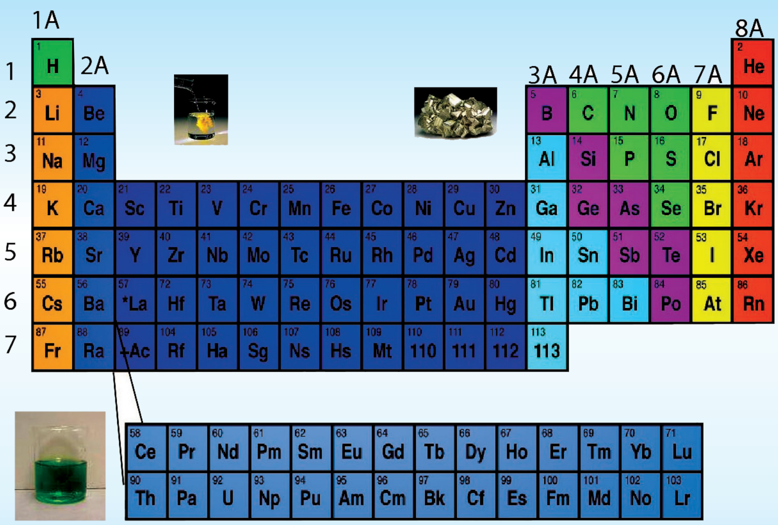
The Periodic Table



Elements, electron configurations, electron dot-structures, and bonding

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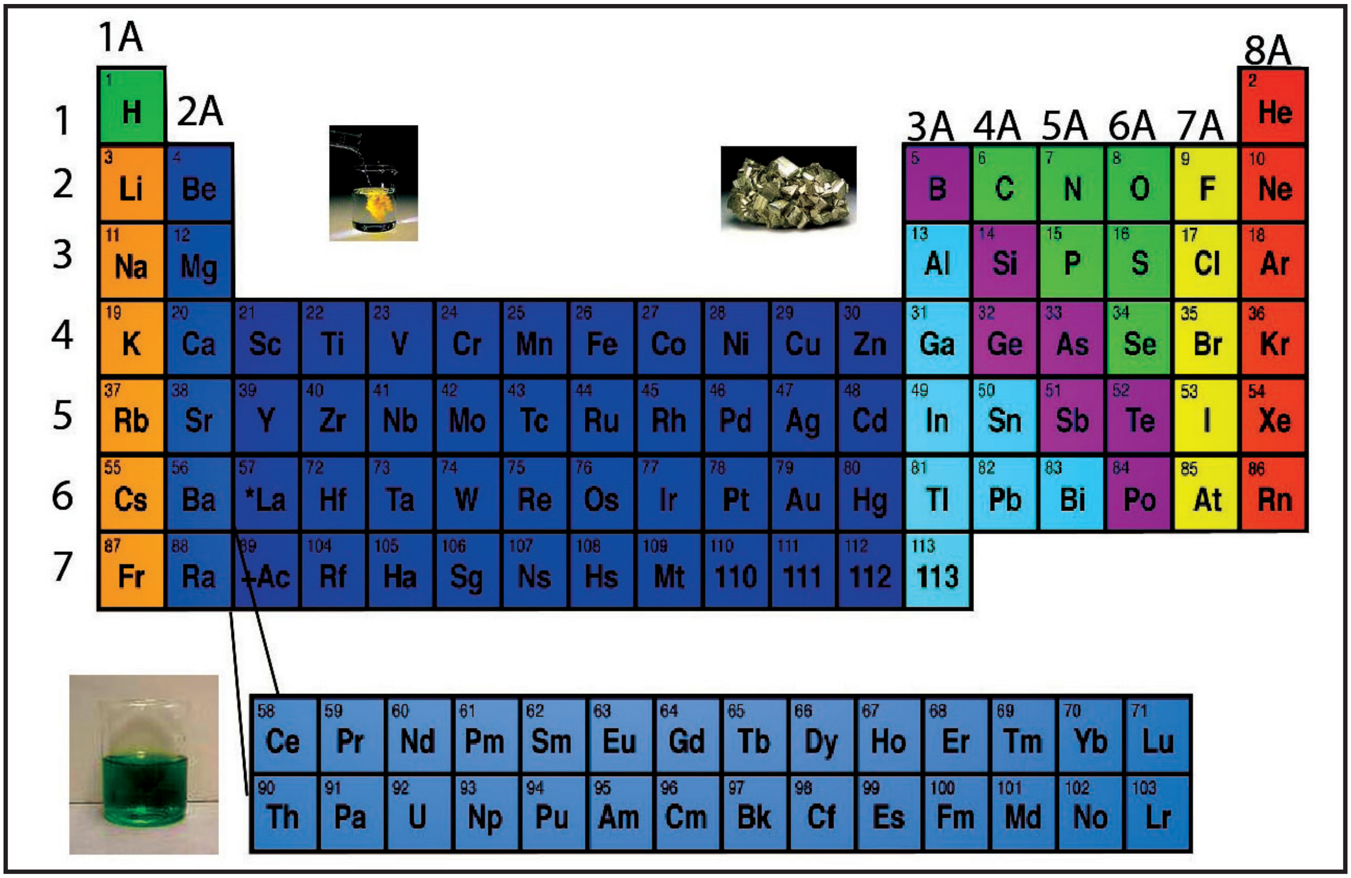
Contents

Development of the Modern Periodic Table

Throughout history man has been intrigued by elements.. Cultures found value as well as recognizing the importance of certain elements such as gold and silver. By the dawn of the industrial revolution chemists were overwhelmed with data on existing and newly discovered elements. Chemists soon realized the importance of being able to categorize elements based upon their properties. By 1864 a chemist by the name of John Newlands recognized that when elements were arranged by increasing atomic mass their properties repeated regularly with every eighth element.



By 1869 Lothar Meyer and Dimitri Mendeleev each demonstrated the relationship between an elements atomic mass and its properties. Mendeleev arranged elements with similar properties by increasing atomic mass into columns. This first periodic table had some flaws that were worked out by an English chemist.

Although Mendeleev is credited with creating the first periodic table it left something to be desired. Around the year 1913 the English chemist Henry Moseley, who had already discovered that an elements protons were equal in number to the elements atomic number, determined that if elements were arranged by increasing atomic number a pattern of repeating properties would arise. We call this the periodic law.

The Periodic Table

Elements are arranged on the modern periodic table by increasing atomic number. Each element has a specific atomic number which relates to the number of protons that are located within its nucleus. Hydrogen has one proton, therefore, its atomic number is one and it is the first element on the periodic table. Helium has two protons and, therefore, has an atomic number of two and is the second element on the periodic table.

The periodic table is divided into two larger groups of elements. Group (A) elements are the representative elements which have a very wide range of chemical and physical properties. Refer to figure 1. Group (B) elements are the transition metals.

Besides increasing atomic number elements are arranged into groups and periods.

Groups

If you look at a periodic table you will notice that elements are arranged into vertical columns. Within the vertical columns elements are arranged by increasing atomic number. For the group (A) elements the group number refers to the number of valence electrons in each elements valence shell.

The valence shell is the outer most energy level and we refer to the electrons in this level as valence electrons. For example group (1A) elements starting with hydrogen and ending with francium all have one valence electron. It does not matter how many electrons elements within the same group have, the group number refers only to the number of valence electrons located within the valence shell. (Refer to figure 2.) Valence electrons are responsible for the bonding of elements and thereby the elements chemical properties. Moving down through each group there is also a trend of increasing atomic radii. The radius of each element increases due to the increasing number of electrons with each succeeding element.

Periods

Elements are also arranged into seven periods by increasing atomic number. Period numbers refer to the energy level occupied by the elements valence electrons. For example sodium is located in period three because its one valence electron is located in the third energy level. This is also the case for all elements within period three. It does not matter how many electrons an element has. Its valence electrons are located in the energy level designated by the period for which it is located.

Moving from left to right across each period another trend arises. As the atomic number increases from left to right across each period atomic radius tends to decrease. This is due to an increasing positive charge within the nucleus.

Therefore, with the arrangement of elements, by increasing atomic number, scientists are able identify and classify substances by their properties.

Classification of Elements

Metals

Metals are located towards the left side of the periodic table from groups (1A) to group (5A) including the inner transition metals (group B elements). Metals tend be in a solid state at room temperature, except for mercury which is naturally found as a liquid, they are shiny, and good conductors of heat and electricity.

Other properties of metals include malleability which refers to a substances ability to be hammered into thin sheets without tearing. Examples of malleable metals are aluminum and tin. Each of these have many uses, however, you are most familiar with aluminum foil because it is commonly used to store left over food. Ductility refers to a substances ability to be drawn into the shape of a wire. Copper is commonly used as speaker wire as well as within electrical wires and cords because of its ability to be drawn into small strands, without breaking, and pass electrical currents. Strength is another common property of metals which refers to their ability to resist pressure and breakage. Within the periodic table metals are arranged into different classes based upon their properties.

Alkali Metals

Alkali metals are located in group 1A on the periodic table. These electrons are very reactive due to their one valence electron. Because alkali metals have only one valence electron they are more than eager to ionically bond with other elements in order to achieve a stable electron configuration. Due to their high reactivity alkali metals are not found existing naturally by themselves in nature and will explode if they come into contact with water.

Alkali metals exhibit the same properties as all other metals in that they are malleable, ductile, and good conductors of heat and electricity.

Alkaline Earth Metals

Alkaline earth metals are located within group 2A of the periodic table. Alkaline earth metals have an oxidation number of 2^+ which makes them very reactive, but not as reactive as group 1A elements. Because of their reactivity they are also eager to bond with other elements and are not found occurring naturally by them selves. Alkaline earth metals share the same properties of all metals in that they are ductile, malleable, and good conductors of heat and electricity.

Transition Metals

Transition metals and inner transition metals are located within group (B) on the periodic table. On some periodic tables they are arranged into groups 3B through 12B. Transition metals share the same properties as all other metals in that they are malleable, ductile, and good conductors of heat and electricity.

Transition and inner transition metals differ from other elements in the way their valence electrons are organized. Valence electrons are often located within different valence shells which are often times only partially filled. This explains why transition metals, within common periods, have different oxidation numbers.

Metalloids

Metalloids are a unique group of elements in that they are neither metals or nonmetals. Metalloids form a stair step pattern on the periodic table which includes boron, silicon, germanium, arsenic, antimony, tellurium, polonium, and astatine and share similar properties of both metals and nonmetals. The most common uses of metalloids are as semiconductors and storage cells for solar energy.

Nonmetals

Nonmetals are the opposite of metals. Where metals are generally solid, shiny, and great conductors of heat and electricity nonmetals are not. Nonmetals are typically in a gaseous state at room temperature. When they are found as naturally occurring solids they tend to be brittle and poses a dull luster. The only nonmetal that is found in a liquid state at room temperature is bromine which belongs to a special group of nonmetals called halogens.

Halogens

Halogens are found in group 7A and are highly reactive. Halogens are very reactive because they only need one electron to fill their valence shells. When halogens come into contact with other elements they tend to bond by removing valence electron from the other element. For example when chlorine, a halogen, comes into contact with sodium, an alkali metal that is very reactive because it strives to lose its one valence electron, the chlorine takes the valence electron away from sodium completely filling its valence shell. This ionic bond forms the ionic compound sodium chloride better known as table salt.

Noble Gases

Noble gases are found in group 8A of the periodic table. Noble gases are inert because they do not react or combine with any other elements. This non-reactivity is because noble gases have valence shells that are naturally completely filled.

Review

1. *How are elements arranged on the periodic table?*
2. *Describe the relationship between an element and its group number.*
3. *What is the significance between an element and its period number?*
4. *What are the three main categories of elements?*
5. *Compare and contrast the properties of metals, nonmetals, and metalloids.*
6. *Describe the group 1A elements in terms of their reactivity.*
7. *In which group are the halogens and why are they so reactive?*
8. *What is unique about the noble gases that make them different from the rest of the elements?*

Electron Configurations

The way in which electrons are configured determines an elements chemical properties. Using the periodic table it is easy to determine how electrons are configured. Remember that group numbers refer to the number of an elements valence electrons and period number refers to the energy level in which they are located.

Valence Electrons

Elements are arranged into groups and periods by increasing atomic number. Elements within the same group share similar chemical properties. Therefore, elements within the same group share similar chemical properties because they have the same number of valence electrons.

Within group 1A you will find hydrogen and the alkali metals lithium, sodium, potassium, rubidium, cesium, and francium. Because each of these group 1A elements has only one electron in its valence shell it readily combines with other elements in order to achieve a stable electron configuration containing eight valence electrons. This rule only applies the group (A) elements. Group (B) metals are a little different when it comes to electron configurations. Besides providing the number of valence electrons group numbers are also helpful in determining how to write electron dot structures which are used to illustrate the bonding of two or more elements.

There is also a connection between valence electrons and period number. The period number designates which energy level the valence electrons located within. Below are examples that show the relationship between period number and valence electrons:

Period1: Hydrogen $1s^1$
Period 2: Lithium $1s^2 2s^1$
Period 3: Sodium $1s^2 2s^2 2p^6 3s^2$

Notice that in each of the above examples the valence electrons are found in the energy level that corresponds to the period number in which each element is located.

s,p,d, and f Block Elements

The periodic table may be further divided into four blocks based upon the configuration of valence electrons. Have you ever looked at a periodic table and wondered why it has that odd shape? Its shape is due to elements being arranged into one of four distinct blocks, either s, p, d, or f. Block letters indicate the sublevel in which the valence electrons are located.

s-block elements consist of groups 1A and 2A including hydrogen and helium. Group 1A elements have partially filled valence shells with one electron in the s-orbital and an electron configuration ending in s^1. Group 2A also have partially filled valence shells consisting of two electrons in the s-orbital and an electron configuration of s^2

p-block elements consist of groups 3A through 8A. Within the valence shell once the s-orbital has been filled electrons will occupy the p-orbitals. Each energy level has a total of three p-orbitals that may hold two electrons each. Together the s- and p- blocks are made up of the representative elements also referred to as group A elements.

d-block elements are comprised of the transition metals. These elements are referred to as d-block elements because their valence electrons are located within the d-orbital of their highest energy level. The d-sublevel may be comprised of five orbitals each being able to hold two electrons, therefore, the d-sublevel may hold a total of ten electrons.

Inner transition metals make up the f-block elements. The f-sublevel consists of seven f-orbitals each being able to hold two electrons for a total fourteen.

Notice that in each block, s-, p-. d-, and f-, the total number of electrons matches the number of columns that span each block. For example the s-block elements span groups 1A and 2A and may hold a total of two electrons. p-block elements span groups 3A through 8A, 6 groups, and may hold a total of six electrons. The same holds true of both the d- and f-block elements as well. Now that we know how electrons are configured lets look at how to write electron configurations.

Writing e^- Configurations

The arrangement of electrons within a stable atom is called the electron configuration. All matter tends to seek stability with the lowest possible energy. Therefore, electrons are arranged in a formation requiring the lowest possible amount of energy to maintain itself. This low energy configuration is called the ground state. Electron configurations are based upon three rules–the Aufbau principle, the Pauli exclusion principle, and Hund's rule–each of which determine how electrons are arranged in any given atom.

Aufbau principle

As you recall our current model of the atom was developed over many years and went through many changes along the way. The Niels Bohr theorized that the atom has a nucleus, containing protons and neutrons, and that electrons orbit the nucleus within specific energy levels. This model helped explain some of the properties of elements, however, did not live completely up to the task. Later Erwin Shroedinger, based on scientific testing and evidence postulated the Quantum Mechanical model of the atom. This changed Bohr's two-dimensional atom to fit a more realistic three dimensional scale and, moreover, assigned electrons within each energy level to sublevels and orbitals.

Electrons are arranged in a specific sequence which is referred to as an atom's electron configuration. In nature all systems seek to be in a state of low energy because low energy systems are more stable than systems requiring higher amounts of energy. Atoms are no different. The electron configuration of any atom is such that electrons occupy specific energy orbitals within energy sublevels requiring the least amount of energy to maintain resulting in an electrically stable atom.

According to the aufbau principle electrons occupy the lowest energy orbital within a given energy sublevel. The diagram above illustrates the atomic orbitals within each energy sublevel. Using the aufbau diagram you can easily determine the correct order in which electrons are arranged. When placing electrons within their respective orbitals it is important to understand the following points.

9. All orbitals of the same type, within an energy sublevel, are of equal energy.

10. Orbitals are arranged in order of increasing energy; s, p, d, and f.

11. Orbitals of higher energy, within one sublevel, may overlap an orbital of lower energy within the next energy sublevel. For example the five orbitals related to an atom's 3d sublevel are of higher energy than the 4s orbital of the next sublevel.

Keep in mind that even though the aufbau principle provides a logical framework for the electron configuration of an atom, electrons are not built into any given atom one electron at a time. Furthermore, electrons do not exist as static points within two-dimensional space. To get a better picture of how electrons behave we need to add on to the aufbau principle.

The Pauli Exclusion Principle

Each electron within an atom not only occupies the lowest possible energy orbital it also rotates on its axis. The axial rotation of an electron will be in one direction or another. We represent this spin with arrows. An arrow pointing up (↑) represents the spin of an electron in one direction and an arrow pointing down (↓)represents the spin of an electron in the opposite direction. The **Pauli exclusion principle** states that a maximum of two electrons, with opposite spins, may occupy an energy orbital. Paired electrons are written as (↑↓)

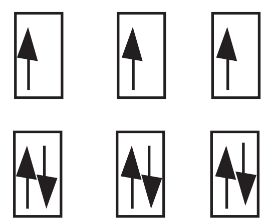


Figure 1. Shows how electrons with opposite spins occupy energy orbitals.

Hund's Rule

Recall that electrons are negatively charged subatomic particles. The fact that the paired electrons within an atomic orbital have the same, negative, charge may seem problematic because after all like charges repel each other. This repulsion problem is overcome by the intervention of Hund's rule. Hund's rule states that electrons with the same spin must occupy orbitals of the same energy. Once each orbital, of equal energy, has one electron then additional electrons with opposite spins may fill the same orbitals. Notice that in the diagrams below each orbital is initially occupied by one electron all of which are spinning in the same direction. Once all of the orbitals have one electron then additional electrons, with opposite spin, may enter and occupy the same orbital.

To sum up how the Aufbau principle, Hund's rule, and the Pauli exclusion principle affect an atom's electron configuration follow these simple steps.

1. Start with the lowest atomic orbital within the lowest energy sublevel.

2. Place one electron in each orbital, of equal energy, with the same spin.

3. Once all equal energy orbitals have one electron, with the

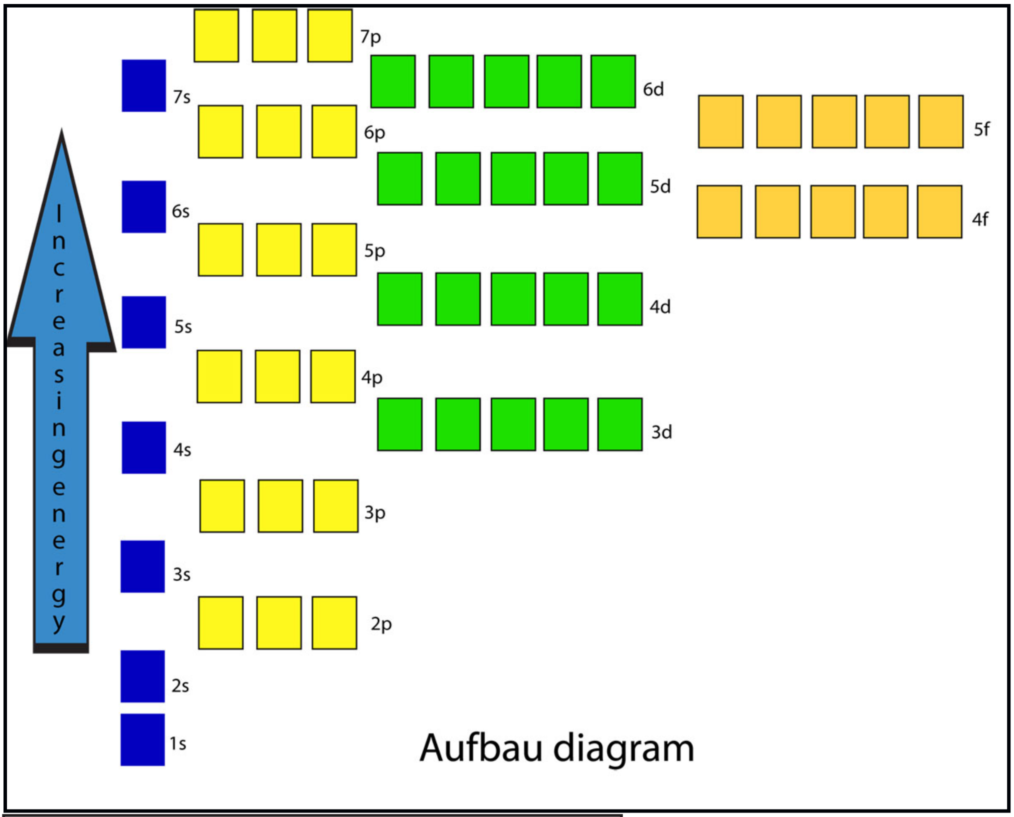


Figure 2. Aufbau diagram

same spin, go back place additional electrons, with opposite spins, in the same orbitals until each has a total of two or you run out of electrons.

Looking at the electron configurations for the elements in figure 3 you can probably guess that their notation can become quite long. In order to shorten things up a bit we may instead use

Element	Atomic Number	Atomic Orbital 1s, 2s, $2p_x$, $2p_y$, $2p_z$	Electron Configuration
H	1	↑	1s
C	6	↑↓ ↑↓ ↑↓	1s 2s 2p
N	7	↑↓ ↑↓ ↑↓ ↑	1s 2s 2p
Ne	10	↑↓ ↑↓ ↑↓ ↑↓	1s 2s 2p

Figure 3. Atomic Orbital and electron configurations.

the noble-gas configuration. For example the electron configuration notation for neon is $1s^2 2s^2 2p^6$. Using the noble-gas electron configurations we simply write out neon's electron configuration notation as (Ne). Using the noble-gas electron configuration notation for elements between neon and argon, period 3 elements, we simply write the symbol for neon (Ne) followed by the notation for the valence electron/s for that element. For example sodium's electron configuration is written as $1s^2 2s^2 2p^6 3s^1$. Writing the electron configuration for sodium using the its noble-gas electron configuration it is written as $(Ne)3s^1$. Magnesium is written as $(Ne)3s^2$. The electron configuration for all other elements in period 3 are written in the same manner up through argon who's noble-gas configuration is (Ar).

Review

1. What are valence electrons?

2. Describe the relationship between the number of valence electrons and group number for the group 1A elements.

3. Valence electrons of the representative elements occupy which block?

4. Write the electron configuration for the following elements:

a. Carbon

b. Nitrogen

c. Oxygen

d. Aluminum

Electron-dot Structures

An element's chemical properties are determined by its valence electrons. Valence electrons are the electrons occupying an elements outermost energy orbital which is also referred to as the valence shell. Being responsible for an elements chemical properties valence electrons form the chemical bonds between atoms. To visually represent the bonding of elements chemists use the atom's electron-dot configuration. **Electron-dot configurations** are represented by the element's symbol, which represents the atomic nucleus and lower level electrons, which is surrounded by dots representing the elements valence electrons. Figure 4 is a table illustrating the electron configuration and electron-dot structures for period two elements.

Bonding

Take a moment and look around you. What do you see? If you are sitting in a classroom you probably see desks, books, posters, windows, and several other objects. If you go outside you will probably notice trees, grass, asphalt, concrete, clouds, and several types of rocks. By now you know that all of these objects are made from chemicals. You also know that the chemicals making all of these objects are also called elements and that there are only about one-hundred and thirteen known elements. How do these one-hundred and thirteen elements make so many things. If you think about it there are so many different types of objects or substances that it would be impossible to try and count. So again the question is how do so few elements form so many different things?

Chemical bonds

The countless numbers of substances found on Earth and throughout the universe are the direct result of the ways in which elements join together. The answer to the question lies in an atom's electron structure as well as the intermolecular forces that hold atoms together. The forces that bind atoms together are called **chemical bonds.** Chemical bonds may form in a couple of different ways. Atoms may bond together due to the attraction between a positively charged nucleus and negatively charged electrons. Other bonds form from the attraction of positive and negatively charged ions. No matter what type of chemical bond forms atoms form chemical bonds in order to achieve stable electron configurations. An atom's valence shell may hold a total of eight valence electrons or an octet. An **octet** is the most stable electron configuration. The **octet rule** states that elements will share, gain, or lose electrons in order to achieve a stable electron configuration of eight valence electrons or an octet. The electron configuration of the noble gases naturally forms an octet resulting in their avoidance of other elements, or being inert. Therefore, once an element achieves an octet it assumes the electron configuration, also called the noble gas configuration, of the noble gas within its period. Keep in mind that by achieving a noble gas electron configuration an atom does not change into the noble gas it just forms the most stable electron configuration possible.

You recall that elements are arranged on the periodic table into groups based upon the number of valence electrons. Elements within the same group, no matter how many electrons they have, posses the same number of valence electrons in their outermost atomic orbitals. The type of chemical bond that an atom will form is dependent upon its valence electrons.

Ionic Bonds

Think back to the last time you ate french fires or anything that needed salt. Do you recall picking up the salt shaker and dispensing the salt onto your food? Table salt is made up of sodium and chlorine. In the formation of table salt a sodium (Na) atom transfers one electron to a chlorine (Cl) atom. By transferring it's one valence electron the sodium (Na) atom gains an overall positive charge due to it now having one more proton than electrons. When an atom transfers an electron, resulting in a greater number of protons than electrons, it becomes a positively charged ion called a **cation.** Conversely, when the chlorine (Cl) atom gains the electron in its valence shell it develops an overall negative charge due to it now having more electrons than protons. When at atom gains an electron, resulting in a greater number of electrons than protons it becomes a negatively charged ion called an **anion**. Therefore **ionic bonds** form as a result of an electrostatic attraction between oppositely charged particles when one atom transfers its valence electrons to another. Elements joined together by ionic bonds are called **ionic compounds** most of which occur between metals and form salts. Sometimes ionic bonds occur between metals and oxygen, a nonmetal, and form oxides such as when a bicycle is left outside for too long. When the iron in the bicycle is over exposed to oxygen, typically in water, it forms iron oxide more commonly known as rust.

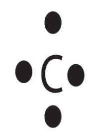
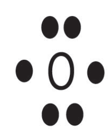
Element	Atomic Number	Electron Configuration	Dot Structure
Li	3	$1s^2 2s^1$	
Be	4	$1s^2 2s^2$	
B	5	$1s^2 2s^2 2p^1$	
C	6	$1s^2 2s^2 2p^2$	
N	7	$1s^2 2s^2 2p^3$	
O	8	$1s^2 2s^2 2p^4$	
F	9	$1s^2 2s^2 2p^5$	
Ne	10	$1s^2 2s^2 2p^6$	

Figure 4. Electron configurations and electron dot structures.

Covalent Bonds

You have learned so far that elements bond in order to gain eight valence electrons in their outer most energy orbitals. The **octet rule** states that elements will share, gain, or lose electrons in order to achieve a stable electron configuration of eight valence electrons or an octet. You learned in the last section that ionic compounds form from the transfer of valence electrons from one atom to another. What about when chemical bonds form between atoms without a transfer of electrons?

In order to achieve a stable electron configuration atoms will sometimes share their valence electrons forming a **covalent bond**. By sharing valence electrons both atoms fill their valence shells achieving stable electrons configurations. Some elements occur naturally as covalent compounds. (O_2) oxygen, (N_2) nitrogen, and (H_2) hydrogen are three examples of elements that naturally occur in pairs and are called **diatomic molecules. Covalent compounds** form as a result of covalent bonds of which there are several types.

When two atoms bond together by sharing a single pair of electrons they form a single covalent bond. For example the diatomic molecule (H_2) forms by each hydrogen atom sharing its one valence electron with the other. The figure below illustrates how water forms from the single covalent bonds between hydrogen and oxygen.

Sometimes atoms share multiple pairs of electrons in order to achieve stable electron configurations. The periodic table is useful in determining which elements are likely to form multiple covalent bonds. Recall that the group number refers to the number of valence electrons in an atom. Elements within group 1A have only one valence electron which also means that they need to gain seven more to form an octet. Gaining seven valence electrons is monumentally more difficult to achieve than gaining one making the group 1A elements over eager in their attempt to transfer it to another element.

Double covalent bonds form when an atom shares two pairs of valence electrons and triple covalent bonds form when an atom shares three. Figure 5 illustrates double and triple covalent bonds.

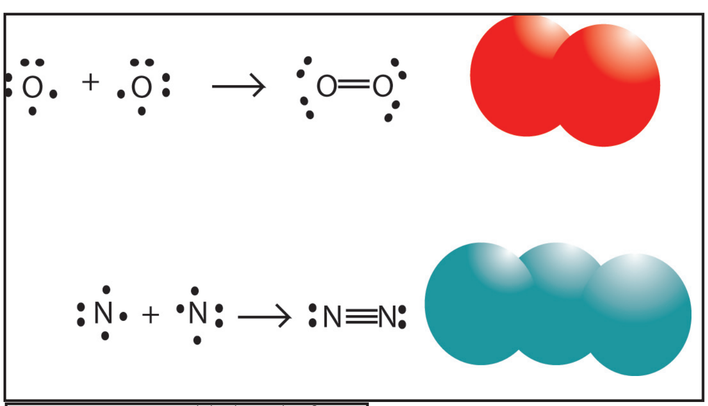


Figure 5. Double bond of oxygen and triple bond of nitrogen

Review

1. *What is a chemical bond?*
2. *When a bond forms from the transfer of valence electrons from one element to another an _____ bond forms.*
3. *Covalent bonds form as a result of _____ valence electrons.*
4. *Why do elements bond to form compounds?*
5. *Explain the octet rule.*
6. *Explain why group 8A elements do not interact to form bonds with other elements.*
7. *Draw the electron dot structures for the following elements:*

a. H
b. Ne
c. O
d. Fr
e. Si

Key Terms	Use this page as a guide to taking notes on the contents of this booklet. Notes
	Development of the periodic table: Group numbers: Periods: Properties of metals: Properties of nonmetals: Metalloids: Halogens: Noble gases: Electron Configurations: Valence Electrons: s, p, d, and f block elements: Writing Electron Configurations: Aufbau principle: Hund's rule: Pauli exclusion principle: Dot structures: Bonding: