

Electron Configuration and the Periodic Table

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Bewick

Parsons

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Authors

Sharon Bewick, Richard Parsons

Editor

Shonna Robinson

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Chapter 1

Electron Configuration and the Periodic Table

1.1 Mendeleev's Periodic Table

Lesson Objectives

The student will:

- identify the person credited for organizing the periodic table.
- state the basis for the organization of Mendeleev's periodic table.

Vocabulary

- periodic table

Introduction

During the 1800s, when most of the elements were being discovered, many chemists tried to classify the elements according to their similarities. In 1829, Johann Dobringer noted chemical similarities in several groups of three elements and placed these elements into what he called triads. His groupings included the triads of 1) chlorine, bromine, and iodine; 2) sulfur, selenium, and tellurium; 3) calcium, strontium, and barium; and 4) lithium, sodium, and potassium.

In 1864, John Newlands saw a connection between the chemical properties of elements and their atomic masses. He stated that if the known elements, beginning with lithium, are arranged in order of increasing mass, the eighth element will have properties similar to the first element, the ninth similar to the second, the tenth similar to the third, and so on. Newlands called his relationship the law of octaves, comparing the elements to the notes in a musical scale. Newlands tried to force all the known elements to fit into his octaves, but many of the heavier elements, when discovered, did not fit into his patterns.

Mendeleev's Table Based on Chemical Behavior

The **periodic table** is a tabular arrangement of the chemical elements according to atomic number. In 1869, Dmitri Mendeleev, unaware of Newlands's work, also developed a table of the elements that showed a relationship between the chemical properties of elements and their atomic masses. His table would be similar to the one we use today. Mendeleev avoided Newlands's mistake of trying to force elements into groups where their chemistry did not match. Instead, he left empty spaces in his table for where no known element fit the properties and atomic mass for that space. He assumed that elements discovered in the future would fit in those spots. Mendeleev even predicted some of the chemical and physical properties the undiscovered elements would have.

Mendeleev's table listed the elements in order of increasing atomic mass and grouped elements with similar chemical behavior together. For example, lithium is a shiny metal that is soft enough to be cut with a spoon. It reacts readily with oxygen and violently with water. When lithium reacts with water, it produces hydrogen gas and lithium hydroxide. As we examine the next few elements in order of increasing mass, we will come to the element sodium. Sodium is also a shiny metal soft enough to be cut with a spoon. It reacts readily with oxygen and violently with water. When sodium reacts with water, it produces hydrogen gas and sodium hydroxide. You should note that the description of the chemical behavior of sodium is very similar to the chemical description of lithium. Whenever Mendeleev found an element whose chemistry was very similar to a previous element, he placed it below that similar element.

3 Li 1.941 Lithium	4 Be 9.0122 Beryllium	5 B 10.811 Boron	6 C 12.011 Carbon	7 N 14.007 Nitrogen	8 O 15.999 Oxygen	9 F 18.998 Fluorine	10 Ne 20.180 Neon	11 Na 22.990 Sodium
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As each succeeding element was considered, Mendeleev continued to place an element underneath other elements with similar chemistry. For this reason, Mendeleev placed magnesium below beryllium, aluminum under boron, and so on.

3 Li 1.941 Lithium	4 Be 9.0122 Beryllium	5 B 10.811 Boron	6 C 12.011 Carbon	7 N 14.007 Nitrogen	8 O 15.999 Oxygen	9 F 18.998 Fluorine	10 Ne 20.180 Neon
11 Na 22.990 Sodium	12 Mg 24.305 Magnesium	13 Al 26.982 Aluminium	14 Si 28.086 Silicon	15 P 30.974 Phosphorus	16 S 32.065 Sulfur	17 Cl 35.453 Chlorine	18 Ar 39.948 Argon
19 K 39.098 Potassium							

When Mendeleev came to potassium, he found that it had very similar chemistry to the chemistry of lithium and sodium, so he placed potassium under sodium. You may notice that the atomic masses are no longer in increasing order with respect to potassium and argon. Although argon has a greater mass than potassium, the chemistry of argon was identical to that of neon, and the chemistry of potassium was almost identical to that of sodium. As a result, Mendeleev did not try to put argon beneath sodium. Instead, he allowed the chemical behavior determine where each element should be placed. As you go through the periodic table, you will find several elements whose atomic masses are not in increasing order.

Mendeleev's procedure was to place the elements in the column with matching chemical behavior regardless of atomic mass, so if no element with appropriate chemistry was known, the space would be left empty. For

3 Li Lithium	4 Be Beryllium																	5 B Boron	6 C Carbon	7 N Nitrogen	8 O Oxygen	9 F Fluorine	10 Ne Neon
11 Na Sodium	12 Mg Magnesium																	13 Al Aluminium	14 Si Silicon	15 P Phosphorus	16 S Sulfur	17 Cl Chlorine	18 Ar Argon
19 K Potassium	20 Ca Calcium	21 Sc Scandium	22 Ti Titanium	23 V Vanadium	24 Cr Chromium	25 Mn Manganese	26 Fe Iron	27 Co Cobalt	28 Ni Nickel	29 Cu Copper	30 Zn Zinc	31 Ga Gallium		32 Ge Germanium	33 As Arsenic	34 Se Selenium	35 Br Bromine	36 Kr Krypton					

- The periodic table in its present form was organized by Dmitri Mendeleev.
- The basis for Mendeleev's organization was the chemical behavior of the elements.

The following websites explore the history behind the periodic table.

1. On what basis did Mendeleev organize his periodic table?
 - (a) atomic number and atomic mass
 - (b) atomic mass and chemical behavior
 - (c) atomic number of number of valence electrons
 - (d) atomic mass and the Latin alphabet
 - (e) none of these

- identify groups in the periodic table.
- state the number of valence electrons for each A group in the periodic table.
- explain the relationship between the chemical behavior of families in the periodic table and their electron configuration.
- identify periods in the periodic table.
- describe the similarities among elements in the same period in the periodic table.

Vocabulary

- actinide series
- alkali metals
- alkaline earth metals
- group (family)
- halogens
- lanthanide series
- noble gases
- period
- transition elements

Introduction

When Mendeleev created his periodic table, he did not know why certain elements had similar chemistry. He placed the elements in their positions because they exhibited similar chemical behaviors. Thus, the vertical columns in Mendeleev's table were composed of elements with similar chemistry. These vertical columns are called **groups**, or families. In this section, you are going to see that the elements in the same groups are related to each other by their electron configurations. Since the families of elements were organized by their chemical behavior, it is predictable that the individual members of each chemical family will have similar electron configurations.

If you examine a periodic table, you will often find a number written above each group (column). These numbers serve as labels, and groups are often referred to by their labels. Depending on the source or age of your periodic table, you may see two different numbering systems for referring to the families on the periodic table. In the older system, the numbers 1 – 8 and the letters A and B were used to label the groups. The newer convention is to label each group from 1 – 18 in sequential order. However, the older labeling scheme helps to provide more insight into the electron configurations of each group. As a result, in this text we will use the older labeling scheme to present each group. The periodic table below shows both numbering systems.

[illegible]

For an introduction to the electronic organization of the periodic table (**1a**, **1c**, **1d**, **1f**), see <http://www.youtube.com/watch?v=5MMWpeJ5dn4> (3:51).



Figure 1.1: ([Watch Youtube Video](http://www.ck12.org/flexbook/embed/view/370))
<http://www.ck12.org/flexbook/embed/view/370>

Group 1A

The electron configuration codes for the elements in Group 1A are:

lithium = $1s^2 2s^1$ or $[\text{He}]2s^1$
sodium = $1s^2 2s^2 2p^6 3s^1$ or $[\text{Ne}]3s^1$
potassium = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ or $[\text{Ar}]4s^1$
rubidium = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$ or $[\text{Kr}]5s^1$
cesium = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$ or $[\text{Xe}]6s^1$
francium = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^1$ or $[\text{Rn}]7s^1$

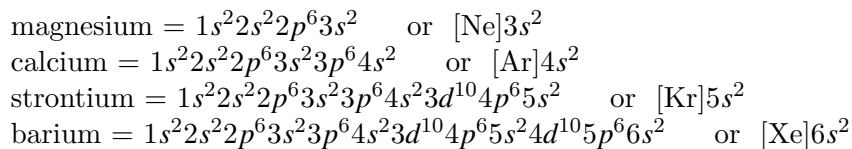
The fact that all 1A elements participate in similar chemistry despite having vastly different nuclear sizes further illustrates the fact that electrons, particularly valence electrons, are the primary contributors to chemical reactivity. The electron configuration for the outermost energy levels of the 1A elements is the same, with the only difference being the energy level involved. Each larger member of the family has its single *s* electron in the next larger principal energy level. As the atomic sizes in this family increase, the valence electrons are located further from the nucleus and are therefore easier to lose. Lithium reacts readily with water, sodium reacts violently with water, potassium reacts so violently that the hydrogen formed begins to burn, and rubidium explodes instantly in water. The differences in these chemical reactions are just a matter of degree. Each larger atom reacts in the same way as the one before, but it reacts faster. The speed and ease of reaction of elements is referred to as reactivity. The relative reactivity of the elements in the 1A family (also called **alkali metals**) increases as the atoms become larger. Note that in the new labeling scheme, this group is labeled as Group 1.

All of the 1A elements have one valence electron and react in similar ways as the other members of the family – that’s why they are a part of the same family. The 1A metals have similar electron configurations, which causes them to have similar chemistry. Although Mendeleev did not know about these elements had similar electron configurations, he recognized the similar chemistry and organized the periodic table based on this similar chemistry.

Group 2A

Here are the electron configuration codes for the first five elements in Group 2A:

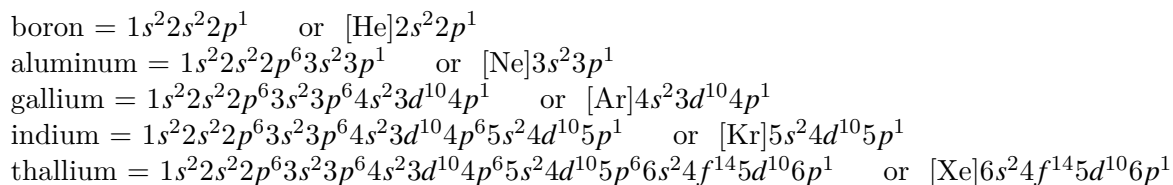
beryllium = $1s^2 2s^2$ or $[\text{He}]2s^2$



All of the elements in this family have two valence electrons and have very similar chemistry. This group of metals is called the **alkali earth metals**. As with the 1A family, the elements in this family also increase in reactivity as the elements become larger and the valence electrons are held more loosely. In the new labeling conventions, this group is labeled as Group 2.

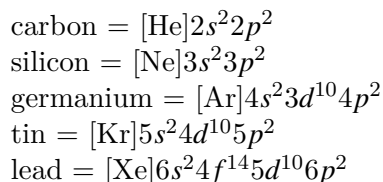
Group 3A

The electron configurations of Group 3A (Group 3 according to new labeling conventions) show that all the members of this family have three valence electrons. The chemical behaviors of the elements in this family are similar to each other, but they are not as consistent from element to element as they are for other families of elements. Information in the following sections will explain why.

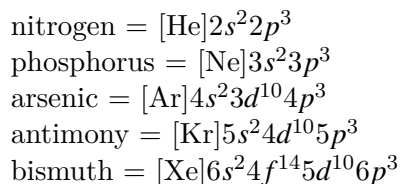


Groups 4A through 8A

Group 4A (also known as Group 14) members have four valence electrons.



Group 5A (or Group 15) members have five valence electrons.



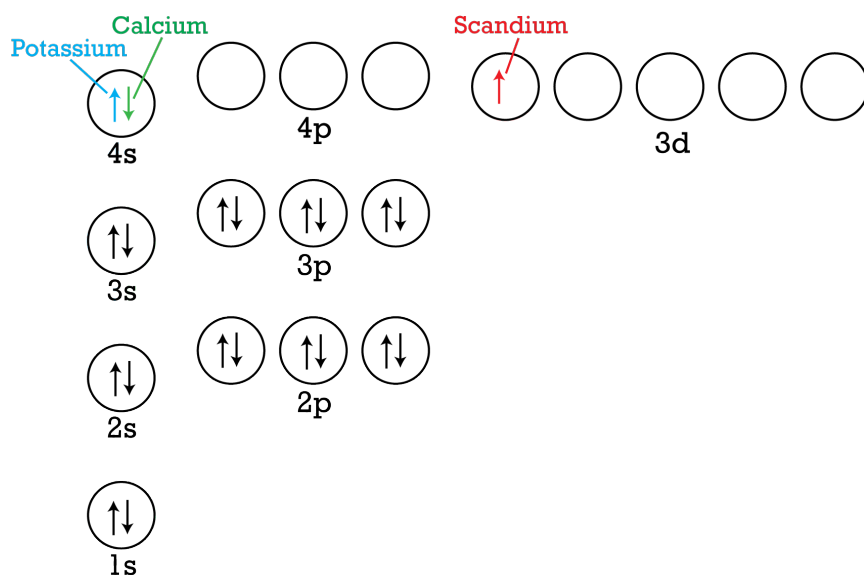
Instead of showing you the electron configurations for Groups 6A (or Group 16), 7A (or 17), and 8A (or 18), it would be good practice for you to practice writing the electron configurations for these families. When you finish writing them, check to make sure that the outer energy levels contain six valence electrons, seven valence electrons, and eight valence electrons for Groups 6A, 7A, and 8A, respectively. The group label provides a hint about the valence electrons: the number preceding the “A” is equal to the number of valence electrons for each atom in the group.

Groups 7A and 8A are also known by the names **halogens** and the **noble gases**, respectively. Groups 7A and 8A were two of the first families identified because the chemistry of their members are so similar to each other.

Transition Elements

If you have been checking where the different groups that have been introduced so far are located on the periodic table, you may have noticed that there are a series of elements that are not part of Groups 1A – 8A (sometimes refer to as the main group). In some periodic tables, these groups of elements are in families called B groups, and in newer periodic tables, these groups are numbered 3 – 12.

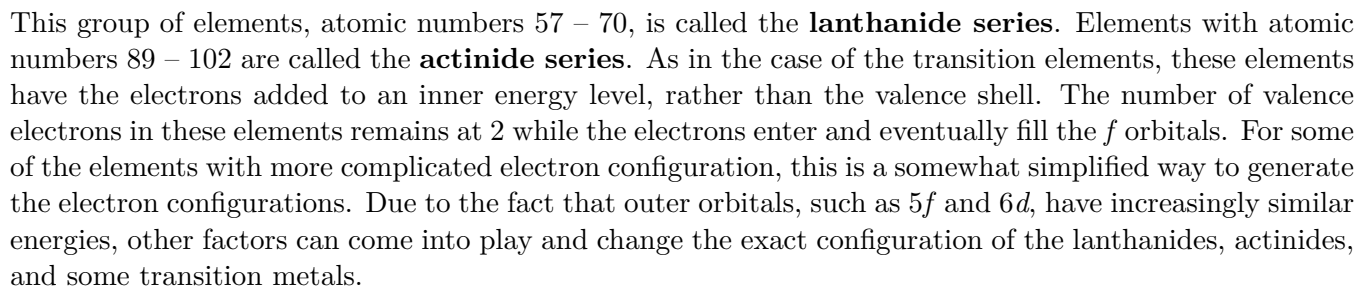
Consider the fourth horizontal row in the periodic table. For potassium, the first element in this row, one electron resides in the $4s$ orbital. For calcium, a second electron is added to the $4s$ orbital. Beyond calcium, however, the pattern in which the electrons are added changes. Beginning with scandium, atomic number 21, the additional electrons do not enter the valence shell but instead enter the **d** sub-level of the third energy level ($n = 3$), as illustrated below. The electron configuration for scandium is $[\text{Ar}]4s^23d^1$.



The ten elements formed by filling in the $3d$ orbitals, as well as all other elements that have between 1 to 10 electrons in d orbitals, are called the **transition elements**. These elements, in general, differ from each other in the electron structure of the next-to-last energy level. For the most part, these elements have similar outer energy levels because they have two valence electrons. As a result, they have somewhat similar chemistry, even though they are not in the same vertical family.

The Lanthanide and Actinide Series

There is still one more block of elements on the periodic table that has not been introduced. This block is usually placed below the periodic table and represents elements with electrons in the f orbitals. In the sixth horizontal row, the first electron goes into the $6s$ orbital and produces the element cesium. The second electron fills the $6s$ orbital producing the element barium. The next 14 electrons that are added then enter the $4f$ orbitals (marked in red in the figure below).



For Groups 1A and 2A, the last electron added enters an s orbital. For this reason, the 1A and 2A families make up an area on the periodic table known as the “s block.” In Groups 3A, 4A, 5A, 6A, 7A, and 8A, the last electron enters a p orbital. These six families of elements, therefore, make up an area called the “p block.” The various sub-level blocks, including the “d block” and “f block,” are indicated in the periodic table shown below.

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Even though the periodic table was organized according to the chemical behavior of the elements, you can now see that the shape and design of the table is a perfect reflection of the electron configuration of the atoms. This is because the chemical behavior of the elements is also dictated by the electron configuration of the atoms.

Periods of the Periodic Table

In addition to providing information about the electron configuration, an element's position on the periodic table can also be used to figure out the energy level of the element's valence electrons. Let's try to figure out what we can learn from an element's period in the periodic table. A **period** is a horizontal row of elements on the periodic table. The figure below shows how the different rows in the periodic table are numbered.

	1 1A							18 8A
ROW 1	1 H [1.00784; 1.00811] HYDROGEN	2 2A	13 3A	14 4A	15 5A	16 6A	17 7A	2 He 4.0026 HELIUM
ROW 2	3 Li [6.938; 6.997] LITHIUM	4 Be 9.0122 BERYLLIUM	5 B [10.806; 10.821] BORON	6 C [12.0096; 12.0116] CARBON	7 N [14.00643; 14.00728] NITROGEN	8 O [15.99903; 15.99977] OXYGEN	9 F 18.998 FLUORINE	10 Ne 20.180 NEON
ROW 3	11 Na 22.990 SODIUM	12 Mg 24.305 MAGNESIUM	13 Al 26.982 ALUMINUM	14 Si [28.084; 28.086] SILICON	15 P 30.974 PHOSPHORUS	16 S [32.059; 32.076] SULFUR	17 Cl [35.446; 35.457] CHLORINE	18 Ar 39.948 ARGON
ROW 4	19 K 39.098 POTASSIUM	20 Ca 40.078 CALCIUM	31 Ga 69.723 GALLIUM	32 Ge 69.723 GERMANIUM	33 As 74.922 ARSENIC	34 Se 78.963 SELENIUM	35 Br 79.904 BROMINE	36 Kr 83.801 KRYPTON
ROW 5	37 Rb 85.468 RUBIDIUM	38 Sr 87.62 STRONTIUM	49 In 114.818 INDIUM	50 Sn 114.818 TIN	51 Sb 121.760 ANTIMONY	52 Te 127.603 TELLURIUM	53 I 126.904 IODINE	54 Xe 131.292 XENON
ROW 6	55 Cs 132.905 CESIUM	56 Ba 137.327 BARIUM	81 Tl [204.382; 204.385] THALLIUM	82 Pb 204.383 LEAD	83 Bi 208.980 BISMUTH	84 Po 208.982 POLONIUM	85 At 209.987 ASTATINE	86 Rn 222.018 RADON
ROW 7	87 Fr 223.020 FRANCIUM	88 Ra 226.0254 RADIUM	113 Uut 284 UNUNTRIUM	114 Uuq 284 UNUNQUADIUM	115 Uup 288 UNUNPENTIUM	116 Uuh 292 UNUNHEXIUM	117 Uus 294 UNUNSEPTIUM	118 Uuo 294 UNUNOCTIUM

To understand what this means in terms of an element's electron configuration, let's consider the Group 1A metals. If we write the electron configuration for the Group 1A metal from each row of the Periodic Table, we have:

row	element	electron configuration
2	Li	$1s^2 2s^1$
3	Na	$1s^2 2s^2 2p^6 3s^1$
4	K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
5	Rb	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$
6	Cs	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$
7	Fr	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^1$

Do you see any patterns? For Group 1A metals, it seems that element's row corresponds to the energy level of that element's valence electron. Lithium (Li), for instance, is found in row **2** of the periodic table, and its valence electron is found in the $n = 2$ energy level. Cesium (Cs) is found in row **6** of the periodic table, and its valence electron is found in the $n = 6$ energy level. Now consider the Group 2A metals:

row	element	electron configuration
2	Be	$1s^2 2s^2$
3	Mg	$1s^2 2s^2 2p^6 3s^2$
4	Ca	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
5	Sr	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$
6	Ba	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$
7	Ra	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2$

For Group 2A metals, the same rule applies. Magnesium (Mg) is found in row **3** of the periodic table, and its valence electrons are found in the $n = 3$ energy level. Similarly, radium (Ra) is found in row **7** of the periodic table, and its valence electrons are found in the $n = 7$ energy level.

Consider the electron configurations for the Group 3A elements in each row:

row	element	electron configuration
2	B	$1s^2 2s^2 2p^1$
3	Al	$1s^2 2s^2 2p^6 3s^2 3p^1$
4	Ga	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$
5	In	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^1$
6	Tl	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 5d^{10} 6p^1$

Even though the valence electrons in Group 3A elements are found in both **s** and **p** orbitals, it turns out that an element's row still corresponds to the energy level of that element's valence electrons. For example, gallium (Ga) is found in row **4** of the periodic table, and its valence electrons are found in the $n = 4$ energy level. Likewise, thallium (Tl) is found in row **6** of the periodic table, and its valence electrons are found in the $n = 6$ energy level.

Let's consider one last example by examining the electron configurations for the Group 7A elements:

row	element	electron configuration
F	B	$1s^2 2s^2 2p^5$
Cl	Al	$1s^2 2s^2 2p^6 3s^2 3p^5$
Br	Ga	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
I	In	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^5$
At	Tl	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 5d^{10} 6p^5$

Once again, it turns out that an element's period can be used to determine the energy level of that element's valence electrons. Chlorine (Cl), for instance, is found in row **3** of the periodic table, and its valence electrons are found in the $n = 3$ energy level. Similarly, iodine (I) is found in row **5** of the periodic table, and its valence electrons are found in the $n = 5$ energy level.

Note, however, that determining the energy level of an element's valence electrons by using the element's row number works only for Groups 1A – 8A. This rule *does not* apply to Groups 1B – 8B (also known as Groups 3 – 12) or to lanthanides and actinides. The elements in these groups behave differently and do not obey the same rules as the Group 1A – 8A elements in terms of valence electrons and valence electron energy levels.

You may have noticed that the first row has not been mentioned at all. In all of the examples considered, only the elements starting from the second row were considered. The first row in the periodic table is a “special row” for several reasons. To begin with, the first row only contains hydrogen and helium. Can you figure out why there are only two elements in the first row?

Recall that an element's period number corresponds to the energy level of that element's valence electrons. Therefore, the first row must contain elements with valence electrons in the $n = 1$ energy level. In the $n = 1$ energy level, there is only one orbital: the $1s$ orbital. Hydrogen (electron configuration $1s^1$) has one valence electron in the $1s$ orbital, and helium (electron configuration $1s^2$) has two valence electrons in the $1s$ orbital. Since it is impossible to fit more than two electrons into the $1s$ orbital, atoms with a total of three or more electrons must have valence electrons in an energy level with $n = 2$ or greater. Atoms with a total of three or more electrons do not belong in the first row of the periodic table.

The first row is also special because its elements have special properties. Hydrogen, for example, is not a metal like the rest of the Group 1A elements. Instead, hydrogen atoms react with each other and form hydrogen gas, H_2 . Hydrogen is different from the other Group 1A elements, so the chemical properties of hydrogen cannot be compared with the chemical properties of the other elements in Group 1A.

The second element in the first row, helium, is also a special atom. Helium is noticeably placed away from hydrogen at the far right-hand side of the periodic table. Helium's placement reflects its special chemical properties. While helium only has two valence electrons when the rest of the Group 8A elements have eight, helium is grouped with the 8A family because helium's chemical behavior is similar to the chemical behaviors of the other noble gases. This is because helium and the other noble gases all have a completed outer energy shell, even though the number of valence electrons is different.

This video describes the organization and family properties within the periodic table (**1b**, **1c**, **1f**): <http://www.youtube.com/watch?v=rD0maRGQhU8> (4:45).



Figure 1.2: ([Watch Youtube Video](http://www.ck12.org/flexbook/embed/view/371))
<http://www.ck12.org/flexbook/embed/view/371>

Lesson Summary

- The vertical columns on the periodic table are called groups or families because the elements in the column exhibit similar chemical behaviors.
- The groups on the periodic table also have similar electron configurations.
- The electron configuration in the outer energy level of the atoms in a family is the same, with the only difference in the electron configurations being the number of the outer energy level.
- All the members of a family of elements have the same number of valence electrons.
- The blocks of elements in the periodic table correspond to the sub-level in which the last electrons are being added.
- The horizontal rows on the periodic table are called periods.
- You can predict the energy level of an element's valence electrons using the element's period number in the periodic table.
- For the elements in Groups 1A, 2A, 3A, 4A, 5A, 6A, 7A, and 8A, the period in which an element is found corresponds to the energy level of that element's valence electrons.

Further Reading / Supplemental Links

This website provides a dynamic periodic table that also provides information about the orbitals and electron configurations of the elements.

- <http://www.dayah.com/periodic>
- A video on electron configuration and the periodic table.
 - <http://video.yahoo.com/watch/193832/925988>

This website reviews the different groups in the periodic table.

- <http://www.chemtutor.com/perich.htm>

This video is a ChemStudy film called “Transuranium Elements.” The film is somewhat dated but the information is accurate.

- <http://www.youtube.com/watch?v=SpQSEN1n4HE>

This video is a ChemStudy film called “Inert Gas Compounds.” The film is somewhat dated but the information is accurate.

- <http://www.youtube.com/watch?v=QASDy4bLQxs>

Review Questions

1. The horizontal rows in the periodic table are called _____.
2. The vertical columns in the periodic table are called _____.
3. Elements in a _____ have similar chemical properties.
 - (a) period
 - (b) family
 - (c) both a and b
 - (d) neither a nor b
4. Which of the following elements is in the same family as fluorine?
 - (a) silicon
 - (b) antimony
 - (c) iodine
 - (d) arsenic
 - (e) None of these.
5. Which of the following is a noble gas?
 - (a) polonium
 - (b) astatine
 - (c) krypton
 - (d) hydrogen
 - (e) francium
6. Name the period (horizontal row) and the sub-level (*s*, *p*, *d*, *f*) of the valence electrons for the following elements. Then indicate the number of valence electrons for that element.
 - (a) hydrogen

- (b) carbon
 - (c) sulfur
 - (d) barium
 - (e) iodine
7. Of the s-block, p-block, d-block, and f-block of elements, which block contains most of the nonmetals?
8. Why are there no transition elements in periods 1 and 2?
9. Why do many of the transition elements have similar chemistry?
10. Use the periodic table to determine the energy level and sub-level of the highest energy electrons in each of the following elements:
- (a) N
 - (b) Ca
 - (c) Rb
 - (d) P
 - (e) In
11. Decide whether each of the following statements is true or false.
- (a) Li has valence electrons in the $n = 1$ energy level.
 - (b) Si has valence electrons in the $n = 3$ energy level.
 - (c) Ga has valence electrons in the $n = 3$ energy level.
 - (d) Xe has valence electrons in the $n = 5$ energy level.
 - (e) P has valence electrons in the $n = 2$ energy level.

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