Periodic Trends in Ionization Energy

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CONCEPT 1 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
the net charge experienced by a specific electron within an atom

ionization energy
the energy required to remove the most loosely held electron from a gaseous atom or ion

Ionization Energy Defined

Consider lithium, which has an electron configuration of $1s^22s^1$ 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:

$$\begin{array}{lcccccc} \text{energy} & + & \text{Li}_{(g)} & \rightarrow & \text{Li}^+_{(g)} & + e^- \\ & & 1s^22s^1 & & 1s^2 & \end{array}$$

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** 1.1 summarizes the electron configuration and the ionization energies for the elements in the second period.

TABLE 1.1: First Ionization Energies for Period 2 Main Group Elements

Element	**Electron Configuration**	**First Ionization Energy, IE$_1$**
Lithium (Li)	$[\text{He}]2s^1$	520 kJ/mol
Beryllium (Be)	$[\text{He}]2s^2$	899 kJ/mol
Boron (B)	$[\text{He}]2s^2 2p^1$	801 kJ/mol
Carbon (C)	$[\text{He}]2s^2 2p^2$	1086 kJ/mol
Nitrogen (N)	$[\text{He}]2s^2 2p^3$	1400 kJ/mol
Oxygen (O)	$[\text{He}]2s^2 2p^4$	1314 kJ/mol
Fluorine (F)	$[\text{He}]2s^2 2p^5$	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.

$$\begin{array}{ccccc} \text{energy} & + & \underset{[\text{He}]2s^1}{\text{Li}_{(g)}} & \rightarrow & \underset{[\text{He}]}{\text{Li}^{+}_{(g)}} & + e^{-} \\ \\ \text{energy} & + & \underset{[\text{Ne}]3s^1 2}{\text{Mg}_{(g)}} & \rightarrow & \underset{[\text{Ne}]}{\text{Mg}^{2+}_{(g)}} & + 2e^{-} \end{array}$$

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** 1.2 shows the electron configuration and the first ionization energies (IE_1) for some of the elements in the first group, the alkali metals.

TABLE 1.2: First Ionization Energies for Some Period 1 Elements

Element	**Electron Configuration**	**First Ionization Energy, IE$_1$**
Lithium (Li)	$[\text{He}]2s^1$	520 kJ/mol
Sodium (Na)	$[\text{Ne}]3s^1$	495.5 kJ/mol
Potassium (K)	$[\text{Ar}]4s^1$	418.7 kJ/mol

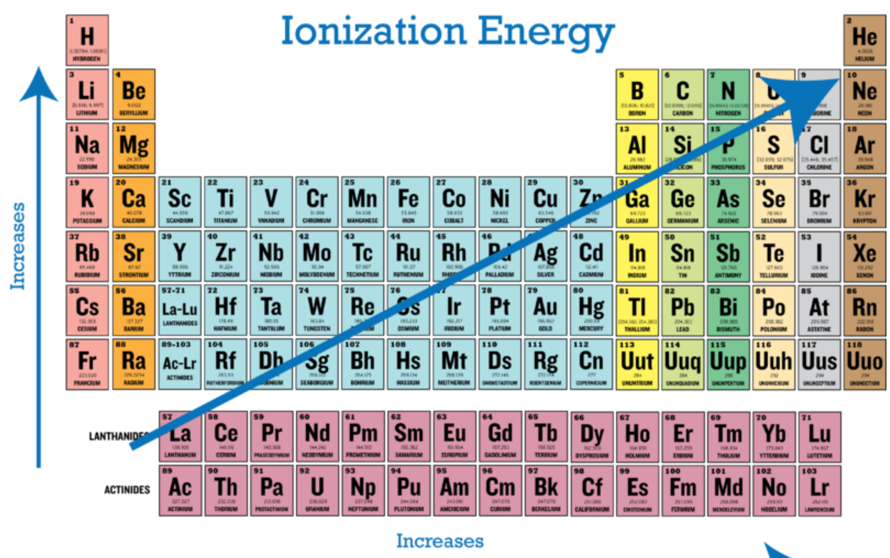
By comparing the electron configurations of lithium to potassium, we know that the valence electron is further away from the nucleus. We know this because the n value is larger, which means that the energy level holding the valence electron is larger. It is easier to remove the most loosely held electron 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** 1.3 shows the effective nuclear charge along with the ionization energy for the elements in period 2.

TABLE 1.3: 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?

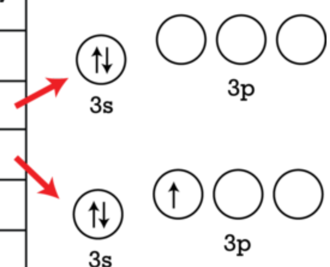
Solution:

Chlorine has the electron configuration: $Cl = [Ne]3s^2 3p^5$. 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)	$1s^2 2s^2 2p^6 3s^1$	495.9 kJ/mol
Magnesium (Mg)	$1s^2 2s^2 2p^6 3s^2$	738.1 kJ/mol
Aluminum (Al)	$1s^2 2s^2 2p^6 3s^2 3p^1$	577.9 kJ/mol
Silicon (Si)	$1s^2 2s^2 2p^6 3s^2 3p^2$	786.3 kJ/mol
Phosphorus (P)	$1s^2 2s^2 2p^6 3s^2 3p^3$	1012 kJ/mol
Sulfur (S)	$1s^2 2s^2 2p^6 3s^2 3p^4$	999.5 kJ/mol
Chlorine (Cl)	$1s^2 2s^2 2p^6 3s^2 3p^5$	1251 kJ/mol
Argon (Ar)	$1s^2 2s^2 2p^6 3s^2 3p^6$	1520 kJ/mol

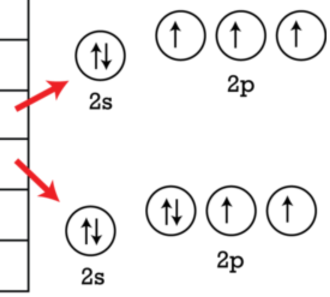


In the table, we see that when we compare magnesium to aluminum, the IE_1 decreases instead of increases. Why is this? Magnesium has its outermost electrons in the $3s$ sub-level. The aluminum atom has its outermost electron in the $3p$ sublevel. Since p electrons have just slightly more energy than 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.

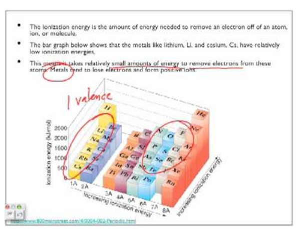
Ionization Energies for Period 2 Main Group Elements

Element	Electron Configuration	Ionization Energy
Lithium (Li)	[He] $2s^1$	520 kJ/mol
Beryllium (Be)	[He] $2s^2$	899 kJ/mol
Boron (B)	[He] $2s^2 2p^1$	801 kJ/mol
Carbon (C)	[He] $2s^2 2p^2$	1086 kJ/mol
Nitrogen (N)	[He] $2s^2 2p^3$	1400 kJ/mol
Oxygen (O)	[He] $2s^2 2p^4$	1314 kJ/mol
Fluorine (F)	[He] $2s^2 2p^5$	1680 kJ/mol
Neon (Ne)	[He] $2s^2 2p^6$	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 $2p$ orbitals. The presence of two electrons in an orbital lead to greater electron-electron repulsion experienced by these $2p$ 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).



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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.