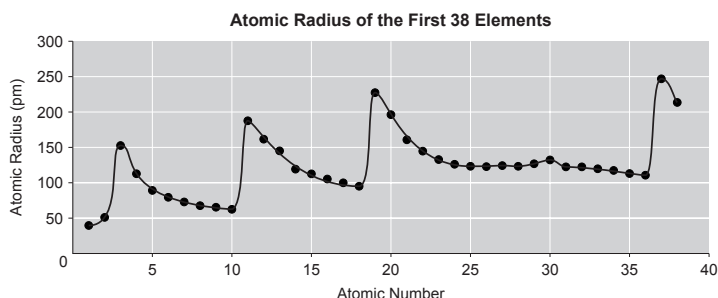
Electron affinity: When an electron is added to an atom of smaller radius, there is a greater attractive force between the nucleus of the atom and the new electron. Electron affinity and atomic radius are therefore inversely related. Smaller atoms should generally have higher electron affinities. Since the size of atoms generally decreases across a period, electron affinity will increase as we move from left to right within a period. And since the size of atoms generally decreases as we move up a group, electron affinity will also increase as we move from the bottom to the top of a group.

Electronegativity: An atom of smaller radius has a greater attractive force between the nucleus of the atom and a new electron. Electronegativity and atomic radius are therefore inversely related. Smaller atoms should generally have higher electronegativities. Since the size of atoms generally decreases across a period, electronegativity will increase as we move from left to right within a period. And since the size of atoms generally decreases as we move up a group, electronegativity will also increase as we move from the bottom to the top of a group.

4. First ionization energies for two isotopes of the same element would be expected to be the same. It is the protons and electrons in atoms that are responsible for determining the element's chemical behaviour. Consequently, isotopes of the same element share chemical properties despite their slight difference in mass.
5. (a)  $\text{Na}^+$ ,  $\text{Mg}^{2+}$ ,  $\text{Al}^{3+}$ . The valence electrons are the ones that usually participate in chemical bonding. Na, Mg, and Al are metals that have relatively low ionization energies and lose their valence electrons during chemical bonding to become positive charged ions.
- (b) From larger ionic radius to smaller ionic radius:  $\text{Na}^+$ ,  $\text{Mg}^{2+}$ ,  $\text{Al}^{3+}$ . Each ion contains the same number of electrons; however, for each successive positive ion, the nuclear charge is greater by 1+. Thus, each electron in an aluminum ion experiences greater attraction than an electron in a magnesium ion, and each electron in a magnesium ion experiences greater attraction than each electron in a sodium ion. The attraction of the nucleus for the electrons in positive ions increases across the period, and the ionic radius gets smaller.
6. (a) 2 valence electrons  
(b) Group 2  
(c) It should exhibit the properties of an alkaline earth metal.  
(d) The first two electrons are valence electrons in the outer shell, and so ionization energies for these electrons is low. The third electron is in a lower shell, and so the ionization energy is very high.
7. As the atomic number increases, atomic radii decrease while ionization energies increase as you move across a period from left to right on the periodic table. Atomic radii increase while ionization energies decrease as you move down a group from top to bottom on the periodic table. Graph shapes for atomic radii as compared to ionization energy would be reversed (slope and direction of graph line), similar to a mirror image.

## Applying Inquiry Skills

8. (a)



High Points: Li, Na, K, Rb

Low Points: H, He, Ar, Kr

- (b) Group 1 corresponds to the elements indicated by the peaks on the graph (excluding hydrogen). Group 18 corresponds to the elements indicated by the lowest points on the graph.
- (c) The atomic radii decrease, as the atomic number increases across a period. There are exceptions — on this graph, elements that fall between atomic number 26 and atomic number 32 fluctuate slightly from this decreasing pattern.
- (d) By taking the individual differences in atomic radius between all of the elements in the same group that precede cesium and barium, and then calculating an averaged difference, and then adding this averaged difference to the value of the element that directly precedes cesium and barium, the predicted atomic radii would be:

cesium 280 pm

barium 250 pm

These predicted values for atomic radii are higher than the actual atomic radii indicated on the periodic table, as shown below:

cesium 265 pm (actual atomic radius)

barium 217 pm (actual atomic radius)

However, lower values for cesium and barium would most likely be predicted if the trend of progressively smaller changes in atomic radii in the two elements directly preceding both cesium and barium was noted, and weighed into the calculations.

## CHAPTER 1 SUMMARY

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### Make a Summary

The student is to summarize the concepts learned so far, by creating a concept map that begins with the word “element” and uses as many words as possible from the list of key words. Refer to page 61 of the text for the list of key words.

### Reflect on your Learning

(Page 61)

By the end of the chapter the student should have developed a more in-depth understanding of the constituents of the atom, how these constituents are arranged, how atoms of various elements differ from each other, how the elements in the periodic table are organized, how there are patterns in properties among the elements, and how these patterns can be explained using the model of the atom. The student should also have developed skill in locating and interpreting information on a periodic table.