SECTION 1.3

Explaining Periodic Trends

In the previous section, you learned that the atomic number increases going from left to right across the periodic table. You also learned that the electron configuration of the valence shell is similar for atoms of the elements within a group. These are only two of the numerous trends in the periodic table. In this section, you will learn about other trends, which vary across a period or down a group in the periodic table, including trends in

- atomic radius
- ionization energy
- electron affinity
- electronegativity

Key Terms

atomic radius
effective nuclear charge
ionization energy
electron affinity
electronegativity

Atomic Radius

Consider the following questions: How would you predict that the size of an atom would change across a period or down a group? On what properties of elements would you base your prediction?

To answer these questions, you need a way to describe the size of an atom. As explained in Section 1.1, electrons move around a nucleus in what is best described as a cloud. In other words, an atom has no clearly defined boundary. To circumvent this problem, chemists define the size of an atom as the space in which the electrons spend 90 percent of their time, as shown in **Figure 1.16**. The size of an atom is usually reported in terms of its **atomic radius**, the distance from the centre of an atom to the boundary within which the electrons spend 90 percent of their time.

The next challenge becomes measuring the size of atoms. There is no way to directly measure the size of the space in which electrons spend 90 percent of their time, so chemists have had to find other methods for measuring atoms. The method that chemists use depends on the type of atom. For elements that can be crystallized in their pure form, chemists use a technique called *X-ray crystallography*. Using X-ray crystallography, chemists can measure the distance between the centres of atoms, as shown in **Figure 1.17 (A)**. The radius of the atom is half the distance between the centres of the atoms. Chemists use similar techniques, called *neutron diffraction* and *electron diffraction*, to measure the distance between the centres of atoms in of the diatomic gas molecules—oxygen, nitrogen, hydrogen, and the halogens. Half of this distance gives the atomic radius of an atom in a diatomic molecule. The method for measuring the radius of a hydrogen atom is illustrated in **Figure 1.17 (B)**.

atomic radius the distance from the centre of an atom to the boundary within which the electrons spend 90 percent of their time

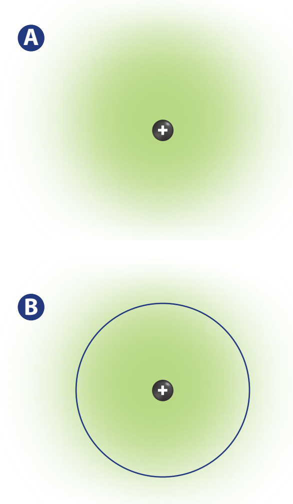


Figure 1.16 (A) This illustration represents the electron cloud of an atom. **(B)** The circle contains the region where the electrons spend 90 percent of their time.

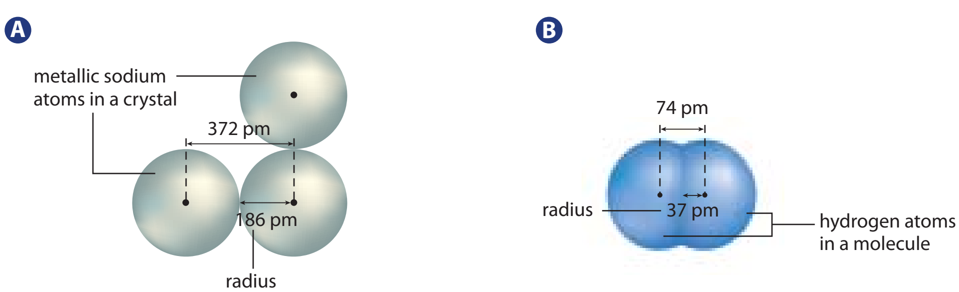


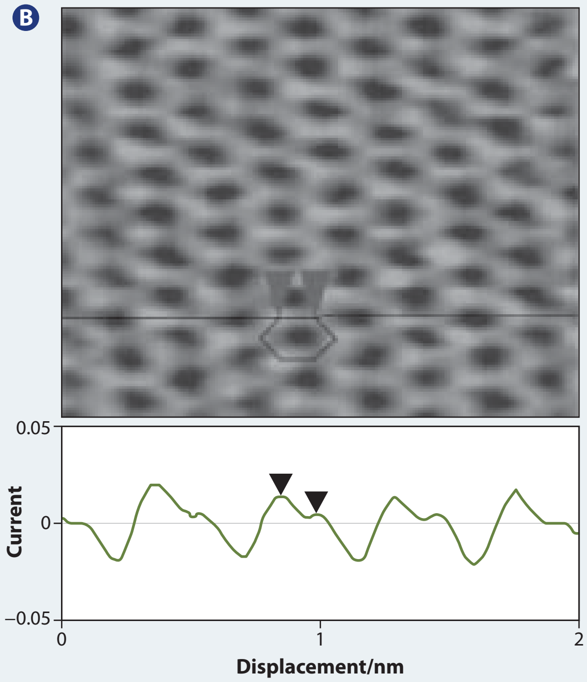
Figure 1.17 (A) These atoms represent sodium atoms in a crystal. **(B)** This molecule represents hydrogen. To obtain the atomic radius of an atom, the centre-to-centre distance is measured and then divided by 2. The unit used in these diagrams is the picometre (pm) which is equal to 1×10^{-12} m.

Activity 1.2 Estimating Atomic Radii

The development of the scanning tunnelling microscope (STM) has given chemists another method for measuring the size of an atom. In this activity, you will use STM data to estimate the radius of a carbon atom in graphite. Graphite consists of sheets of carbon atoms that are in hexagonal arrays, as shown in the diagram labelled A.

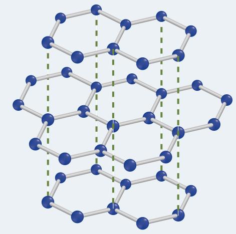
The picture labelled B is a STM image of graphite. The graph below the image shows the electric current through the STM while it was scanning across the black line on the image. The two pointers on the graph are the same as those in the image. They point to the centres of two carbon atoms in one hexagonal unit.

B



Data obtained from: Chaun-Jian Zhong et al. 2003. Atomic scale imaging: a hands-on scanning probe microscopy laboratory for undergraduates. *Journal of Chemical Education* 80: 194–197.

A



Materials

- ruler
- calculator

Procedure

1. Use the ruler to measure, in centimetres, the horizontal distance between the pointers on the graph. Estimate the measurement to two decimal places.
2. Measure, in centimetres, the distance that represents 1 nm (1×10^{-9} m) on the scale.
3. Calculate the centre-to-centre distance between the two adjacent carbon atoms by using the following formula:

   $$\text{actual distance} = \frac{\text{distance between pointers on graph}}{\text{distance representing 1 nm of scale}}$$

4. From the centre-to-centre distance you calculated in step 3, calculate the radius of a carbon atom.
5. The accepted radius of a carbon atom is 77×10^{-12} m. Calculate the percent error for the radius you calculated in step 4. Go to Measurement in Appendix A for help with calculating percent error.

Questions

1. Compare the structure of graphite to the STM image. What do the bright spots on the STM image represent? What do the dark spots represent?
2. What was your percent error? Give possible reasons for this error.
3. Using the same image and scale, propose another method for calculating the radius of a carbon atom that might be more accurate.

Analyzing Atomic Radii Data

Using a variety of techniques, chemists have been able to measure the atomic radii of atoms of all of the elements with stable isotopes. The results of these measurements are shown in **Figure 1.18** in the form of a periodic table. The values for atomic radii are given in picometres. The spheres in the table show the relative sizes of atoms of the elements.

To analyze the data for size versus atomic number of the elements, recall that all of the positive charge and nearly all of the mass are located in an extremely small volume at the centre of the atom. The electrons, with their negative charge, account for almost the entire volume of the atom. Both the positive charge in the nucleus and the negative charge of the electrons increase as the atomic number increases. The number of occupied shells is the same across any given period. Positive and negative charges attract one another. As the charge becomes greater, the attractive force becomes greater. All of this information helps to explain the pattern shown in **Figure 1.18**.

Trends in Atomic Radius within a Period

As you can see in **Figure 1.18**, the size of an atom decreases going from left to right across a period. In contrast, the number of positive charges in the nucleus increases. The number of electrons also increases, so you might expect the sizes of the atoms to get larger. However, all of the electrons in atoms of elements in the same period are in the same shell. With the increase in the number of positive charges, the attractive force on each electron becomes stronger. As the attractive force becomes stronger, the electrons are drawn closer to the nucleus, and thus the atom is smaller.

Go to **scienceontario** to find out more

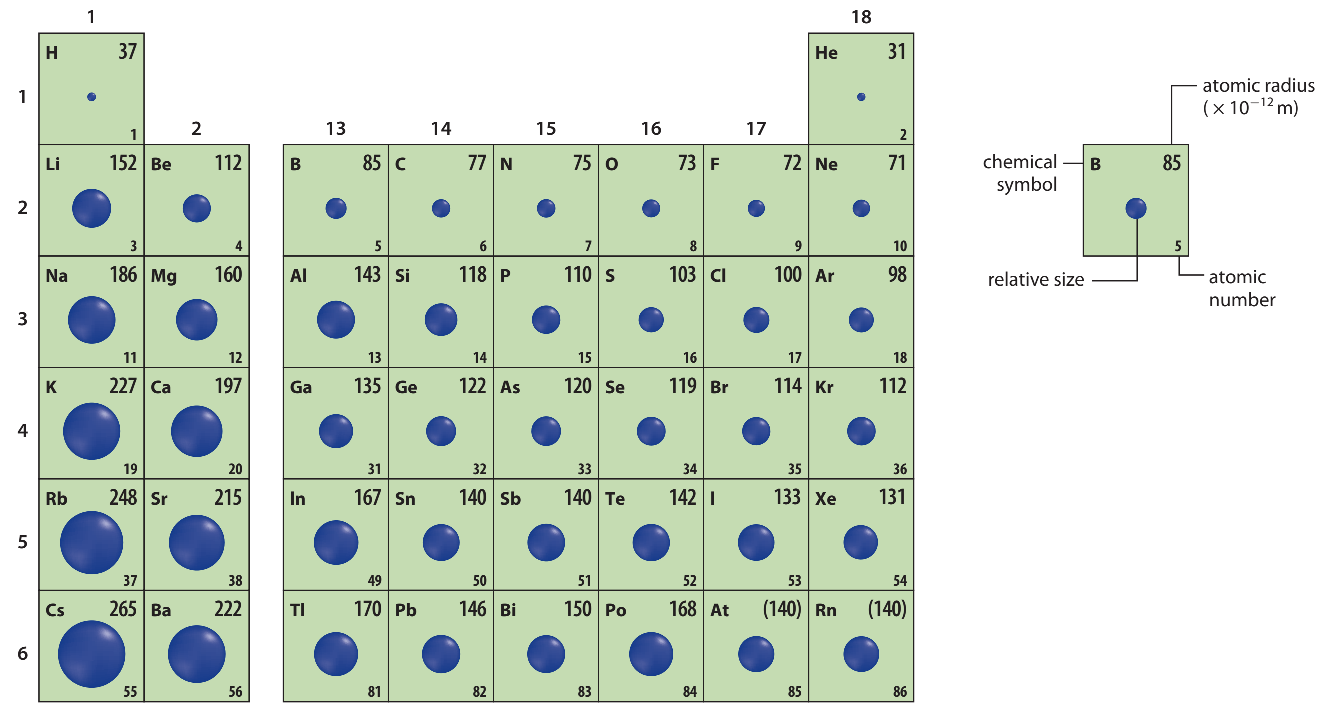


Figure 1.18 This partial periodic table shows the atomic radii of atoms of most of the main-group elements.
***Analyze** this periodic table with atomic radii and **describe** any trends that you see.*

Activity 1.3 Plotting Atomic Radius versus Atomic Number

Data in graphical form are often easier to analyze than data in table form. In this activity, you will plot a graph using data for atomic radius and atomic number.

Materials

- graph paper
- ruler
- pencil

Procedure

1. Plot a graph of atomic radius versus atomic number using the data in the periodic table in **Figure 1.18**. Atomic number will be your independent variable (*x*-axis), and atomic radius will be your dependent variable (*y*-axis). Decide on a scale for your *x*-axis. Notice that elements with atomic numbers 21 through 30, 39 through 48, and 57 through 80 are missing. Decide how you will show that they are missing on your graph.
2. Analyze the atomic radius data in **Figure 1.18** to decide on a suitable scale for your *y*-axis.
3. Plot the data. Connect the data points with straight lines.
4. Label each peak and valley on your graph with the chemical symbol for the element that matches the atomic number on the scale below.

Questions

1. What feature stands out the most on your graph? What does it tell you about the atomic radii of the elements in a particular period or group?
2. How does the atomic radius change going across a period?
3. How does the atomic radius change going down a group?
4. What characteristic of atomic radii is made more obvious by plotting the data in a graph rather than observing the data in the partial periodic table in **Figure 1.18**?
5. What characteristic is more obvious in the periodic table than in your graph?

effective nuclear charge the apparent nuclear charge, as experienced by the outermost electrons of an atom, as a result of the shielding by the inner-shell electrons

Trends in Atomic Radius within a Group

As you go down a group, the atomic number becomes greater, and thus the amount of positive charge also becomes greater. With each step down a group, however, the number of occupied electron shells increases. The inner shells that are filled shield the outer electrons from the positive charge of the nucleus. Because of this shielding, the **effective nuclear charge** appears smaller than the actual charge in the nucleus. As a result, the outer electrons are not attracted to the nucleus as strongly as they would have been without the shielding. Therefore, each additional electron shell, and thus the atomic radius, becomes larger going down a group.

Learning Check

13. How do chemists define the radius of an atom?

14. Why is it not possible to measure the size of an atom directly?

15. How does the amount of charge in the nucleus of an atom affect the size of the atom?

16. Explain the shielding effect of electrons, and describe how it affects the size of an atom.

17. List the following elements in order of increasing atomic number and then in order of increasing size: oxygen, tin, potassium, krypton. Explain why the orders of the elements are different in your two lists.

18. List at least three factors that affect atomic radius.

Ionization Energy

In chemical reactions, atoms can lose, gain, or share electrons with other atoms. When an atom loses an electron, the remaining ion is positively charged. A generalized equation is shown below. The A in the equation represents any atom.

$$\text{A(g)} + \text{energy} \rightarrow \text{A}^{+}\text{(g)} + \text{e}^{-} \text{ (first ionization)}$$

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

The amount of energy required to remove an electron from an atom is an important property that plays a role in determining the types of reactions in which atoms of that element might be involved. The amount of energy required to remove the outermost electron from an atom or ion in the gaseous state is called the **ionization energy**. The reason that the definition states that the atom and ion are in the gaseous state is to eliminate any effect of nearby atoms. If ionization energy were measured in the solid or liquid state, the adjacent atoms would affect the measurement.

After one electron has been removed, it is still possible to remove more electrons, leaving the ion with a larger positive charge. A generalized equation for the removal of a second electron is shown below.

$$\text{A}^{+}\text{(g)} + \text{energy} \rightarrow \text{A}^{2+}\text{(g)} + \text{e}^{-} \text{ (second ionization)}$$

The removal of any subsequent electron requires more energy than the removal of the previous electron because there are fewer negative charges repelling each subsequent electron but the same number of positive charges attracting it.

A graph of first ionization energy versus atomic number is shown in **Figure 1.19**. Notice that the elements at all of the peaks in the graph are noble gases. Atoms of the noble gases have filled valence shells. The filled shells make the atoms very stable so you would expect that it would take more energy to remove electrons from atoms of these elements. Notice, also, that the elements at the lowest points on the graph are Group 1 elements. Recall that all of their atoms have only one electron in their valence shell. When that electron is removed, the remaining ion has a filled outer shell, which is a stable configuration.

Suggested Investigation

ThoughtLab Investigation 1-B, Analyzing Ionization Energy Data

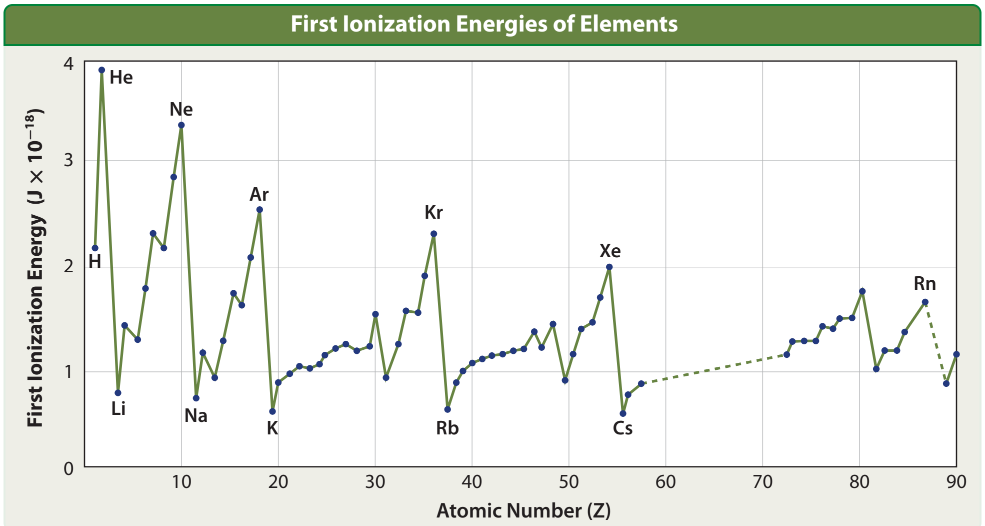


Figure 1.19 The data points on this graph represent the amount of energy required to remove one electron from a neutral atom of the various elements. The numbers on the scale must be multiplied by 10^{-18} to give the correct value in Joules. The dashed lines represent sections in which data points are missing.

Electron Affinity

If one atom loses an electron, another atom must gain this electron. Neutral atoms can gain electrons and become negatively charged. In the process of gaining an electron, a neutral atom can release energy or energy might be needed to add the electron. If energy is needed to add an electron to a neutral atom, the resulting negatively charged ion will be unstable and will soon lose the electron. If energy is released when an electron is added to a neutral atom, the resulting negatively charged ion will be stable. An equation for the generalized reaction in which a stable ion is formed is shown here.

$$A(g) + e^- \longrightarrow A^-(g) + \text{energy}$$

The energy that is either absorbed or released during the addition of an electron to a neutral atom is called the atom's **electron affinity**. **Figure 1.20** shows a partial periodic table of most of the main-group elements, with electron affinities of elements that become stable ions after gaining an electron. The electron affinities are expressed as negative values because energy is released in the formation of the ion. As the values become more negative, the ions become more stable. The elements with "unstable" in place of a numerical value are the ones that have a positive electron affinity.

electron affinity the energy absorbed or released when an electron is added to a neutral atom

1	2	13	14	15	16	17	18
1 H −1.21							2 He unstable
3 Li −0.989	4 Be unstable	5 B −0.448	6 C −2.02	7 N unstable	8 O −2.34	9 F −5.44	10 Ne unstable
11 Na −0.877	12 Mg unstable	13 Al −0.683	14 Si −2.22	15 P −1.19	16 S −3.32	17 Cl −5.78	18 Ar unstable
19 K −0.802	20 Ca −0.00393	31 Ga −0.689	32 Ge −1.97	33 As −1.30	34 Se −3.24	35 Br −5.39	36 Kr unstable
37 Rb −0.778	38 Sr −0.00769	49 In −0.481	50 Sn −1.78	51 Sb −1.68	52 Te −3.16	53 I −4.90	54 Xe unstable
55 Cs −0.756	56 Ba −0.232	81 Tl −0.320	82 Pb −0.583	83 Bi −1.51	84 Po −3.04	85 At −4.49	86 Rn unstable

Figure 1.20 The values in this table represent electron affinities. The values must be multiplied by 10^{-19} to get the correct values in Joules.

Trends in Electron Affinity

Analyze the electron affinities in **Figure 1.20**, and look for trends. In general, although not always, the electron affinities become increasingly negative going across a period and up a group. Also, notice that the noble gases do not form a stable ion if an electron is added. This is understandable when you consider that the outer shell of a noble gas is a filled shell. An added electron would be an unpaired electron in a higher shell, which would be very unstable. Next, consider the halogens. This group has the most negative electron affinities. Compared with all the other groups, the largest amount of energy is released when an electron is added, indicating that the ion is very stable. Remember that atoms of the halogens have seven electrons in their outer shell. The addition of one electron fills this shell. As a result, the ion is quite stable.

Electronegativity

Consider what happens when an atom of one element has lost an electron and an atom of another element has gained this electron. The result is oppositely charged ions, which attract each other. **Figure 1.21 (A)** shows a negatively charged bromide ion beside a positively charged potassium ion. The arrows represent the forces acting on the electrons. Notice that the electron lost by the potassium atom and gained by the bromine atom is attracted by both positive nuclei. Therefore, the overall effect is similar to shared electrons, although they are not truly shared. Recall that atoms of some elements do share electrons equally. **Figure 1.21 (B)** represents two chlorine atoms sharing two electrons. The arrows indicate that both electrons are attracted by both positively charged nuclei.

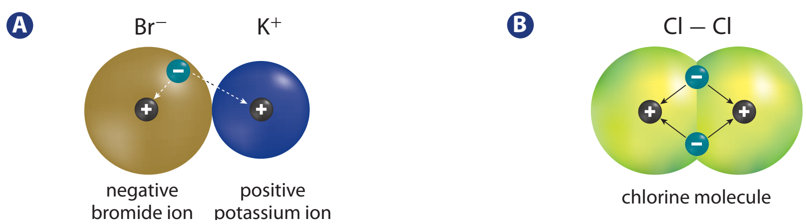


Figure 1.21 (A) When oppositely charged ions attract each other, their positively charged nuclei attract each other's electrons. **(B)** The nuclei of both atoms in the molecule attract the shared electrons.

electronegativity an indicator of the relative ability of an atom to attract shared electrons

The interactions in **Figure 1.21** lead to another important property of atoms, called electronegativity. An element's **electronegativity** is an indicator of the relative ability of an atom of this element to attract shared electrons. Because electronegativity is a relative value, it typically has no units. **Figure 1.22** shows a periodic table with electronegativities. Notice that Group 18 elements have no values. The noble gases do not share electrons with other atoms, so they do not have values for electronegativity.

Figure 1.22 Electronegativity is relative, so units are often not used. However, some chemists use the unit Paulings in honour of Linus Pauling, a Nobel prize-winning chemist.

***Identify** the element with the greatest electronegativity and the element with the smallest electronegativity. **Describe** their relationship with each other in the periodic table.*

Electronegativity Values

electronegativity < 1.0
1.0 ≤ electronegativity < 2.0
2.0 ≤ electronegativity < 3.0
3.0 ≤ electronegativity < 4.0

1 **H** 2.2																	2 **He**
3 **Li** 1.0	4 **Be** 1.6											5 **B** 2.0	6 **C** 2.6	7 **N** 3.0	8 **O** 3.4	9 **F** 4.0	10 **Ne**
11 **Na** 0.9	12 **Mg** 1.3											13 **Al** 1.6	14 **Si** 1.9	15 **P** 2.2	16 **S** 2.6	17 **Cl** 3.2	18 **Ar**
19 **K** 0.8	20 **Ca** 1.0	21 **Sc** 1.4	22 **Ti** 1.5	23 **V** 1.6	24 **Cr** 1.7	25 **Mn** 1.6	26 **Fe** 1.8	27 **Co** 1.9	28 **Ni** 1.9	29 **Cu** 1.9	30 **Zn** 1.7	31 **Ga** 1.8	32 **Ge** 2.0	33 **As** 2.2	34 **Se** 2.6	35 **Br** 3.0	36 **Kr**
37 **Rb** 0.8	38 **Sr** 1.0	39 **Y** 1.2	40 **Zr** 1.3	41 **Nb** 1.6	42 **Mo** 2.2	43 **Tc** 2.1	44 **Ru** 2.2	45 **Rh** 2.3	46 **Pd** 2.2	47 **Ag** 1.9	48 **Cd** 1.7	49 **In** 1.8	50 **Sn** 2.0	51 **Sb** 2.1	52 **Te** 2.1	53 **I** 2.7	54 **Xe**
55 **Cs** 0.8	56 **Ba** 0.9	57 **La** 1.1	72 **Hf** 1.3	73 **Ta** 1.5	74 **W** 1.7	75 **Re** 1.9	76 **Os** 2.2	77 **Ir** 2.2	78 **Pt** 2.2	79 **Au** 2.4	80 **Hg** 1.9	81 **Tl** 1.8	82 **Pb** 1.8	83 **Bi** 1.9	84 **Po** 2.0	85 **At** 2.2	86 **Rn**
87 **Fr** 0.7	88 **Ra** 0.9	89 **Ac** 1.1	104 **Rf**	105 **Db**	106 **Sg**	107 **Bh**	108 **Hs**	109 **Mt**	110 **Ds**	111 **Rg**	112 **Cn**	113 **Uut**	114 **Uuq**	115 **Uup**	116 **Uuh**		118 **Uuo**

Analyzing Trends in Electronegativity

Analyzing **Figure 1.22** reveals that the electronegativity of the elements increases going up a group and going from left to right across a period. Thus, fluorine has the largest electronegativity, and francium has the smallest. To discover one of the reasons for these trends, examine **Figure 1.23**. It shows a silicon atom beside a cesium atom, and a silicon atom beside a fluorine atom.

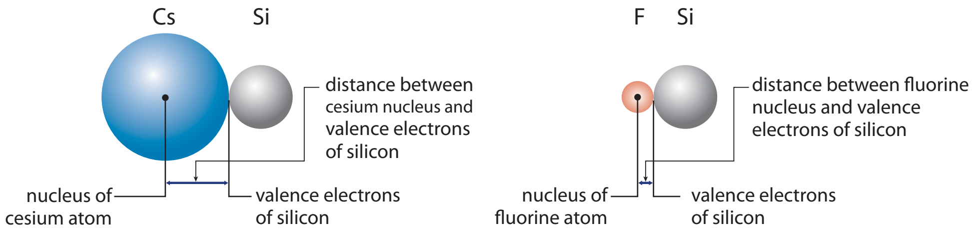


Figure 1.23 As you can see, the nucleus of a small atom can get much closer to the outer-shell electrons of another atom than the nucleus of a larger atom can.

***Predict** how the distance of the nucleus of an atom from the electrons of an adjacent atom will affect the strength of the attraction between that nucleus and the outer electrons of the adjacent atom.*

The elements in **Figure 1.23** were chosen because silicon is a medium-sized atom, cesium is very large, and fluorine is very small. Notice how far the nucleus of the cesium atom is from the outer electrons of the silicon atom and how close the nucleus of the fluorine atom is to the outer electrons of the silicon atom. Because the positively charged nucleus of a small atom can get much closer to the electrons of another atom, this nucleus should be able to exert a stronger attractive force on those electrons. This point raises a new question: Is atomic radius a critical factor in determining the electronegativity of atoms? Examine **Figure 1.24** to find out. The heights of the columns show the electronegativities of the elements, while the spheres on top of the columns show the relative sizes of the atoms. The correlation between size and electronegativity appears to be strong. As the atom becomes smaller, its electronegativity becomes larger.

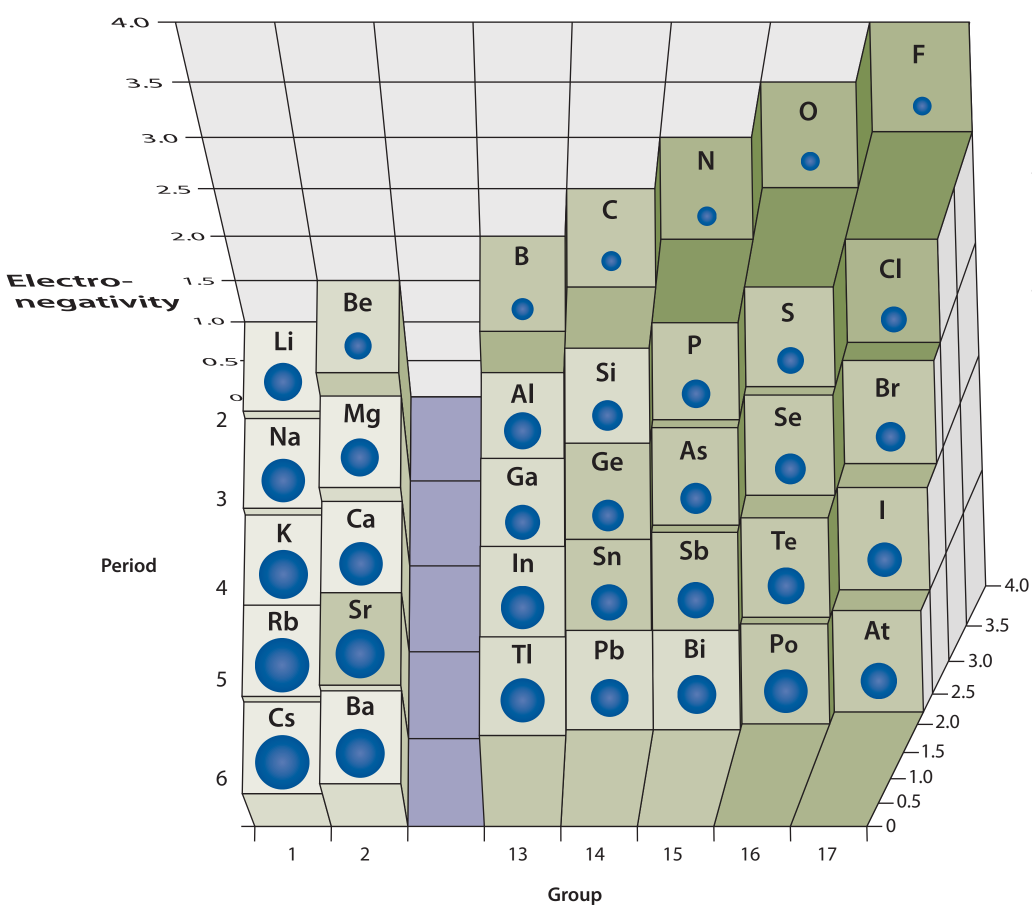


Figure 1.24 This periodic table is shown in three dimensions, so you can easily compare the sizes and electronegativities of the elements.

Activity 1.4 Analyzing Trends in the Halogens

The halogens are a group of reactive non-metals. Although the atoms all have seven electrons in their valence shell, their properties vary somewhat. The following table lists values for several properties of the halogens.

Properties of Halogens

Halogen	Atomic Radius ($\times 10^{-12}$ m)	Ionization Energy ($\times 10^{-18}$ J)	Electron Affinity ($\times 10^{-19}$ J)	Electro-negativity	Reactivity
fluorine	72	2.79	−5.44	3.98	Highest
chlorine	100	2.08	−5.78	3.16	↓
bromine	114	1.89	−5.39	2.96	↓
iodine	133	1.67	−4.90	2.66	↓
astatine	140	1.53	−4.49	2.2	Lowest

Procedure

1. Before plotting points for graphs, predict what they will look like, by sketching graphs of (A) atomic radius, (B) ionization energy, (C) electron affinity, and (D) electronegativity versus atomic number, for the halogens.
2. On four separate graphs, plot the data for atomic radius, ionization energy, electron affinity, and electronegativity against atomic number.
3. Describe any trends that you see.

Questions

1. Compare your sketches with the graphs plotted from the data. Describe any differences.
2. Suggest a possible explanation for each trend that you listed in step 3, based on what you know about atoms.
3. Suggest the property (or properties) that might be most responsible for the trend in reactivity of the halogens. Explain your reasoning.

Summarizing Trends in the Periodic Table

Figure 1.25 summarizes the trends in atomic radius, ionization energy, electron affinity, and electronegativity within periods and groups. The arrows indicate the direction of increasing values. These concepts are all factors in how atoms of different elements react, or do not react, with one another. In the next chapter, you will learn about bond formation and the nature of the bonds that form when chemical reactions occur.

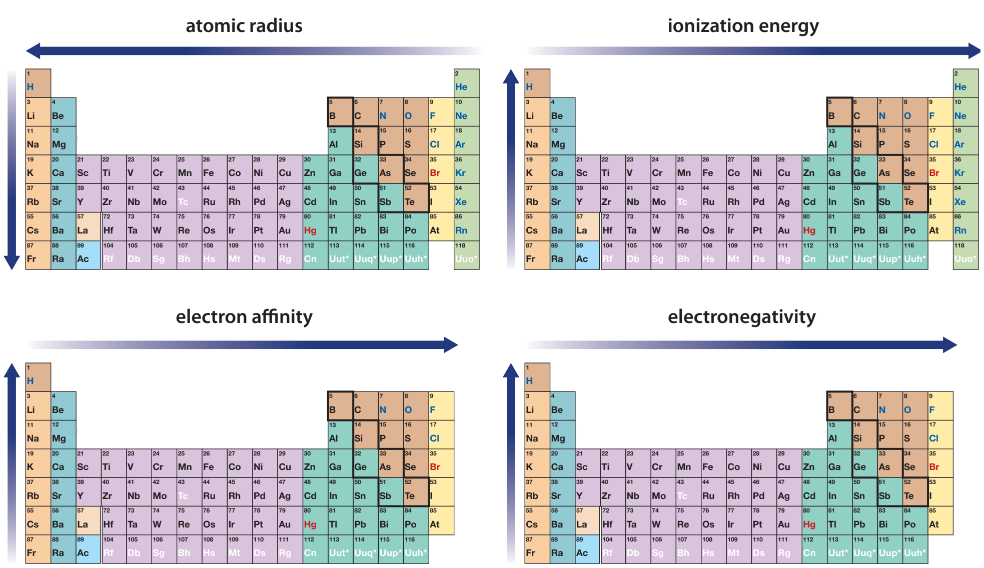


Figure 1.25 You can use these miniature periodic tables as a quick reference when you need to remember the trends in the groups and periods of elements. The two bottom tables do not include Group 18 because the noble gases do not have values for these properties.

***Identify** the three properties that follow the same trends. **State** which property follows trends opposite to the other three.*

STSE

CHEMISTRY Connections

Elements of the Body

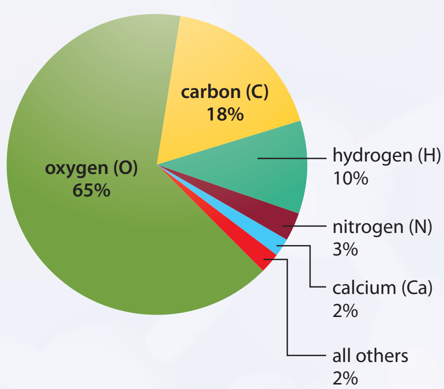
Every time you eat a sandwich or take a breath of air, you are taking in elements that your body needs to function normally. These elements have specific properties, depending on their location in the periodic table. The circle graph below shows the percent by mass of elements in cells in the human body.

OXYGEN In an adult body, there are more than 14 billion billion billion oxygen atoms! Without a constant input of oxygen into the blood, the human body could die in just a few minutes.

CARBON Carbon can form strong bonds with itself and other elements. Carbon forms the long-chained carbon backbones that are an essential part of biological molecules such as carbohydrates, proteins, and lipids. The DNA molecule that determines your characteristics relies on the versatility of carbon and its ability to bond with many different elements.

HYDROGEN Although there are more hydrogen atoms in the human body than there are atoms of all the other elements combined, hydrogen represents only 10 percent of the composition by mass because of its significantly lower mass. Your body requires hydrogen in a variety of essential compounds, like water, rather than in its elemental form. With oxygen and carbon, hydrogen is also a crucial component of carbohydrates and other biological molecules that your body needs for energy.

Percent by Mass of the Elements in the Human Body



The human body is composed of many different elements.

The entire human body is covered with muscle tissue.

NITROGEN As shown in the illustration above, the human body is entirely covered with muscle tissue. Nitrogen atoms are found in the compounds that make up the muscle protein.

OTHER ELEMENTS IN THE BODY Oxygen, carbon, hydrogen, and nitrogen are the most abundant elements in your body and yet are but a few of the elements that your body needs to live and grow. Trace elements, which together make up less than 2 percent of the body's mass, are a critical part of your body. Your bones and teeth could not grow without the constant intake of calcium. Although sulfur comprises less than 1 percent of the human body by mass, it is an essential component in some of the proteins, such as those in your fingernails. Sodium and potassium are crucial for the transmission of electrical signals in your brain.

Connect to Society

Many food manufacturers add sulfites to their products. What are the benefits of adding sulfites to food? How much of the sulfites are needed to provide these benefits? What are the risks of adding sulfites to food? Do sulfites occur naturally in any foods? Carry out research to answer these questions. Then state why you do or do not think that sulfites should be used as a food additive.

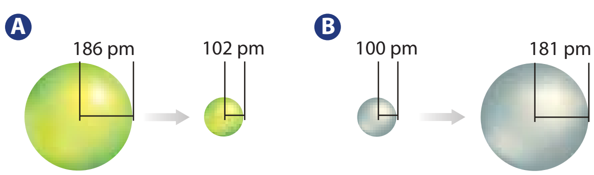
Section 1.3 REVIEW

Section Summary

- The atomic radius of the atoms of an element is influenced by the amount of charge in the atom's nucleus and by the number of occupied electron shells. The atomic radius increases when going down a group and decreases when going across a period from left to right.
- The ionization energy of the atoms of an element is influenced by the distance between its outermost electron and its nucleus. The ionization energy decreases when going down a group and increases when going across a period.
- The electron affinity of the atoms of an element is influenced by whether the valence shell of the atoms is filled. The electron affinity decreases when going down a group and increases when going across a period.
- The electronegativity of the atoms of an element is influenced by the atomic radius. The electronegativity decreases when going down a group and increases when going across a period.

Review Questions

1. K/U What is atomic radius and how is this value obtained?
2. T/I List the following elements in order of increasing radius: Ba, Cs, O, Sb, Sn.
3. K/U State the trend in atomic radius within a group. Describe the factor that accounts for this trend.
4. T/I The following figures represent neutral atoms gaining or losing an electron to become ions. One of these figures represents a chlorine atom, Cl, gaining an electron to become a negatively charged ion, Cl^{-}. The other figure represents a sodium atom, Na, losing an electron to become a positively charged ion, Na^{+}. Identify which figure represents each element. Provide possible explanations for the change in size upon ionization. **Note:** The unit pm (picometre) is 10^{-12} m.



5. K/U Define ionization energy. Use a chemical equation to clarify your definition.
6. T/I Review the graph of First Ionization Energies in **Figure 1.19** on page 35. Notice that the differences in ionization energy for the first 20 elements is quite significant. Then, for elements 20 to 30, the values are very similar. Provide a possible explanation for the similarity in these values.
7. T/I Write a chemical equation for an atom losing a third electron. Would the ionization energy for this reaction be larger or smaller than the ionization energy for the loss of the first or second electron? Explain.
8. T/I Explain why helium does not have a third ionization energy.
9. T/I List the following elements in the order of increasing ionization energy for the first ionization: arsenic, cesium, fluorine, helium, phosphorus, strontium.
10. K/U What is the significance of an element having a positive electron affinity? Is the resulting ion stable or unstable? Explain why.
11. K/U Atoms of what group have the most negative electron affinities? Explain why.
12. C Using diagrams, show how electronegativity and electron affinity are different.
13. K/U Explain the relationship between electronegativity and the size of an atom.
14. T/I Copy the following table in your notebook. Fill in the data for atomic mass, using the periodic table in **Figure 1.12** on page 24. Fill in the data for atomic radius, using **Figure 1.18** on page 33. Compare the data for the two elements. Would a piece of aluminum or a piece of lead be denser? Explain your reasoning.

Data for Aluminum and Lead

Element	Atomic Mass	Atomic Radius
aluminum		
lead		

15. A Use the data in **Figure 1.19** and **Figure 1.20** on page 35 to answer the following questions. First ionization data can also be found in the data table in Investigation 1-B on page 42.
    a. What would the total change in energy be to remove an electron from a sodium atom and add the electron to a chlorine atom?
    b. Which of the following combinations would be more stable, a sodium atom and a chlorine atom, or a positively charged sodium ion and a negatively charged chloride ion? Explain your reasoning.