

Periodic Trends

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CHAPTER

1

Periodic Trends

TEKS

C.5.C Use the Periodic Table to identify and explain periodic trends, including atomic and ionic radii, electronegativity, and ionization energy.

Lesson Objectives

- Learn the periodic trends for atomic radius.
- Know the relationship between group number and valence electrons.
- Describe how ions are formed.
- Learn the periodic trends for ionization energy.
- Explain how multiple ionization energies are related to noble gas electron configurations.
- Describe electron affinity.
- Predict the effect that ion formation has on the size of an atom.
- Learn the periodic trends for electronegativity.

So far, you have learned that the elements are arranged in the periodic table according to their atomic number, and elements in vertical groups share similar electron configurations and chemical properties. In this lesson, we will explore various measurable properties of the elements and how their variation is related to the position of each element on the periodic table. Specifically, we will examine trends within periods and groups. A trend is a general increase or decrease in a particular measurable quantity. For example, as the calendar moves from August to December in the northern hemisphere, the trend is for the average daily temperature to decrease. That doesn't mean that the temperature drops every single day, just that the overall direction is generally downward.

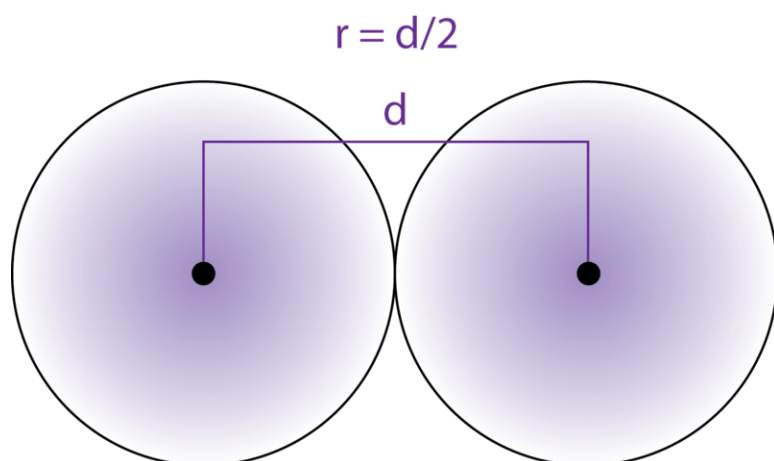
Atomic Radius

One way to define the size of an atom might be to determine the distance from the nucleus to the edge of its electron cloud. However, orbital boundaries are fuzzy, and this distance can vary depending on the conditions under which it is measured. A value that is less variable and easier to measure is the distance between the nuclei of atoms that are bonded together. The **atomic radius** is defined as one-half the distance between the nuclei of identical atoms that are bonded together (**Figure 1.1**).

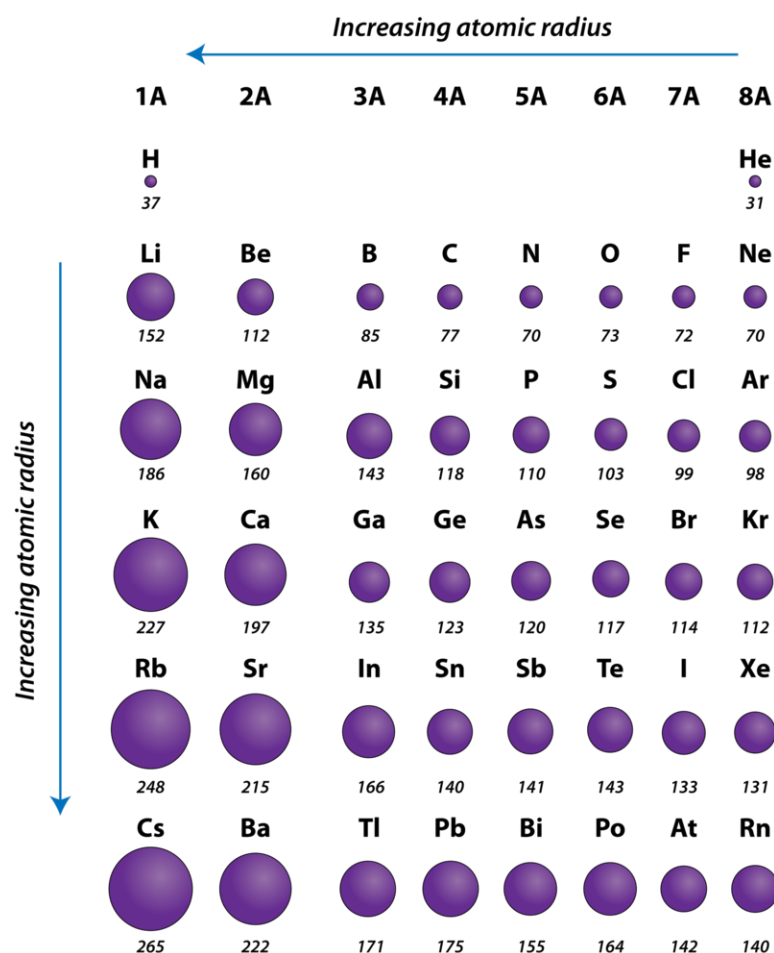
Atomic radii have been measured for most elements, some of which are shown below (**Figure 1.2**). They are commonly reported in units of picometers (recall that $1 \text{ pm} = 10^{-12} \text{ m}$). As an example, the internuclear distance between the two hydrogen atoms in an H_2 molecule is measured to be 74 pm. Therefore, the atomic radius of a hydrogen atom is $74/2 = 37 \text{ pm}$.

Periodic Trend

As you can see from the previous figure (**Figure 1.2**), atomic radius generally decreases from left to right across a period, although there are some small exceptions to this trend, such as the relative radii of oxygen and nitrogen. Within a period, protons are added to the nucleus as electrons are being added to the same principal energy level. These electrons are gradually pulled closer to the nucleus because of its increased positive charge. Since the force of attraction between the nuclei and electrons increases, the size of the atoms decreases. The effect lessens as one moves farther to the right in a period because of electron-electron repulsions that would otherwise cause the atom's size to increase.

**FIGURE 1.1**

The atomic radius (r) of an atom can be defined as one half the distance (d) between two nuclei in a diatomic molecule.

**FIGURE 1.2**

Atomic radii of the representative elements measured in picometers.

Group Trend

As shown above (**Figure 1.2**), the atomic radius of atoms generally increases from top to bottom within a group. As the atomic number increases down a group, there is again an increase in the positive nuclear charge. However, there is also an increase in the number of occupied principal energy levels. Higher principal energy levels consist

of orbitals that are larger in size than the orbitals from lower energy levels. This effect outweighs the increase in nuclear charge, so atomic radius increases down a group.

Pictured below (**Figure 1.3**) is a graph of atomic radius plotted versus atomic number. Each successive period is shown in a different color. You can see that as the atomic number increases within a period, the atomic radius decreases. However, there is a sudden increase in size whenever a new principal energy level is first occupied by electrons.

Atomic radius plotted against atomic number

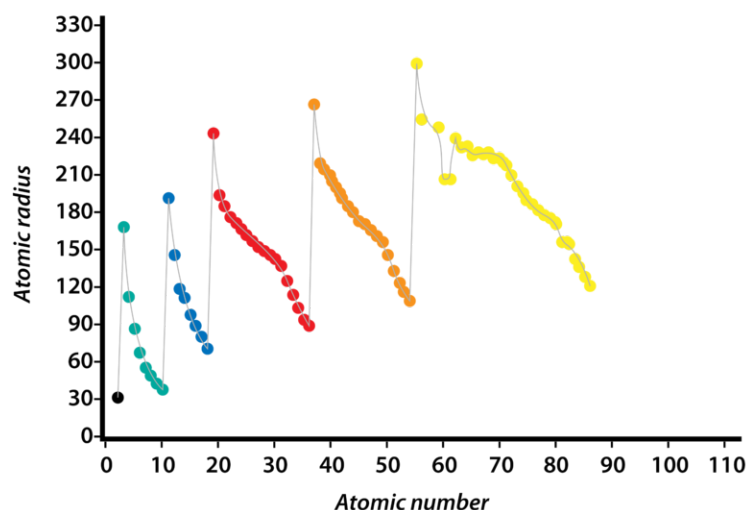


FIGURE 1.3

Graph of atomic radii plotted against atomic number.

Valence Electrons

We have previously defined the valence electrons as those in the outermost occupied principal energy level. Since the *d* and *f* sublevels do not fill until electrons are placed into orbitals from a higher principal energy level, valence electrons will always be in either *s* or *p* sublevels. As a result, the maximum number of valence electrons for an atom in its ground state is eight. Listed below (**Table 1.1**) is the relationship between valence electrons and group number for the representative elements.

TABLE 1.1: Valence Electrons for the Representative Elements

Group Number	Outer Electron Configuration	Number of Valence Electrons
1	ns^1	1
2	ns^2	2
13	ns^2np^1	3
14	ns^2np^2	4
15	ns^2np^3	5
16	ns^2np^4	6
17	ns^2np^5	7
18	ns^2np^6	8

You can see that the number of valence electrons and the outer electron configuration is constant within a group. This is the reason why elements within a group share similar chemical properties.

These concepts are expanded upon in the following video lecture:

- Valence Electrons at <http://www.youtube.com/watch?v=1TZA171yxY4>



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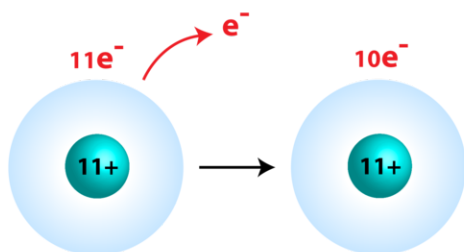
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Forming Ions

Many chemical compounds consist of particles called ions. An **ion** is an atom or group of bonded atoms that has a positive or negative charge. For now, we will only consider monatomic ions, which are single atoms that carry an electrical charge. How do atoms obtain this charge? In theory, there are two possibilities: (1) gaining or losing positively charged protons, or (2) gaining or losing negatively charged electrons. However, the nucleus of an atom is very stable, and the number of protons cannot be changed by chemical reactions. On the other hand, electrons are very capable of movement within an atom. We have already seen this in the form of electron energy level transitions, which are responsible for atomic emission spectra. Ions are formed when an atom gains or loses electrons (**Figure 1.4**).

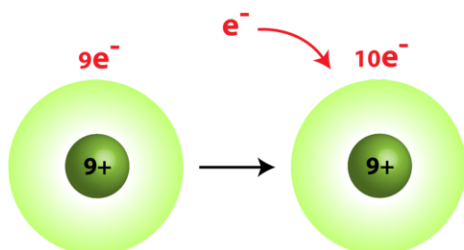
Sodium Atom (Na) \longrightarrow Sodium Ion (Na^+)



Cation Formation

FIGURE 1.4

Fluorine Atom (F) \longrightarrow Fluoride Ion (F^-)



Anion Formation

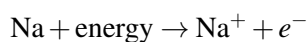
When an atom loses one or more electrons, it becomes positively charged because it now has more protons than electrons. A *positively charged ion* is called a **cation**. The charge for a cation is written as a numerical superscript after the chemical symbol, followed by a plus sign. If the ion carries a single unit of charge, the number “1” is assumed and is not written. For example, a sodium atom that loses one electron becomes a sodium ion, which is

written as Na^+ . A magnesium atom that loses two electrons becomes a magnesium ion, which is written as Mg^{2+} . This magnesium ion carries a 2+ charge because it now has two more protons than electrons.

When an atom gains one or more electrons, it becomes negatively charged because it now has more electrons than protons. A negatively charged ion is called an **anion**. The charge of an anion is written in the same way as the charge of a cation, except a minus sign is used instead of a plus sign. A chlorine atom that gains one electron becomes a chloride ion, which is written as Cl^- . Note that the names of monatomic anions end with an “-ide” suffix. A sulfur atom that gains two electrons becomes a sulfide ion, which is written as S^{2-} .

Ionization Energy

To make an electron jump from a lower energy level to a higher energy level, there must be an input of energy. It stands to reason then, that removing the electron from the atom entirely requires even more energy. This is called an ionization process. **Ionization energy** is the energy required to remove an electron from an atom. An equation can be written to illustrate this process for a sodium atom.



The equation shows that energy added to a sodium atom results in a sodium ion plus the removed electron (e^-). The lost electron is always a valence electron because the electrons in the outermost principal energy level are farthest from the nucleus. The ionization energies of various elements (**Figure 1.5**) are influenced by the size of the atom, the nuclear charge, and the electron energy levels. Ionization energies are measured in units of kilojoules per mole (kJ/mol).

Period Trend

As can be seen in the figures above (**Figure 1.5**), the ionization energy of atoms generally increases from left to right across each row of the periodic table. The reason for this increase in ionization energy is the increase in nuclear charge. A nucleus containing more protons has a larger total positive charge, which results in a greater attractive force being applied to each electron. If the valence electrons are held more tightly to the nucleus by this stronger force, they are more difficult to remove, and more ionization energy is required.

Note that there are several exceptions to the general increase in ionization energy across a period. For example, the Group 13 elements B, Al, and Ga have lower ionization energies than the Group 2 elements from the same period (Be, Mg, and Ca). This is an illustration of a concept called *electron shielding*. Outer electrons are partially shielded from the attractive force of the protons in the nucleus by inner electrons (**Figure 1.7**).

To explain how shielding works, consider a lithium atom, which has three protons and three electrons. Two of its electrons are in the first principal energy level, and its valence electron is in the second. The valence electron is partially shielded from the attractive force of the nucleus by the two inner electrons. Removing that valence electron is easier because of this shielding effect. There is also a shielding effect that occurs between sublevels within the same principal energy level. Specifically, an electron in the *s* sublevel is capable of shielding electrons in the *p* sublevel of the same principal energy level. This is because of the spherical shape of the *s* orbital. The reverse is not true—electrons in *p* orbitals do not shield electrons in *s* orbitals from the same principal energy level (**Figure 1.8**).

The first electron to be removed from an Al atom is a *3p* electron, which is shielded by the two *3s* electrons in addition to all the inner core electrons. The electron being removed from a Mg atom is a *3s* electron, which is only shielded by the inner core electrons. Since there is a greater degree of electron shielding in the Al atom, it is slightly easier to remove its first valence electron. As a result, its ionization energy is less than that of Mg, despite the fact that the nucleus of the Al atom contains one more proton than the nucleus of the Mg atom.

Another anomaly can be found between Groups 15 and 16. Atoms of Group 16 (O, S, etc.) often have lower ionization energies than atoms of Group 15 (N, P, etc.). This can be explained in terms of Hund’s rule. In a nitrogen

1 1A																	18 8A
1 H 1312																	2 He 2372.3
3 Li 520.2	4 Be 899.5											5 B 800.6	6 C 1086.5	7 N 1402.3	8 O 1313.9	9 F 1681.0	10 Ne 2080.7
11 Na 495.8	12 Mg 737.7	3	4	5	6	7	8	9	10	11	12	13 Al 577.5	14 Si 786.5	15 P 1011.8	16 S 999.6	17 Cl 1251.2	18 Ar 1520.6
19 K 418.8	20 Ca 589.8	21 Sc 633.1	22 Ti 658.8	23 V 650.9	24 Cr 652.9	25 Mn 717.3	26 Fe 762.5	27 Co 760.4	28 Ni 737.1	29 Cu 745.5	30 Zn 906.4	31 Ga 578.8	32 Ge 762	33 As 947.0	34 Se 941.0	35 Br 1139.9	36 Kr 1350.8
37 Rb 403.0	38 Sr 549.5	39 Y 600	40 Zr 640.1	41 Nb 652.1	42 Mo 684.3	43 Tc 702	44 Ru 710.2	45 Rh 719.7	46 Pd 804.4	47 Ag 731.0	48 Cd 867.8	49 In 558.3	50 Sn 708.6	51 Sb 834	52 Te 869.3	53 I 1008.4	54 Xe 1170.4
55 Cs 375.7	56 Ba 502.9	57 La* 538.1	72 Hf 658.5	73 Ta 761	74 W 770	75 Re 760	76 Os 840	77 Ir 880	78 Pt 870	79 Au 890.1	80 Hg 1007.1	81 Tl 589.4	82 Pb 715.6	83 Bi 703	84 Po 812.1	85 At 920	86 Rn 1037
87 Fr 380	88 Ra 509.3	89 Ac** 499	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	117 Uus	118 Uuo
LANTHANIDES*		58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu		
ACTINIDES**		90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr		

FIGURE 1.5

A periodic table showing the first ionization energies of the elements in units of kJ/mol.

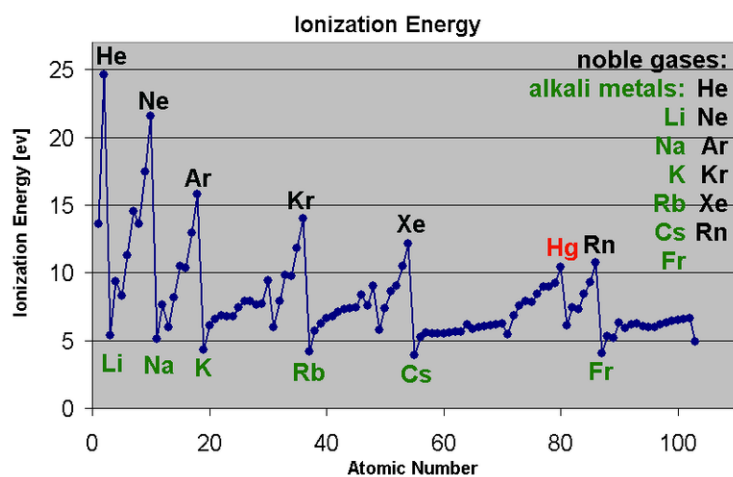


FIGURE 1.6

Graph of first ionization energy plotted against atomic number.

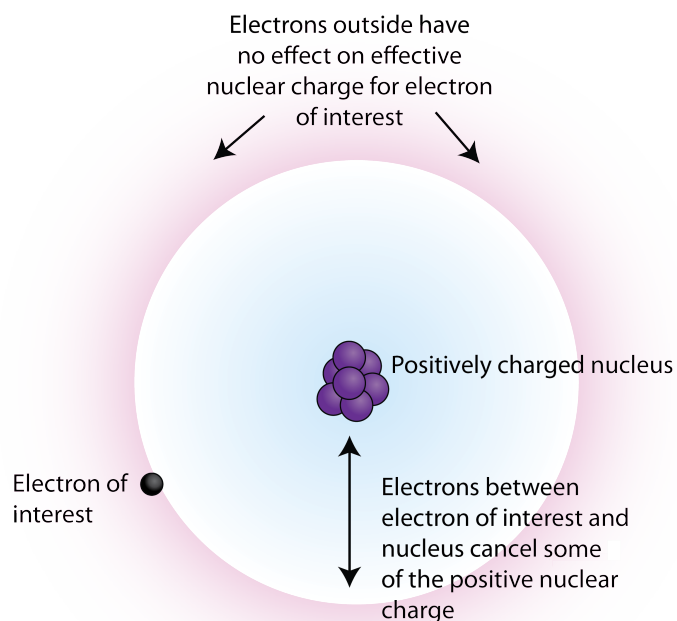


FIGURE 1.7

The interior electron cloud (in blue) shields the outer electrons from the full attractive force of the nucleus. A larger shielding effect results in a decrease in ionization energy.

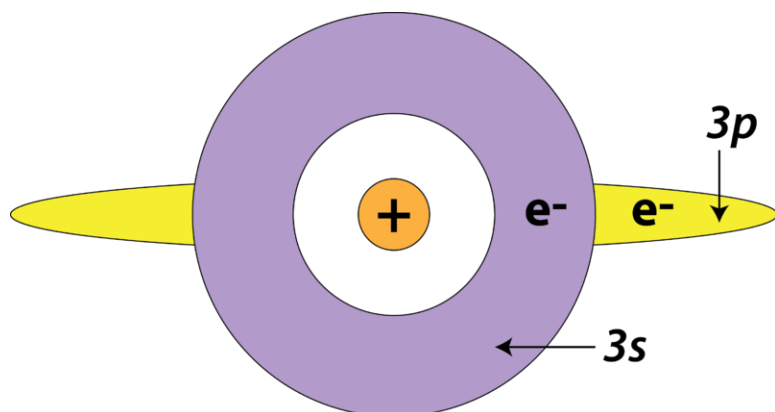


FIGURE 1.8

The spherical 3s orbital exhibits a shielding effect on the dumbbell-shaped 3p orbital, which is slightly higher in energy. This reduces the ionization energy of a 3p electron compared to that of a 3s electron.

atom, there are three unpaired electrons in the $2p$ sublevel. In an oxygen atom, there are four electrons in the $2p$ sublevel, so one orbital must contain a pair of electrons. One electron from this pair will be removed first during the ionization of an oxygen atom. Since electrons repel each other, it is slightly easier to remove an electron from a doubly occupied orbital than to remove an unpaired electron. Consequently, removing a paired electron from oxygen requires slightly less energy than removing an unpaired electron from the nitrogen atom.

Group Trend

The ionization energies of the representative elements generally decrease from top to bottom within a group. This trend is explained by the increase in size of the atoms within a group. The valence electron that is being removed is farther from the nucleus in the case of a larger atom. The attractive force between the valence electron and the nucleus weakens as the distance between them increases, resulting in a lower ionization energy for the larger atoms within a group. Although the nuclear charge is increased for larger atoms, the shielding effect also increases due to

the presence of a larger number of inner electrons. This is particularly easy to see in the alkali metals, where the single valence electron is shielded by all of the inner core electrons.

Multiple Ionizations

So far, we have described first ionization energy and its trends for various atoms. However, in many cases, multiple electrons can be removed from an atom. If an atom loses two electrons, it acquires a 2+ charge. If an atom loses three electrons, it acquires a 3+ charge, and so on. The energies required for subsequent ionizations are called the second ionization energy (IE_2), the third ionization energy (IE_3), and so on. The first six ionization energies are shown for the elements of the first three periods listed below (**Table 1.2**).

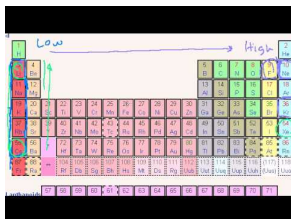
TABLE 1.2: Ionization Energies (kJ/mol) of the First 18 Elements

Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6
H	1312					
He	2373	5251				
Li	520	7300	11,815			
Be	899	1757	14,850	21,005		
B	801	2430	3660	25,000	32,820	
C	1086	2350	4620	6220	38,000	47,261
N	1400	2860	4580	7500	9400	53,000
O	1314	3390	5300	7470	11,000	13,000
F	1680	3370	6050	8400	11,000	15,200
Ne	2080	3950	6120	9370	12,200	15,000
Na	496	4560	6900	9540	13,400	16,600
Mg	738	1450	7730	10,500	13,600	18,000
Al	578	1820	2750	11,600	14,800	18,400
Si	786	1580	3230	4360	16,000	20,000
P	1012	1904	2910	4960	6240	21,000
S	1000	2250	3360	4660	6990	8500
Cl	1251	2297	3820	5160	6540	9300
Ar	1521	2666	3900	5770	7240	8800

Notice that the second ionization energy of an element is always higher than the first, the third is always higher than the second, and so on. This is because after one ionization, a positively charged ion is formed. At this point, there is a greater overall attractive force on the remaining electrons because the protons now outnumber the electrons. Removing a second electron is therefore more difficult.

The first ionization energies for the noble gases (He, Ne, Ar) are higher than those of any other element within that period. The noble gases have full outer s and p sublevels, which gives them extra stability and makes them mostly nonreactive. The stability of the noble gas electron configuration applies to other elements as well. Consider the element lithium, which has a configuration of $1s^2 2s^1$. As an alkali metal, its first ionization energy is very low. After it loses its valence electron (the $2s$ electron), it becomes a lithium ion, Li^+ , which has an electron configuration of $1s^2$. This is the electron configuration of the noble gas helium. The second ionization energy of lithium (shaded above) shows an extremely large jump compared to the first because the removal of a second electron requires breaking apart the noble gas electron configuration. The pattern continues across each period of the table. Beryllium shows a large jump after IE_2 , boron after IE_3 , and so on.

Watch a video lecture about Ionization Energy at <http://www.youtube.com/watch?v=ywqq9PorTAW>



MEDIA

Click image to the left for use the URL below.

URL: <http://www.ck12.org/flx/render/embeddedobject/54714>

Electron Affinity

In most cases, the formation of an anion by the addition of an electron to a neutral atom releases energy. This can be shown for chloride ion formation below:



The energy change that occurs when a neutral atom gains an electron is called its **electron affinity**. When energy is released in a chemical reaction or process, that energy is expressed as a negative number. The figure below (**Figure 1.9**) shows electron affinities in kJ per mole for the representative elements.

1A 1 H -73								8A 2 He >0
3 Li -60	2A 4 Be >0	3A 5 B -27	4A 6 C -122	5A 7 N >0	6A 8 O -141	7A 9 F -328	10 Ne >0	
11 Na -53	12 Mg >0	13 Al -43	14 Si -134	15 P -72	16 S -200	17 Cl -349	18 Ar >0	
19 K -48	20 Ca -4	31 Ga -30	32 Ge -119	33 As -78	34 Se -195	35 Br -325	36 Kr >0	
37 Rb -47	38 Sr -11	49 In -30	50 Sn -107	51 Sb -103	52 Te -190	53 I -295	54 Xe >0	

FIGURE 1.9

Electron affinities (in kJ/mol) for representative elements in the first five periods. Electron affinities are written as negative numbers because energy is being released.

The elements of the halogen group (Group 17) gain electrons most readily, as can be seen from their large negative electron affinities. This means that more energy is released in the formation of a halide ion than for the anions of any other elements. Considering electron configuration, it is easy to see why. The outer configuration of all halogens is ns^2np^5 . The addition of one more electron gives the halide ions the same electron configuration as a noble gas, which we have seen is particularly stable.

Period and group trends for electron affinities are not nearly as regular as those for ionization energy. In general, electron affinities increase (become more negative) from left to right across a period and decrease (become less negative) from top to bottom down a group. However, there are many exceptions.

Ionic Radius

The figure below (**Figure 1.10**) compares the radii of commonly formed ions to the sizes of their parent atoms for Groups 1, 2, 13, 16 and 17. The atoms are shown in gray. Groups 1, 2, and 13 are metals that lose electrons to form

cations, which are shown in green. Groups 16 and 17 are nonmetals that gain electrons to form anions, which are shown in purple.

Size of Atoms and Their Ions in pm

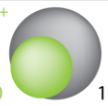
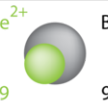
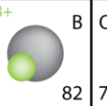
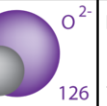
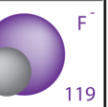

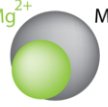
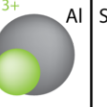
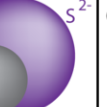
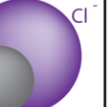
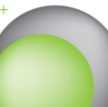
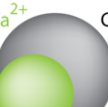
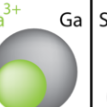
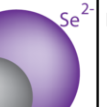
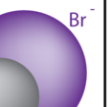
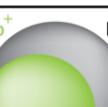
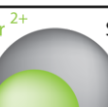
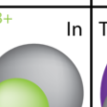
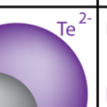
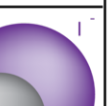
Group 1	Group 2	Group 13	Group 16	Group 17
Li^+  Li 90 134	Be^{2+}  Be 59 90	B^{3+}  B 41 82	O^{2-}  O 73 126	F^-  F 71 119
Na^+  Na 116 154	Mg^{2+}  Mg 86 130	Al^{3+}  Al 68 118	S^{2-}  S 102 170	Cl^-  Cl 99 167
K^+  K 152 196	Ca^{2+}  Ca 114 174	Ga^{3+}  Ga 76 126	Se^{2-}  Se 116 184	Br^-  Br 114 182
Rb^+  Rb 166 211	Sr^{2+}  Sr 132 192	In^{3+}  In 94 144	Te^{2-}  Te 135 207	I^-  I 133 206

FIGURE 1.10

Atomic and ionic radii of the first five elements in Groups 1, 2, 13, 16, and 17. Atoms are shown in gray. The most common ion for each element is shown in either green (for cations) or purple (for anions).

The removal of electrons always results in a cation that is smaller than the parent atom. This is true for any cation because the remaining electrons are drawn closer to the nucleus, now that the protons outnumber the electrons. Additionally, if all of the valence electrons from a given atom are removed, the resulting ion has one fewer occupied principal energy levels, so the electron cloud that remains is considerably smaller.

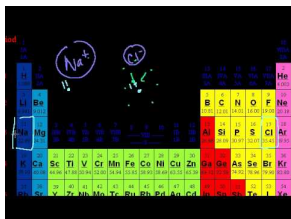
The addition of electrons always results in an anion that is larger than the parent atom. More electrons results in greater electron-electron repulsions, and without any additional protons to cancel this effect, the electron cloud spreads out over a larger volume to minimize repulsive interactions.

Trends

Period and group trends for ionic radii are similar to the trends for atomic radii for the same basic reasons. Going from left to right across the second period, the cations decrease in size because of greater nuclear charge. Starting in Group 15, a nitrogen atom becomes more stable by gaining three electrons to become a nitride ion, N^{3-} , which has a noble gas electron configuration. The nitride ion is larger than the previous cations, but the anions then decrease in size as we move on to Groups 16 and 17. Both types of ions increase in size from top to bottom within a group due to an increase in the number of occupied principal energy levels.

These concepts are expanded upon in the following video lecture:

- Ion Size at <http://www.youtube.com/watch?v=HBi8xjMchZc>



MEDIA

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Electronegativity

Valence electrons of both atoms are always involved when those two atoms come together to form a chemical bond. Chemical bonds are the basis for how elements combine with one another to form compounds. When these chemical bonds form, atoms of some elements have a greater ability to attract the valence electrons involved in the bond than other elements. **Electronegativity** is a measure of the ability of an atom to attract shared electrons when the atom is part of a compound. Electronegativity differs from electron affinity because electron affinity is a measure of the actual energy released when an atom gains an electron. In contrast, electronegativity is a relative scale, so it is not measured in units of energy. All elements are compared to one another, and the most electronegative element, fluorine, is assigned an electronegativity value of 3.98. Fluorine attracts electrons better than any other element. Pictured below (**Figure 1.11**) are the electronegativity values of most elements.

PAULING ELECTRONEGATIVITY VALUES

1 H 2.20																	5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98	
3 Li 0.98	4 Be 1.57																	13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16
11 Na 0.93	12 Mg 1.31	21 Sc	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						
19 K 0.82	20 Ca 1.00	39 Y	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	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 Tl 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																					

FIGURE 1.11

The electronegativity scale was developed by Nobel Prize winning American chemist Linus Pauling. The largest electronegativity (3.98) is assigned to fluorine, and all other electronegativity measurements are made relative to that value.

Since metals have few valence electrons, they tend to increase their stability by losing electrons to become cations. Consequently, the electronegativities of metals are generally low. Nonmetals have more valence electrons and increase their stability by gaining electrons to become anions. The electronegativities of nonmetals are generally high.

Trends

Electronegativities generally increase from left to right across a period. This is due to an increase in nuclear charge. Alkali metals have the lowest electronegativities, while halogens have the highest. Because most noble gases do not form compounds, they are generally not assigned electronegativity values. Note that there is little variation among the transition metals. Electronegativities generally decrease from top to bottom within a group due to the larger atomic size.

Metallic and Nonmetallic Character

Pure elements with a high metallic character are generally very reactive. Metals tend to lose electrons in chemical reactions, as indicated by their low ionization energies. Within a compound, metal atoms have a relatively low attraction to shared electrons, as indicated by their low electronegativity values. By following the trend summary pictured below (**Figure 1.12**), you can see that the most reactive metals would reside in the lower left portion of the periodic table. The most reactive metal that occurs naturally in reasonable quantities is cesium, which is always found in nature as a compound, never as a free element. It reacts explosively with water and will ignite spontaneously in air. Francium is below cesium in the alkali metal group, but it is so rare that many of its properties have never even been observed.

Nonmetals tend to gain electrons in chemical reactions and have a high attraction to electrons within a compound. The most reactive nonmetals reside in the upper right portion of the periodic table. Since the noble gases are an unusually nonreactive group, the element fluorine is the most reactive nonmetal. It is also not found in nature as a free element. Fluorine gas reacts explosively with many other elements and compounds and is considered to be one of the most dangerous known substances.

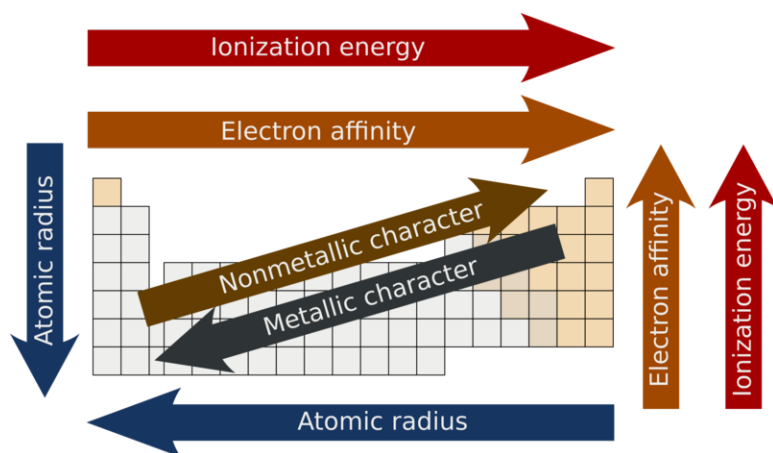


FIGURE 1.12

Summary of periodic trends within periods and groups.

Look at the reactivity of metals in the form of Sumo Wrestlers at <http://freezeray.com/flashFiles/ReactivitySumo.htm>

Lesson Vocabulary

- anion
- atomic radius
- cation
- electron affinity
- electronegativity

- ion
- ionization energy

Lesson Summary

- Atomic radius generally decreases from left to right across a period and increases from top to bottom within a group.
- The number of valence electrons varies from one to eight among the groups of representative elements.
- Metals tend to lose electrons easily to form positively-charged cations, while nonmetals tend to gain electrons to form negatively-charged anions.
- Ionization energy is the energy required to remove an electron. It generally increases from left to right across a period and decreases from top to bottom within a group.
- Electron affinity is the energy released when an atom gains an electron. Its values are largest (most negative) for elements in the halogen group.
- Cations are smaller than their parent atom, while anions are larger than their parent atom.
- Electronegativity is a measure of the attraction for electrons. It generally increases from left to right across a period and decreases from top to bottom within a group.
- Metallic reactivity is greatest for elements in the lower left area of the periodic table, while nonmetallic reactivity is greatest for elements in the upper right area.

Lesson Review Questions

Reviewing Concepts

1. Answer the following:
 - a. How does atomic radius change from left to right across a period? Explain.
 - b. How does atomic radius change from top to bottom within a group? Explain.
2. Answer the following:
 - a. How does ionization energy change from left to right across a period? Explain.
 - b. How does ionization energy change from top to bottom within a group? Explain.
3. Why is the second ionization energy of an element always larger than the first?
4. Why are most electron affinities negative numbers?
5. Answer the following:
 - a. Which group of elements has the highest electronegativities?
 - b. Which group of elements has the lowest electronegativities?

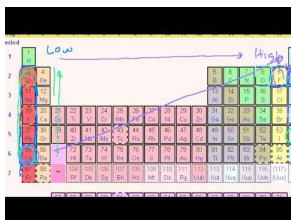
Problems

6. Which element in each pair below has the larger atomic radius?
 - a. K, Na
 - b. S, Cl
 - c. F, Br
 - d. Ba, Cs
7. Which equation below shows the second ionization of an alkaline earth metal?
 - a. $\text{Sr} \rightarrow \text{Sr}^+ + e^-$
 - b. $\text{Sr}^+ \rightarrow \text{Sr}^{2+} + e^-$
 - c. $\text{Cs} \rightarrow \text{Cs}^+ + e^-$
 - d. $\text{Cs}^+ \rightarrow \text{Cs}^{2+} + e^-$

8. Which element in each pair has the higher first ionization energy?
 - a. Mg, P
 - b. Se, O
 - c. Li, Rb
 - d. Ne, N
9. Why is the difference between the second and third ionization energies (IE_2 and IE_3) of calcium so much larger than the difference between its first and second ionization energies?
10. Which atom/ion of each pair is larger?
 - a. Na, Na^+
 - b. Br, Br^-
 - c. Be^{2+} , Ca^{2+}
 - d. F^- , O^{2-}
11. Which element in each pair has a higher electronegativity value?
 - a. F, Cl
 - b. P, S
 - c. Sr, Be
 - d. Al, Na
12. The electron affinity of a halogen atom is the energy released by which of the following processes?
 - a. $Cl \rightarrow Cl^+ + e^-$
 - b. $S + e^- \rightarrow S^-$
 - c. $Br + e^- \rightarrow Br^-$
 - d. $Na \rightarrow Na^+ + e^-$
13. Arrange the following elements in order of increasing metallic character: K, Al, Cs, Na.
14. Arrange the following elements in order of increasing nonmetallic character: O, P, F, N.
15. Explain why the Na^{2+} ion is very unlikely to form.

Further Reading / Supplemental Links

- Watch a video lecture on various Periodic Table Trends at <http://www.youtube.com/watch?v=XMLd-O6PgVs>



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3. Christopher Auyeung. [CK-12 Foundation](#) . CC BY-NC 3.0
4. Christopher Auyeung. [CK-12 Foundation](#) . CC BY-NC 3.0

5. Christopher Auyeung. [CK-12 Foundation](#) . CC BY-NC 3.0
6. User:JJnoDog/De.Wikipedia. <http://commons.wikimedia.org/wiki/File:IonizationEnergyAtomicWeight.PNG> . Public Domain
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