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Chemical Periodicity

Forsythe **Parsons**

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Chapter 1

Chemical Periodicity

1.1 The Periodic Table

Lesson Objectives

The student will:

- explain the periodic law.
- describe the differences among metals, nonmetals, and metalloids.
- draw a rough sketch of the modern periodic table and indicate the portion of the table occupied by metals, nonmetals, and metalloids.
- identify the stair-step line that separates the metallic elements from the nonmetallic ones.

Vocabulary

- ductile
- malleable
- periodic law

Introduction

In the periodic table, the elements are arranged according to similarities in their properties. The elements are listed in order of increasing atomic number as you read from left to right across a period. In this chapter, you will learn the general behavior and trends within the periodic table that result from this arrangement in order to predict the properties of the elements.

The Periodic Law

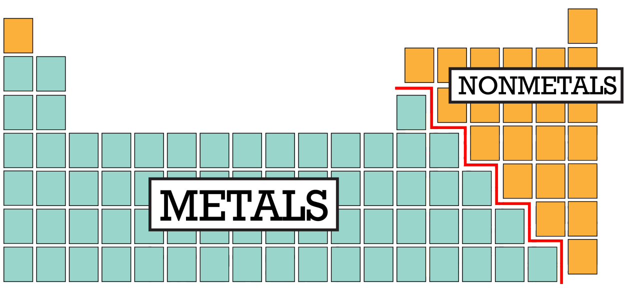
The periodic table is a powerful tool that provides a way for chemists to organize the chemical elements. The word "periodic" means happening or recurring at regular intervals. The **periodic law** states that the properties of the elements recur periodically as their atomic numbers increase. This is because the electron configurations of the atoms vary periodically with their atomic number. Since the physical and chemical properties of elements depend on their electron configurations, many of the physical and chemical properties of the elements also tend to repeat in a pattern.

You may recall that the periodic table was created by Russian scientist Dmitri Mendeleev. Mendeleev used similarities in properties to construct his periodic table of the elements. He was even able to predict the properties of several elements missing from the periodic table by using the properties of neighboring elements. Mendeleev arranged the elements in a table by increasing atomic weight, although he sometimes had to ignore the atomic weights in order to group elements with similar chemical behaviors together.

The work of Henry Moseley led to the arrangement of elements based on their properties and atomic numbers, *not* atomic masses, which cleared up the inconsistencies of Mendeleev's arrangement. Moseley's periodic table is now considered the current periodic table.

Metals, Nonmetals, and Metalloids

There is a progression from metals to nonmetals across each period of elements in the periodic table. The diagonal line at the right side of the table shown below separates the elements into two groups: the metals and the nonmetals. The elements that are on the left of this line tend to be metals, while those to the right tend to be nonmetals. The elements that are directly on the diagonal line are metalloids. Metallic character generally increases from top to bottom down a group and right to left across a period. The noticeable exception is hydrogen, which is grouped with the metals but is actually a nonmetal.



Most of the chemical elements are metals. Most metals have the common properties of being shiny, being very dense, and having high melting points. Metals tend to be **ductile** (can be drawn out into thin wires) and **malleable** (can be hammered into thin sheets). Metals are good conductors of heat and electricity. All metals are solids at room temperature except for mercury. In chemical reactions, metals easily lose electrons to form positive ions. Examples of metals are silver, gold, and zinc.

Nonmetals are generally brittle, dull, and have low melting points. They are generally poor conductors of heat and electricity. In chemical reactions, they tend to gain electrons to form negative ions. Examples of nonmetals are hydrogen, carbon, and nitrogen.

Metalloids have properties of both metals and nonmetals. Metalloids can be shiny or dull. Electricity and heat can travel through metalloids, although not as easily as they can through metals. They are also called semi-metals. They are typically semiconductors, which means that they conduct electricity better than insulators, but not as well as conductors. Semiconductors are valuable in the computer chip industry. Examples of metalloids are silicon and boron.

Lesson Summary

- The word "periodic" means happening or recurring at regular intervals.
- The periodic law states that the properties of the elements recur periodically as their atomic numbers increase.
- There is a progression from metals to nonmetals across each period of elements in the periodic table.
- Metallic character generally increases from top to bottom down a group and right to left across a period.

Further Reading / Supplemental Links

This video is on the periodic table.

- http://www.youtube.com/watch?v=1geccHiylcU&feature=fvw

Review Questions

1. Why is the table of elements called “the periodic table”?
   (a) It describes the periodic motion of celestial bodies.
   (b) It describes the periodic recurrence of chemical properties.
   (c) Because the rows are called periods.
   (d) Because the elements are grouped as metals, metalloids, and nonmetals.
   (e) None of these.
2. Which of the following elements is a nonmetal?
   (a) oxygen
   (b) lead
   (c) iron
   (d) zinc
   (e) All of these are metals.
3. Which of the following metals is *not* an element?
   (a) gold
   (b) silver
   (c) copper
   (d) bronze
   (e) All of these are elements.
4. Which of the following elements is a metalloid?
   (a) chlorine
   (b) magnesium
   (c) rhenium
   (d) boron
   (e) None of these.

1.2 Periodic Trends in Atomic Size

Lesson Objectives

The student will:

- define atomic radius.
- define the shielding effect.
- describe the factors that determine the trend in atomic size.
- describe the general trend in atomic size for groups and periods.
- use the general trends to predict the relative sizes of atoms.
- describe variations that occur in the general trend of atomic size in the transition metals.

Vocabulary

- atomic radius
- diatomic molecule
- nuclear charge
- shielding effect

Introduction

In the periodic table, there are a number of physical properties that are trend-like. This means that as you move down a group or across a period, you will see the properties changing in a general direction. The actual trends that are observed with atomic size have to do with three factors. These factors are:

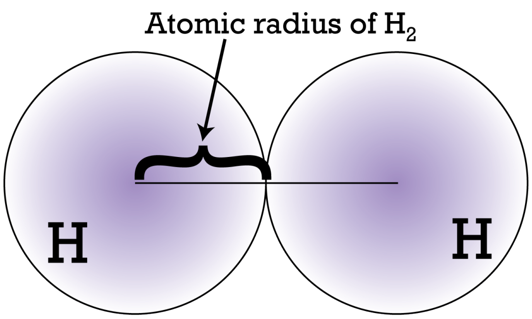
1. the number of protons in the nucleus (called the **nuclear charge**).
2. the number of energy levels holding electrons and the number of electrons in the outer energy level.
3. the number of electrons held between the nucleus and its outermost electrons (called the **shielding effect**).

Atomic Radius Defined

The gold foil experiment performed by Rutherford in 1911 (see the chapter "The Atomic Theory" for more details), was the first experiment that gave scientists an approximate measurement for the size of the atom. The atomic size is the distance from the nucleus to the valence shell, where the valence electrons are located. Using the technology available in the early part of the 1900s, Rutherford was able to determine quantitatively that the nucleus had a radius size smaller than 3×10^{-12} cm. The size of the atom is significantly larger, being approximately 2×10^{-8} cm in diameter.

The region in space occupied by the electron cloud of an atom is often thought of as a probability distribution of the electron positions. Consequently, there is no well-defined outer edge of the electron cloud. Because it is so difficult to measure atomic size from the nucleus to the outermost edge of the electron cloud, chemists use other approaches to get consistent measurements of atomic sizes.

Atomic size is defined in several different ways, which often produce some variations in the measurement of atomic sizes. One way that chemists define atomic size is by using the atomic radius. The **atomic radius** is one-half the distance between the centers of a homonuclear diatomic molecule, as illustrated below. A **diatomic molecule** is a molecule made of exactly two atoms, while homonuclear means both atoms are the same element.



Group Trends in Atomic Radii

Let's now look at how the atomic radii changes from the top of a family to the bottom. Take, for example, the Group 1A metals (see **Table** 1.1). Every atom in this family has the same number of electrons in the outer energy level (true for all main group families). Each period in the periodic table represents another added energy level. When we first learned about principal energy levels, we learned that each new energy level was larger than the one before. Therefore, as we move down the periodic table, each successive period represents the addition of a larger energy level, thus increasing the atomic radius.

Table 1.1: **Group 1A Data**

Element	Number of Protons	Electron Configuration
Lithium (Li)	3	[He]$2s^1$
Sodium (Na)	11	[Ne]$3s^1$
Potassium (K)	19	[Ar]$4s^1$
Rubidium (Rb)	37	[Kr]$5s^1$
Cesium (Cs)	55	[Xe]$6s^1$

One other contributing factor to atomic size is the shielding effect. The protons in the nucleus attract the valence electrons in the outer energy level, but the strength of this attraction depends on the size of the charges, the distance between the charges, and the number of electrons between the nucleus and the valence electrons. The presence of the core electrons weakens the attraction between the nucleus and the valence electrons. This weakening is called the shielding effect. Note that although valence electrons do participate in shielding, electrons in the same energy level do not shield each other as effectively as the core electrons do. As a result, the amount of shielding primarily depends on the number of electrons between the nucleus and the valence electrons. When the nucleus pulls strongly on the valence electrons, the valence shell can be pulled in tighter and closer to the nucleus. When the attraction is weakened by shielding, the valence shell cannot be pulled in as close. The more shielding that occurs, the further the valence shell can spread out.

For example, if you are looking at the element sodium, it has the electron configuration:

Na $1s^22s^22p^63s^1$

The outer energy level is $n = 3$. There is one valence electron, but the attraction between this lone valence electron and the nucleus, which has 11 protons, is shielded by the other 10 inner (or core) electrons.

When we compare an atom of sodium to one of cesium, we notice that the number of protons increases, as well as the number of energy levels occupied by electrons. The increase in the number of protons, however, is also accompanied by the same increase in the number of shielding electrons.

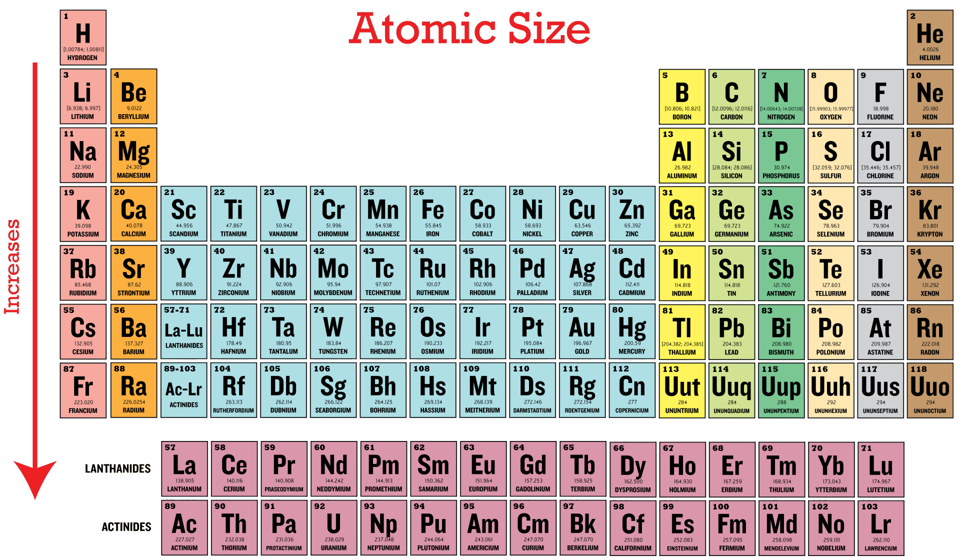
Cs $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^1$

The result is that the valence electron in both atoms feels a similar pull from the nucleus, but the valence electron in the cesium atom is further from the nucleus because it is in a higher energy level. Compared to the shielding effect, the increase in the number of energy levels has a greater impact on the atom's size. Consequently, the size of a cesium atom is larger than that of a sodium atom.

This is true for not only Group 1A metals, but for all of the groups across the periodic table. For any given group, as you move downward in the periodic table, the size of the atoms increases. For instance,

the largest atoms in the halogen family are bromine and iodine (astatine is radioactive and only exists for short periods of time, so we won't include it in the discussion). You can imagine that with the increase in the number of energy levels, the size of the atom must increase. The increase in the number of energy levels in the electron cloud takes up more space.

The periodic table below shows the trend of atomic size for groups, with the arrow indicating the direction of the increase.



Example:

Which of the following is larger? Explain.

1. As or Sb
2. Ca or Be
3. polonium or sulfur

Solution:

1. Sb, because it is below As in Group 15.
2. Ca, because it is below Be in Group 2.
3. Polonium, because it is below sulfur in Group 16.

Period Trends in Atomic Radii

In order to determine the trend for the periods, we need to look at the number of protons (nuclear charge), the number of energy levels, and the shielding effect. For a row in the periodic table, the atomic number still increases (as it did for the groups), and thus the number of protons would increase. For a given period, however, we find that the outermost energy level does not change as the number of electrons increases. In period 2, for example, each additional electron goes into the second energy level, so the total number of energy levels does not go up. **Table** 1.2 shows the electron configuration for the elements in period 2.

Table 1.2: **Electron Configurations for Elements in Period 2**

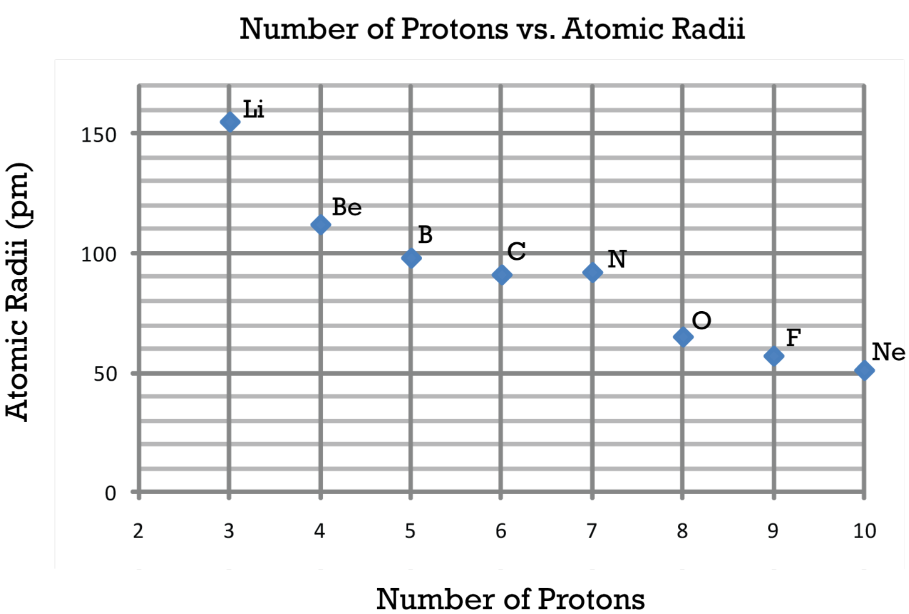
Element	Number of Protons	Electron Configuration
Lithium (Li)	3	$1s^2 2s^1$

Table 1.2: (continued)

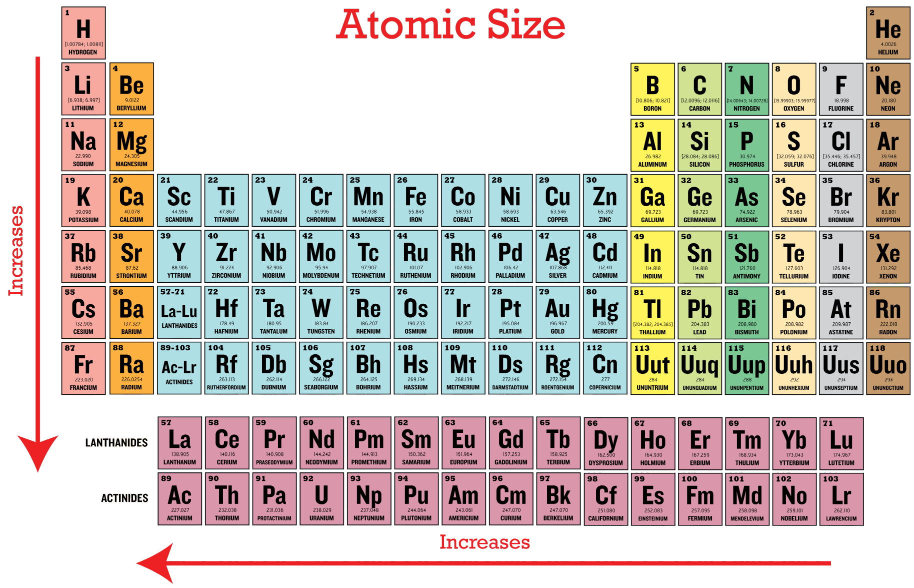
Element	Number of Protons	Electron Configuration
Beryllium (Be)	4	$1s^2 2s^2$
Boron (B)	5	$1s^2 2s^2 2p^1$
Carbon (C)	6	$1s^2 2s^2 2p^2$
Nitrogen (N)	7	$1s^2 2s^2 2p^3$
Oxygen (O)	8	$1s^2 2s^2 2p^4$
Fluorine (F)	9	$1s^2 2s^2 2p^5$

Looking at the elements in period 2, the number of protons increases from three (for lithium) to nine (for fluorine). Therefore, the nuclear charge increases across a period. Meanwhile, the number of energy levels occupied by the electrons remains the same. How will this affect the radius? We know that every one of the elements in this period has two core electrons in the inner energy level ($n = 1$). The core electrons shield the outer electrons from the charge of the nucleus. Since the number of protons attracting the outer electrons increases while the shielding remains the same, the valence electrons are pulled closer to the nucleus, making the atom smaller.

Consider the elements lithium, beryllium, and fluorine from period 2. With lithium, the two core electrons will shield the one valence electron from three protons. Beryllium has four protons being shielded by the $1s^2$ electrons. For fluorine, there are nine protons and nine electrons. All three of these elements have the same core electrons: the $1s^2$ electrons. As the number of protons increases, the nuclear charge increases. With an increase in nuclear charge, there is an increase in the pull between the protons and the outer level, pulling the outer electrons toward the nucleus. The amount of shielding from the nucleus does not increase because the number of core electrons remains the same. The net result is that the atomic size decreases going across the row. In the graph below, the values are shown for the atomic radii for period 2.



Let's add this new trend to the periodic table. In the diagram below, you will notice that the trend arrow for the period shows the atomic radii increase going from right to left, which is the equivalent to saying that the atomic radii decrease from left to right.



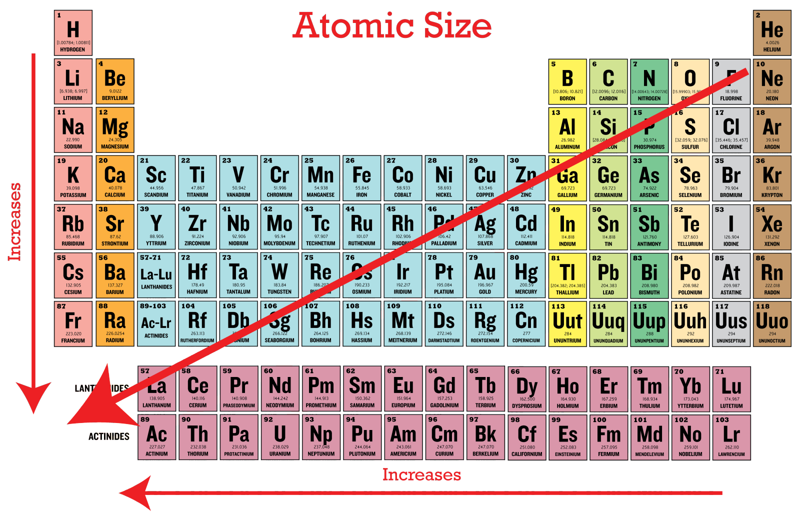
Considering these two trends, you will recognize that the largest atom, francium (atomic number 87), is at the bottom left-hand corner of the periodic table, while the smallest atom, helium (atomic number 2) is at the top right-hand corner of the table.

For an introduction to the electronic organization of the periodic table **(1c)**, see http://www.youtube.com/watch?v=35cWAxtHUGw (4:20).

Figure 1.1: (Watch Youtube Video)
http://www.ck12.org/flexbook/embed/view/372

Atomic Radii of Transition Elements

The general trend for atomic radii in the periodic table would look similar to that illustrated in the diagram below. The elements with the smallest atomic radii are found in the upper right; those with the largest atomic radii are found in the lower left.



Until now, we have worked solely with the main group elements. Let's consider how three factors affecting atomic size affect transition metals. The first row of the transition metals all contain electrons in the $3d$ sublevel and are referred to as the $3d$ metals. **Table** 1.3 shows the electron configuration for the ten elements in this row. The number of protons is increasing, so the nuclear charge is increasing.

Table 1.3: **Electron Configuration for 3*d* Metals**

Element	Number of Protons	Electron Configuration
Scandium (Sc)	21	$[Ar]3d^1 4s^2$
Titanium (Ti)	22	$[Ar]3d^2 4s^2$
Vanadium (V)	23	$[Ar]3d^3 4s^2$
Chromium (Cr)	24	$[Ar]3d^5 4s^1$
Manganese (Mn)	25	$[Ar]3d^5 4s^2$
Iron (Fe)	26	$[Ar]3d^6 4s^2$
Cobalt (Co)	27	$[Ar]3d^7 4s^2$
Nickel (Ni)	28	$[Ar]3d^8 4s^2$
Copper (Cu)	29	$[Ar]3d^{10} 4s^1$
Zinc (Zn)	30	$[Ar]3d^{10} 4s^2$

You may notice that some of these configurations are not what you would expect based on the information presented so far. Both chromium and copper have one of the $4s$ electrons moved to a $3d$ orbital. A simplified explanation for these unusual electron configurations is that the d sub-level is particularly stable when it is half-full (5 electrons) or completely full (10 electrons). Since the $4s$ and $3d$ orbitals are close in energy, this added stabilization is enough to change the location of one valence electron.

Table 1.4: **Atomic Radii for 3d Metals**

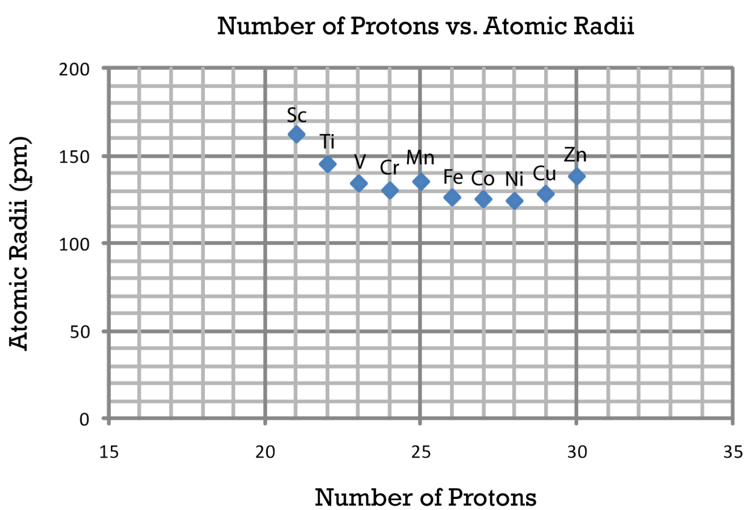
Element	Atomic Radii (pm)
Scandium (Sc)	164
Titanium (Ti)	147
Vanadium (V)	135
Chromium (Cr)	129
Manganese (Mn)	137
Iron (Fe)	126

Table 1.4: (continued)

Element	Atomic Radii (pm)
Cobalt (Co)	125
Nickel (Ni)	125
Copper (Cu)	128
Zinc (Zn)	137

Table 1.4 lists the atomic radii for the first row of the transition metals. It can be seen from this table that the period trend in atomic radii is not followed as closely by the transition metals. Since we are adding electrons to the 3d orbitals, we are actually adding to the core electrons and *not* to the valence orbitals. Although the nuclear charge is going up, the shielding is also increasing with each added electron. Because of this, there is less atomic contraction throughout the transition metals.

The graph of the number of protons versus the atomic radii for the 3d metals is shown below. Compared to the same graph for the elements in period 2, the graph for transition metals shows the trend for atomic radii is not as straightforward.



Lesson Summary

- Atomic size is the distance from the nucleus to the valence shell.
- Atomic size is difficult to measure because it has no definite boundary.
- Atomic radius is a more definite and measurable way of defining atomic size. It is half the distance from the center of one atom to the center of another atom in a homonuclear diatomic molecule.
- There are three factors that help in the prediction of the trends in the periodic table: number of protons in the nucleus, number of energy levels, and the shielding effect.
- The atomic radii increase from top to the bottom in any group.
- The atomic radii decrease from left to right across a period.
- This trend is not as systematic for the transition metals because other factors come into play.

Review Questions

1. Why is the atomic size considered to have “no definite boundary”?
2. How is atomic size measured?

(a) using a spectrophotometer
(b) using a tiny ruler (called a nano ruler)
(c) indirectly
(d) directly

3. Which of the following would be smaller: indium or gallium?
4. Which of the following would be smaller: potassium or cesium?
5. Which of the following would be smaller: titanium or polonium?
6. Explain why iodine is larger than bromine.
7. What are three factors that affect atomic size?
8. Which of the following would have the largest atomic radius?
   (a) Si
   (b) C
   (c) Sn
   (d) Pb
9. Which of the following would have the smallest atomic radius?
   (a) $1s^2 2s^2$
   (b) $1s^2 2s^2 2p^6 3s^1$
   (c) $1s^2$
   (d) $1s^1$
10. Arrange the following in order of increasing atomic radius: Tl, B, Ga, Al, In.
11. Arrange the following in order of increasing atomic radius: Ga, Sn, C.
12. Which of the following would be larger: Rb or Sn?
13. Which of the following would be larger: Ca or As?
14. Describe the trend for the atomic size of elements in a row in the periodic table.
15. Which of the following would have the largest atomic radius?
    (a) Sr
    (b) Sn
    (c) Rb
    (d) In
16. Which of the following would have the smallest atomic radius?
    (a) K
    (b) Kr
    (c) Ga
    (d) Ge
17. Arrange the following in order of decreasing atomic radius: Ba, Tl, Se, Bi, Cs.

1.3 Periodic Trends in Ionic Size

Lesson Objectives

The student will:

- explain what an ion is.
- describe how cations and anions are formed.
- describe the factors that determine the trend in ionic size.
- describe the trend in ionic size for elements.
- use the general trends to predict the relative sizes of ions.

Vocabulary

- anion
- cation
- ion

Introduction

An atom is electrically neutral, which means that the number of protons is equal to the number of electrons. In chemical reactions, however, atoms can gain or lose electrons. This results in the formation of an ion. An **ion** is an atom with a positive or negative charge.

Atoms and Ions

Atoms of metallic elements tend to form positive ions by losing one or more electrons. A positive ion is called a **cation** (pronounced CAT-ion) and has fewer electrons than an electrically neutral atom. For example, an atom of sodium has eleven protons and eleven electrons. Its electron configuration is $[Ne]3s^1$. Sodium has one valence electron surrounding a stable core of ten electrons. In chemical reactions, a sodium atom tends to lose its one valence electron to become a sodium cation. Because this sodium ion has eleven protons and only ten electrons, it has a net charge of $+1$. An atom that loses two electrons will become an ion with a charge of $+2$, and an atom that loses three electrons will become an ion with a charge of $+3$.

Atoms of nonmetallic elements tend to form negative ions by gaining one or more electrons. A negative ion is called an **anion** (pronounced AN-ion). For example, an atom of fluorine has seventeen protons and seventeen electrons. Its electron configuration is $1s^2 2s^2 2p^5$, and it has seven valence electrons. In chemical reactions, a fluorine atom tends to gain one valence electron, becoming a fluoride anion. (Notice that the name of anions typically end in "-ide.") Because the fluoride ion has seventeen positive protons and eighteen negative electrons, it has a net charge of -1. An atom that gains two electrons will become an ion with a charge of -2, and an atom that gains three electrons will become an ion with a charge of -3.

Group and Period Trends in Ionic Size

Cations are smaller than the atoms from which they are formed. The loss of outer shell electrons results in increased attraction between the nucleus and the remaining electrons. This results in less electron-electron repulsion and allows the nucleus and the electrons to come closer together. When compared to a neutral atom of sodium, Na, a sodium cation, symbolized by Na^+, has a smaller size.

Anions are larger than the atoms from which they are formed. The gain of outer shell electrons results in decreased attraction between the nucleus and the remaining electrons, and electron-electron repulsion forces them to spread apart. When compared to a neutral atom of fluorine, F, a fluoride anion, symbolized by F^-, has a larger size.

If we examine the ionic sizes of just the metals in the main group, we will find that the trends are the same as the trends in atomic sizes for the neutral elements. All the elements for each group of metals lose the same number of electrons, which means that the ionic sizes will be primarily affected by the number of energy levels in the electron cloud. Since the number of energy levels still increases from top to bottom, the ionic size also increases down a group of elements in the periodic table. For similar reasons, the trend across a period is the same for both ions and neutral atoms. All the metal elements in a given period will lose their outer shell electrons but still have the same number of core electrons. As a result, the nuclear charge increases from left to right, while the number of core electrons remains the same. This means that

the ion size will decrease from left to right across a period.

Nonmetals also see the same trends in size as the neutral elements. The negative ions increase in size as you move down a group and decrease in size as you move from left to right across a period. In other words, as you go from top to bottom down a group or left to right across a period, the ionic size decreases *as long as* you are comparing all metals or all nonmetals. Between the metals and the nonmetals, the ionic size increases because you are switching from cations, which lose electrons, to anions, which gain electrons.

Lesson Summary

- In chemical reactions, atoms can gain or lose electrons. This results in the formation of an ion. An ion is basically an atom with a positive or negative charge.
- Atoms of metallic elements tend to form positive ions (cations) by losing one or more electrons.
- Atoms of nonmetallic elements tend to form negative ions (anions) by gaining one or more electrons.
- Cations are smaller than the atoms from which they are formed.
- Anions are larger than the atoms from which they were formed.
- Ionic size increases from top to bottom down a group of elements in the periodic table.
- From left to right across a period, the ionic size decreases as long as you are comparing all metals or all nonmetals. Between the metals and nonmetals, the ionic size increases as you switch from cations to anions.

Review Questions

1. How is the size of a cation different from the size of the atom from which it was formed? Why?
2. How is the size of an anion different from the size of the atom from which it was formed? Why?
3. Mg^{2+} has the same number of electrons as F^- (they are said to be isoelectric). Which ion is larger and why?
4. Which of the following has the smallest ionic radius?
   (a) O^{2-}
   (b) S^{2-}
   (c) Mg^{2+}
   (d) Ca^{2+}
5. Which of the following has the largest ionic radius?
   (a) Ba^{2+}
   (b) Cs^+
   (c) I^-
   (d) Te^{2-}

1.4 Periodic Trends in Ionization Energy

Lesson Objectives

The student will:

- define ionization energy.
- describe the trends that exist in the periodic table for ionization energy.
- use the general trends to predict the relative ionization energies of atoms.

Introduction

Atoms are capable of forming ions by either losing or gaining electrons. Since the electrons are attracted to the positively charged nucleus, energy is needed to pull the electron away from the nucleus. In this lesson, we will gain an understanding of the energy required to remove an electron and recognize its trend on the periodic table.

Vocabulary

- effective nuclear charge
- ionization energy

Ionization Energy Defined

Consider lithium, which has an electron configuration of $1s^2 2s^1$ and has one electron in its outermost energy level. In order to remove this electron, energy must be added to the system. Look at the equation below:

$$\begin{array}{ccccccc} \text{energy} & + & \underset{1s^2 2s^1}{\text{Li}_{(g)}} & \rightarrow & \underset{1s^2}{\text{Li}^+_{(g)}} & + & e^- \end{array}$$

With the addition of energy, a lithium atom can lose one electron and form a lithium ion. This energy is known as the ionization energy. The **ionization energy** is the energy required to remove the most loosely held electron from a gaseous atom or ion. The higher the value of the ionization energy, the harder it is to remove that electron. In the equation above, the subscript "g" indicates that the element is in the form of a gas. The definition for ionization energy specifies "in the gaseous phase" because when the atom or ion is in the liquid or solid phases, other factors are involved. The general equation for the ionization energy is as follows.

$$A_{(g)} + \text{energy} \rightarrow A^+_{(g)} + e^- \quad \text{IE}_1$$

If a second electron is to be removed from an atom, the general equation for the ionization energy is as follows:

$$A^+_{(g)} + \text{energy} \rightarrow A^{2+}_{(g)} + e^- \quad \text{IE}_2$$

After the first electron is removed, there are a greater number of protons than electrons. As a result, when a second electron is being removed, the energy required for the second ionization (IE_2) will be greater than the energy required for the first ionization (IE_1). In other words, $\text{IE}_1 < \text{IE}_2 < \text{IE}_3 < \text{IE}_4$.

Group and Period Trends in Ionization Energy

We can see a trend when we look at the ionization energies for the elements in period 2. **Table** 1.5 summarizes the electron configuration and the ionization energies for the elements in the second period.

Table 1.5: **First Ionization Energies for Period 2 Main Group Elements**

Element	**Electron Configuration**	**First Ionization Energy, IE_1**
Lithium (Li)	$[He]2s^1$	520 kJ/mol
Beryllium (Be)	$[He]2s^2$	899 kJ/mol
Boron (B)	$[He]2s^22p^1$	801 kJ/mol
Carbon (C)	$[He]2s^22p^2$	1086 kJ/mol
Nitrogen (N)	$[He]2s^22p^3$	1400 kJ/mol
Oxygen (O)	$[He]2s^22p^4$	1314 kJ/mol
Fluorine (F)	$[He]2s^22p^5$	1680 kJ/mol

We can see that as we move across the period from left to right, in general the ionization energy increases. At the beginning of the period with the alkali metals and the alkaline earth metals, losing one or two electrons allows these atoms to become ions.

$$\begin{array}{ccccc} \text{energy} & + & \underset{[He]2s^1}{Li_{(g)}} & \rightarrow & \underset{[He]}{Li^+_{(g)}} & + e^- \end{array}$$

$$\begin{array}{ccccc} \text{energy} & + & \underset{[Ne]3s^12}{Mg_{(g)}} & \rightarrow & \underset{[Ne]}{Mg^{2+}_{(g)}} & + 2e^- \end{array}$$

As we move across the period, the atoms become smaller, which causes the nucleus to have greater attraction for the valence electrons. Therefore, the electrons are more difficult to remove.

A similar trend can be seen for the elements within a family. **Table** 1.6 shows the electron configuration and the first ionization energies (IE_1) for some of the elements in the first group, the alkali metals.

Table 1.6: **First Ionization Energies for Some Period 1 Elements**

Element	**Electron Configuration**	**First Ionization Energy, IE_1**
Lithium (Li)	$[He]2s^1$	520 kJ/mol
Sodium (Na)	$[Ne]3s^1$	495.5 kJ/mol
Potassium (K)	$[Ar]4s^1$	418.7 kJ/mol

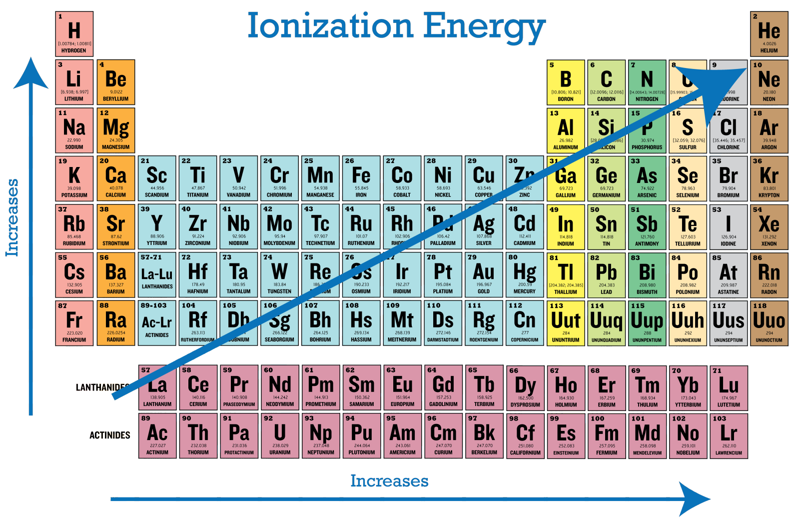
By comparing the electron configurations of lithium to potassium, we know that the valence electron is further away from the nucleus. We know this because the n value is larger, which means that the energy level holding the valence electron is larger. It is easier to remove the most loosely held electron when the electron is further away from the nucleus, because the attractive pull between the nucleus and an electron decreases as the distance between the two increases. Therefore, IE_1 for potassium is less than IE_1 for lithium.

Why does the ionization energy increase going across a period? It has to do with two factors. One factor is that the atomic size decreases. The second factor is that the effective nuclear charge increases. The **effective nuclear charge** is the charge experienced by a specific electron within an atom. Recall that the nuclear charge was used to describe why the atomic size decreased going across a period. **Table** 1.7 shows the effective nuclear charge along with the ionization energy for the elements in period 2.

Table 1.7: **Effective Nuclear Charge for Period 2 Main Group Elements**

Element	Electron Configuration	Number of Protons	Number of Core Electrons	Effective Nuclear Charge	Ionization Energy
Lithium (Li)	$[He]2s^1$	3	2	1	520 kJ/mol
Beryllium (Be)	$[He]2s^2$	4	2	2	899 kJ/mol
Boron (B)	$[He]2s^22p^1$	5	2	3	801 kJ/mol
Carbon (C)	$[He]2s^22p^2$	6	2	4	1086 kJ/mol
Nitrogen (N)	$[He]2s^22p^3$	7	2	5	1400 kJ/mol
Oxygen (O)	$[He]2s^22p^4$	8	2	6	1314 kJ/mol
Fluorine (F)	$[He]2s^22p^5$	9	2	7	1680 kJ/mol

The electrons that are shielding the nuclear charge are the core electrons, which are the $1s^2$ electrons for period 2. The effective nuclear charge is approximately the difference between the total nuclear charge and the number of core electrons. Notice that as the effective nuclear charge increases, the ionization energy also increases. Overall, the general trend for ionization energy is summarized in the diagram below.



Example:

What would be the effective nuclear charge for chlorine? Would you predict the ionization energy to be higher or lower than the ionization energy for fluorine?

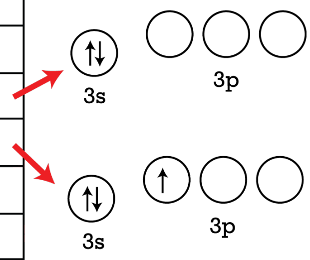
Solution:

Chlorine has the electron configuration: $Cl = [Ne]3s^23p^5$. The effective nuclear charge is 7, which is the same as the nuclear charge for fluorine. Predicting the ionization energy with just this information would be difficult. The atomic size, however, is larger for chlorine than it is for fluorine because chlorine has three energy levels (chlorine is in period 3). Now we can conclude that the ionization energy for chlorine should be lower than that of fluorine because the electron would be easier to pull off when it is further away from the nucleus. (Indeed, the value for the first ionization energy of chlorine is 1251 kJ/mol, compared to 1680 kJ/mol for fluorine.)

A few anomalies exist with respect to the ionization energy trends. Going across a period, there are two ways in which the ionization energy may be affected by the electron configuration. When we look at period 3, we can see that there is an anomaly as we move from the $3s$ sublevels to the $3p$ sublevel. The table below shows the electron configurations and first ionization energy for the main group elements in period

3.

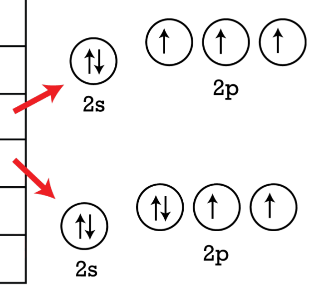
Ionization Energies for Period 3 Main Group Elements		
Element	**Electron Configuration**	**Ionization Energy**
Sodium (Na)	$1s^22s^22p^63s^1$	495.9 kJ/mol
Magnesium (Mg)	$1s^22s^22p^63s^2$	738.1 kJ/mol
Aluminum (Al)	$1s^22s^22p^63s^23p^1$	577.9 kJ/mol
Silicon (Si)	$1s^22s^22p^63s^23p^2$	786.3 kJ/mol
Phosphorus (P)	$1s^22s^22p^63s^23p^3$	1012 kJ/mol
Sulfur (S)	$1s^22s^22p^63s^23p^4$	999.5 kJ/mol
Chlorine (Cl)	$1s^22s^22p^63s^23p^5$	1251 kJ/mol
Argon (Ar)	$1s^22s^22p^63s^23p^6$	1520 kJ/mol



In the table, we see that when we compare magnesium to aluminum, the IE_1 decreases instead of increases. Why is this? Magnesium has its outermost electrons in the 3*s* sub-level. The aluminum atom has its outermost electron in the 3*p* sublevel. Since *p* electrons have just slightly more energy than *s* electrons, it takes a little less energy to remove that electron from aluminum. One other factor is that the electrons in $3s^2$ shield the electron in $3p^1$. These two factors allow the IE_1 for aluminum to be less than IE_1 for magnesium.

When we look again at the table, we can see that the ionization energy for nitrogen also does not follow the general trend.

Ionization Energies for Period 2 Main Group Elements		
Element	**Electron Configuration**	**Ionization Energy**
Lithium (Li)	[He] $2s^1$	520 kJ/mol
Beryllium (Be)	[He] $2s^2$	899 kJ/mol
Boron (B)	[He] $2s^22p^1$	801 kJ/mol
Carbon (C)	[He] $2s^22p^2$	1086 kJ/mol
Nitrogen (N)	[He] $2s^22p^3$	1400 kJ/mol
Oxygen (O)	[He] $2s^22p^4$	1314 kJ/mol
Fluorine (F)	[He] $2s^22p^5$	1680 kJ/mol
Neon (Ne)	[He] $2s^22p^6$	2081 kJ/mol



While nitrogen has one electron occupying each of the three *p* orbitals in the second sub-level, oxygen has an additional electron in one of the three 2*p* orbitals. The presence of two electrons in an orbital lead to greater electron-electron repulsion experienced by these 2*p* electrons, which lowers the amount of energy needed to remove one of these electrons. Therefore, IE_1 for oxygen is less than that for nitrogen.

This video discusses the ionization energy trends in the periodic table **(1c)**: http://www.youtube.com/watch?v=xE9YOBXdTSo (9:25).

Figure 1.2: (Watch Youtube Video)
http://www.ck12.org/flexbook/embed/view/373

Lesson Summary

- Ionization energy is the energy required to remove the most loosely held electron from a gaseous atom or ion. Ionization energy generally increases across a period and decreases down a group.
- Once one electron has been removed, a second electron can be removed, but $IE_1 < IE_2$. If a third electron is removed, $IE_1 < IE_2 < IE_3$, and so on.
- The effective nuclear charge is the charge of the nucleus felt by the valence electrons.
- The effective nuclear charge and the atomic size help explain the trend of ionization energy. Going down a group, the atomic size gets larger and the electrons can be more readily removed. Therefore, ionization energy decreases down a group. Going across a period, both the effective nuclear charge and the ionization energy increases, because the electrons are harder to remove.

Review Questions

1. Define ionization energy and write the general ionization equation.
2. Which of the following would have the largest ionization energy?
   (a) Na
   (b) Al
   (c) H
   (d) He
3. Which of the following would have the smallest ionization energy?
   (a) K
   (b) P
   (c) S
   (d) Ca
4. Place the following elements in order of increasing ionization energy: Na, O, Mg, Ne, K.
5. Place the following elements in order of decreasing ionization energy: N, Si, P, Mg, He.
6. Using experimental data, the first ionization energy for an element was found be 600 kJ/mol. The second ionization energy was found to be 1800 kJ/mol. The third, fourth, and fifth ionization energies were found to be, respectively, 2700 kJ/mol, 11,600 kJ/mol, and 15,000 kJ/mol. To which family of elements does this element belong? Explain.
7. Using electron configurations and your understanding of ionization energy, which would you predict to have a higher second ionization energy: Na or Mg?
8. Comparing the first ionization energies of Ca and Mg,
   (a) calcium has a higher ionization energy because its radius is smaller.
   (b) magnesium has a higher ionization energy because its radius is smaller.
   (c) calcium has a higher ionization energy because it outermost sub-energy level is full.

(d) magnesium has a higher ionization energy because it outermost sub-energy level is full.
(e) they have the same ionization energy because they have the same number of valence electrons.

9. Comparing the first ionization energies of Be and B,

   (a) beryllium has a higher ionization energy because its radius is smaller.
   (b) boron has a higher ionization energy because its radius is smaller.
   (c) beryllium has a higher ionization energy because it outermost sub-energy level is full.
   (d) boron has a higher ionization energy because it outermost sub-energy level is full.
   (e) they have the same ionization energy because boron only has one extra valence electron.

1.5 Periodic Trends in Electronegativity

Lesson Objectives

The student will:

- define electronegativity.
- describe the trends that exist in the periodic table for electronegativity.
- use the general trends to predict the relative electronegativities of atoms.

Vocabulary

- electronegativity

Introduction

Around 1935, the American chemist Linus Pauling developed a scale to describe the attraction an element has for electrons in a chemical bond. In this lesson, we will gain an understanding of this concept and recognize its trend on the periodic table.

Electronegativity Defined

In a molecule, some electrons are shared between the atoms making up the molecule. The ability of an atom in a molecule to attract shared electrons is called **electronegativity**. The higher the electronegativity of an atom, the greater its ability to attract shared electrons. The electronegativity of atoms has been defined in several ways. One method that is widely accepted is that developed by Linus Pauling.

On the Pauling scale, shown below, fluorine is the most electronegative element with an electronegativity of close to 4.0, and cesium and francium are the least electronegative with electronegativities of around 0.7.

Pauling Electronegativity Values

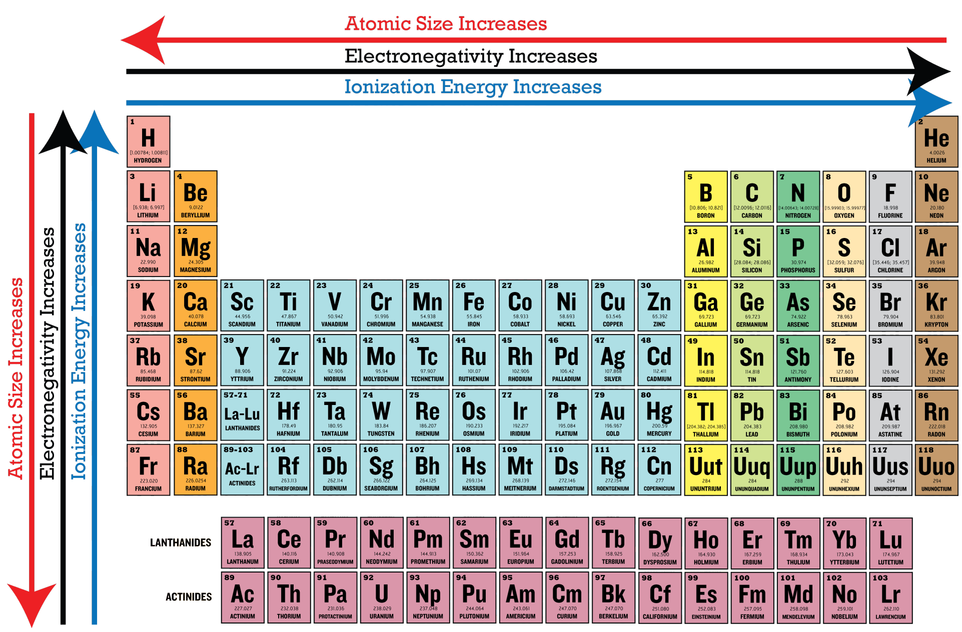
1 H 2.20																
3 Li 0.98	4 Be 1.57											5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98
11 Na 0.93	12 Mg 1.31											13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16
19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96
37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 1.9	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66
55 Cs 0.79	56 Ba 0.89	57 La 1.1	72 Hf 1.3	73 Ta 1.5	74 W 2.36	75 Re 1.9	76 Os 2.2	77 Ir 2.20	78 Pt 2.28	79 Au 2.54	80 Hg 2.00	81 Ti 1.62	82 Pb 2.33	83 Bi 2.02	84 Po 2.0	85 At 2.2
87 Fr 0.7	88 Ra 0.9															

Group and Period Trends in Electronegativity

The electronegativity of atoms increases as you move from left to right across a period in the periodic table. This is because as you go from left to right across a period, the nuclear charge is increasing faster than the electron shielding, so the attraction that the atoms have for the valence electrons increases.

The electronegativity of atoms decreases as you move from top to bottom down a group in the periodic table. This is because as you go from top to bottom down a group, the atoms of each element have an increasing number of energy levels. The electrons in a bond are thus farther away from the nucleus and are held less tightly.

Atoms with low ionization energies have low electronegativities because their nuclei do not have a strong attraction for electrons. Atoms with high ionization energies have high electronegativities because the nucleus has a strong attraction for electrons.



Here is another video that describes ionization energy trends in the periodic table **(1c)**: http://www.youtube.com/watch?v=q3AiM1BYX-c (9:39).

Figure 1.3: (Watch Youtube Video)
http://www.ck12.org/flexbook/embed/view/374

Lesson Summary

- American chemist Linus Pauling developed the electronegativity scale to describe the attraction an element has for electrons in a chemical bond.
- The higher the electronegativity of an atom, the greater its ability to attract shared electrons.
- The electronegativity of atoms increases as you move from left to right across a period in the periodic table.
- The electronegativity of atoms decreases as you move from top to bottom down a group in the periodic table.

Further Reading / Supplemental Links

A series of selectable videos that show the properties and discuss the bonding of various elements.

- http://www.periodicvideos.com/#

Review Questions

1. Define electronegativity.
2. Choose the element in each pair that has the lower electronegativity.
   (a) Li or N
   (b) Cl or Na
   (c) Ca or K
   (d) Mg or F
3. Which of the following will have the largest electronegativity?
   (a) Se
   (b) F
   (c) Ne
   (d) Br
4. Which of the following will have the smallest electronegativity?
   (a) Na
   (b) Ne

(c) Al
(d) Rb

5. Describe the general trend for electronegativity in period 2.

1.6 Periodic Trends in Electron Affinity

Lesson Objectives

The student will:

- define electron affinity.
- describe the trends for electron affinity in the periodic table.

Vocabulary

- electron affinity

Introduction

We have talked about atomic structure, electronic configurations, size of atoms and ions, ionization energy, and electronegativity. The final periodic trend that we will examine is how atoms gain electrons.

Electron Affinity Defined

Atoms can gain or lose electrons. When an atom gains an electron, energy is given off and is known as the electron affinity. **Electron affinity** is defined as the energy released when an electron is added to a gaseous atom or ion.

$$T_{(g)} + e^- \rightarrow T^-_{(g)}$$

For most elements, the addition of an electron to a gaseous atom releases potential energy.

$$\text{Br}_{(g)} + e^- \rightarrow \text{Br}^-_{(g)} \qquad \Delta H = -325 \text{ kJ/mol}$$

Group and Period Trends in Electron Affinity

Let's look at the electron configurations of a few elements and the trend that develops within groups and periods. **Table** 1.8 shows the electron affinities for the halogen family.

Table 1.8: **Electron Affinities for Group 7A**

Element	**Electron Configuration**	**Electron Affinity, kJ/mol**
Fluorine, F	$[\text{He}]2s^22p^5$	−328
Chlorine, Cl	$[\text{Ne}]3s^23p^5$	−349
Bromine, Br	$[\text{Ar}]4s^24p^5$	−325
Iodine, I	$[\text{Kr}]5s^25p^5$	−295

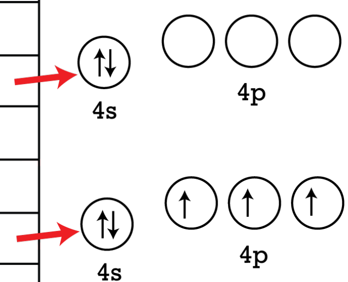
Going down a group, the electron affinity generally decreases because of the increase in size of the atoms. Remember that within a family, atoms located lower on the periodic table are larger because there are more filled energy levels. When an electron is added to a large atom, less energy is released because the electron cannot move as close to the nucleus as it can in a smaller atom. Therefore, as the atoms in a family get larger, the electron affinity gets smaller.

There are exceptions to this trend, especially when comparing the electron affinity of smaller atoms. In **Table** 1.8, the electron affinity for fluorine is less than that for chlorine. This phenomenon is observed in other families as well. The electron affinity of all the elements in the second period is less than the the electron affinity of the elements in the third period. For instance, the electron affinity for oxygen is less than the electron affinity for sulfur. This is most likely due to the fact that the elements in the second period have such small electron clouds ($n = 2$) that electron repulsion of these elements is greater than that of the rest of the family.

Overall, each row in the periodic table shows a general trend similar to the one below.

Electron Affinities for Period 4 Main Group Elements

Element	Electron Configuration	Electron Affinity
Potassium (K)	[Ar] $4s^1$	-48 kJ/mol
Calcium (Ca)	[Ar] $4s^2$	-2.4 kJ/mol
Gallium (Ga)	[Ar] $4s^2 4p^1$	-29 kJ/mol
Germanium (Ge)	[Ar] $4s^2 4p^2$	-118 kJ/mol
Arsenic (As)	[Ar] $4s^2 4p^3$	-77 kJ/mol
Selenium (Se)	[Ar] $4s^2 4p^4$	-195 kJ/mol
Bromine (Br)	[Ar] $4s^2 4p^5$	-325 kJ/mol
Krypton (Kr)	[Ar] $4s^2 4p^6$	0 kJ/mol



The general trend in the electron affinity for atoms is almost the same as the trend for ionization energy. This is because both electron affinity and ionization energy are highly related to atomic size. Large atoms have low ionization energy and low electron affinity. Therefore, they tend to lose electrons. In general, the opposite is true for small atoms. Since they are small, they have high ionization energies and high electron affinities. Therefore, the small atoms tend to gain electrons. The major exception to this rule is the noble gases. Noble gases follow the general trend for ionization energies, but do not follow the general trend for electron affinities. Even though the noble gases are small atoms, their outer energy levels are completely filled with electrons. Any added electron cannot enter their outer most energy level and would have to be the first electron in a new (larger) energy level. This causes the noble gases to have essentially zero electron affinity.

When atoms become ions, the process involves either releasing energy (through electron affinity) or absorbing energy (ionization energy). Therefore, the atoms that require a large amount of energy to release an electron will most likely be the atoms that give off the most energy while accepting an electron. In other words, nonmetals will gain electrons most easily since they have large electron affinities and large ionization energies. Metals will lose electrons since they have the low ionization energies and low electron affinities.

Lesson Summary

- Electron affinity is the energy released when an electron is added to a gaseous atom or ion.
- Electron affinity generally decreases going down a group and increases left to right across a period.
- Nonmetals tend to have the highest electron affinities.

Further Reading / Supplemental Links

This video shows the relationships between atomic size, ionization energy, and electron affinity.

- http://www.youtube.com/watch?v=iCwYjpl8eeY&feature=channel

This pdf document reviews the causes and relationships of the trends in atomic size, ionization energy, electronegativity, and electron affinity.

- http://www.oakland.k12.mi.us/Portals/0/Learning/PeriodicTable.pdf

Review Questions

1. Define electron affinity and write an example equation.
2. Choose the element in each pair that has the lower electron affinity.
   (a) Li or N
   (b) Cl or Na
   (c) Ca or K
   (d) Mg or F
3. Why is the electron affinity for calcium higher than that of potassium?
4. Which of the following will have the largest electron affinity?
   (a) Se
   (b) F
   (c) Ne
   (d) Br
5. Which of the following will have the smallest electron affinity?
   (a) Na
   (b) Ne
   (c) Al
   (d) Rb
6. Place the following elements in order of increasing electron affinity: Tl, Br, S, K, Al.
7. Describe the general trend for electron affinities in period 2.
8. Why does sulfur have a greater electron affinity than phosphorus does?

Image Sources