Name: _______________________________________ Date: _________________________

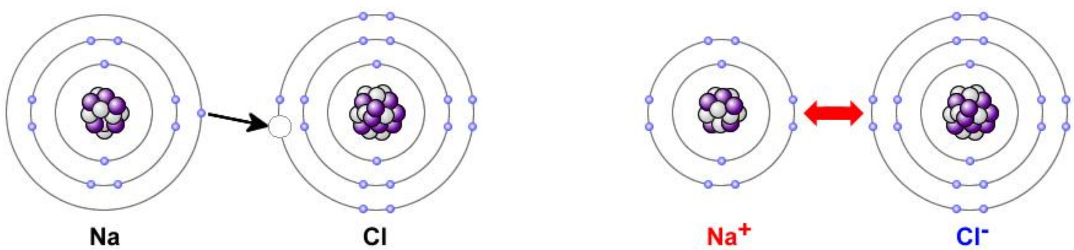
Guided Learning: Chemical Bonding (Part 2)

[Note to students: Have a copy of the periodic table handy as you complete this activity.]

Activity D: Ionic bonds
When atoms form ions, they gain or lose electrons until they reach a stable configuration with a full valence shell. The first energy level is stable with a total of two electrons, while the second and third energy levels are considered full when they have eight valence electrons. (Even though the third energy level isn't full with eight electrons, it is very stable in that configuration.) The stability of many atoms when they have eight valence electrons is known as the **octet rule**.

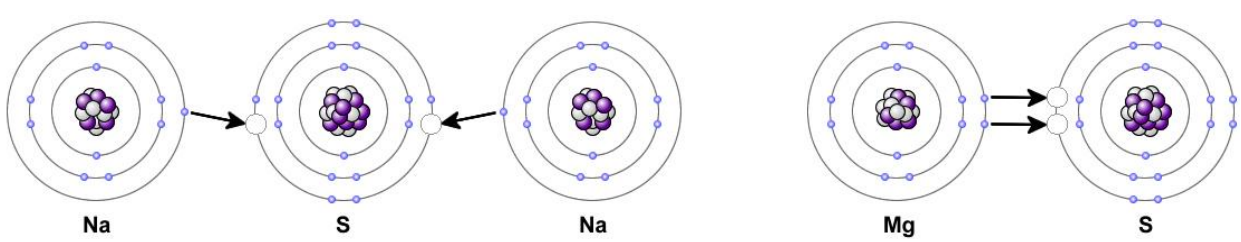
Just as negatively charged electrons are attracted to the positively charged nuclei, negative ions are attracted to positive ions by the electromagnetic force. The resulting chemical bonds that form between ions are called **ionic bonds**.

Consider the meeting of a neutral sodium atom (Na) and a neutral chlorine atom (Cl). The sodium atom has one valence electron in its third energy level. The chlorine atom has seven valence electrons in the third energy level. The sodium atom has low ionization energy, so the single valence electron is easily jostled free. The chlorine atom has high electron affinity and attracts the loose electron (below left).



After the electron is transferred (above right), both atoms have stable valence shells with 8 electrons. The sodium atom is now a positively charged ion while the chlorine atom is a negatively charged ion. The ions are attracted to one another and form an ionic bond. The result is the chemical compound sodium chloride (NaCl), also known as table salt.

The number of valence electrons determines how many bonds an atom can form. Sulfur, for example, has six valence electrons. It can form bonds with two sodium atoms (below left) or with one magnesium atom, which has two valence electrons (below right). The resulting compounds have the formulas Na_2S and MgS, respectively.



1. In your own words, summarize how ionic bonds form. ______________________________

____

____

2. Nitrogen (atomic number 7) forms ionic bonds with lithium (atomic number 3). Draw the electrons for each atom in the diagrams below:

Nitrogen **Lithium**

   A. How many valence electrons does nitrogen have? ___________

   B. How many valence electrons does nitrogen need for a full *p* subshell? ___________

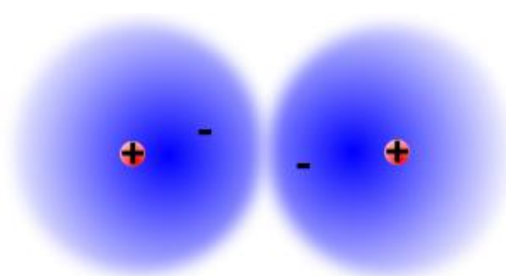
   C. How many valence electrons does lithium have? ___________

   D. How many lithium ions will bond with a nitrogen atom? ___________

   E. What is the chemical formula of lithium nitride? ___________

Activity E: Covalent bonds

In an ionic bond, a valence electron is transferred from one atom to another. Valence electrons can also be shared between two atoms to form a **covalent bond**.



To see how a covalent bond forms, imagine two hydrogen atoms approaching one another. As they move closer together, each electron is attracted to the positively charged nucleus of the other atom. As a result the electrons spend more time in the region between the two atoms than they otherwise would. The attraction between each electron and the opposite nucleus bonds the two atoms together to form a **molecule** of hydrogen gas, H_2.

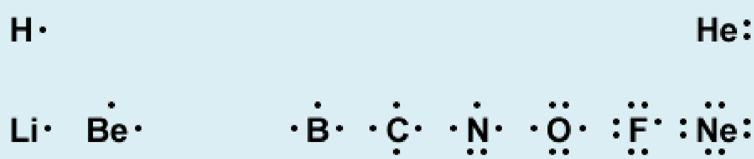
As in ionic bonds, nonmetal atoms (with the exception of hydrogen) are most stable in covalent bonds when their outer *p* subshell is full. Atoms will share pairs of electrons until they achieve a stable configuration with a full *p* subshell.

For example, fluorine atoms have seven valence electrons. Two fluorine atoms can share a single pair of electrons to create a stable molecule of fluorine gas, F_2. Oxygen atoms have six valence electrons. Therefore, two oxygen atoms must share two pairs of electrons to reach a stable octet of valence electrons in the molecule O_2.

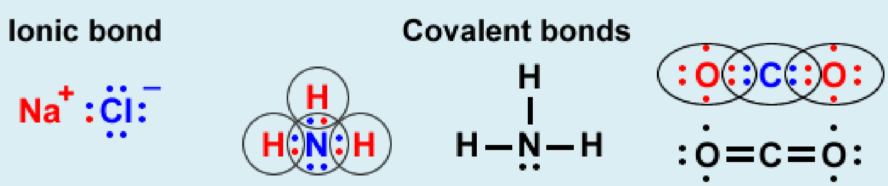
In general, the number of valence electrons determines the **valence**, or number of bonds, an atom can form. For nonmetals other than hydrogen, the valence is equal to the difference between the number of valence electrons and eight. For metals, the valence is equal to the number of valence electrons.

Electron dot diagrams

Because the inner electrons do not take part in chemical bonds, they can generally be ignored. Chemists use **electron dot diagrams** as a shorthand method to represent the valence electrons for a given element. Here are some examples:



Electron dot diagrams can be used to show ionic or covalent bonds. In an ionic bond, use a different color or symbol to represent the transferred electrons, and label the charge of each ion. In a covalent bond, circle each atom that has a stable electron configuration (two electrons for hydrogen, eight electrons for other elements). In covalent molecules, pairs of shared electrons can be shown with a dash, as illustrated in the two right diagrams. In carbon dioxide, carbon and oxygen form double bonds consisting of two pairs of shared electrons.



1. In your own words, summarize how covalent bonds form. ______________________________

____

____

2. Use electron dot diagrams to represent the following molecules: H_2, F_2, H_2O, and CH_4.

H H   F F   H O H   H C H (with H above and below C)

3. Use electron dot diagrams to represent the following ionic compounds: LiCl, BeF_2, and Na_2O. Use a different color or symbol for the transferred electrons, and indicate the charge of each ion.

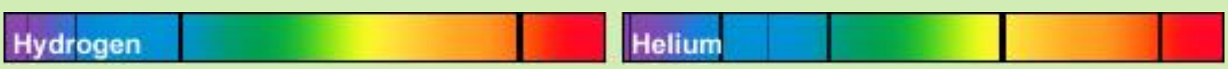
Li Cl F Be F Na O Na

Molecular structure and gases

In an “ideal” gas, particles are modeled as perfectly hard spheres that constantly collide with one another. This is close to the truth for gases composed of noble gas elements such as helium, neon, and argon because the these gases are *monatomic*—they are composed of individual atoms. When energy is added to a noble gas, all of the energy is expressed by the kinetic energy of the atoms. As a result, noble gases have very low specific heat capacities—they are relatively easy to heat up.

Most gases are made of molecules, however. When molecules are heated, some of the energy goes into their motion, but some of the energy causes them to rotate, twist, or vibrate. The more complex the molecule, the greater the number of ways the energy can be absorbed without increasing the temperature of the gas. This makes these gases harder to heat up. For example, argon (Ar) and fluorine (F_2) have nearly identical densities. However, the specific heat of fluorine (31.3 J/mol·°C) is much greater than the specific heat capacity of argon (20.8 J/mol·°C).

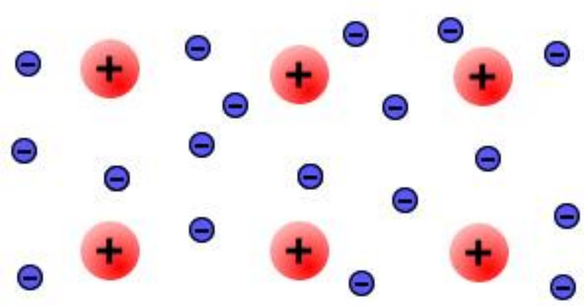
Molecular structure also affects the *absorption spectra* of gases, or the wavelengths of light that the gases can absorb. Simpler gases such as hydrogen and helium can only absorb a few different wavelengths of light (their absorption spectra are shown below).



Gases with more complex molecules—e.g., carbon dioxide, water vapor, and methane—can absorb a much greater variety of wavelengths. This explains why these gases tend to act as “greenhouse gases,” absorbing radiation emitted from Earth’s surface.

Activity F: Metallic bonds

A third type of chemical bond occurs between metal atoms. When metal atoms interact, their loosely-held valence electrons form a freely flowing “sea” that binds positively charged nuclei together in **metallic bonds**, shown at right. The fact that valence electrons can flow easily in metallic bonds accounts for many properties of metals such as **luster** (shininess), **malleability** (the ability to be flattened), **ductility** (the ability to be stretched into a wire), and **conductivity** (the ability to transmit heat or electrical current).



In metallic bonds, positively charged ions are surrounded by freely moving electrons.

For example, suppose a metal wire is connected to a positive charge at one end and a negative charge at the other end. The valence electrons would be attracted to the positive charge and will drift toward that end of the wire. This causes a current of electricity to flow in the wire. The mobility of electrons also allows metals to conduct heat from one place to another.
When a metal is pushed or pulled, the positive ions tend to slide over one another but still remain in contact and bonded together. This causes the metal to deform rather than break or shatter, allowing the shape of metals to be changed. Thus, metals can be beaten into sheets, drawn into wires, or changed in many other ways.

The luster of metals is related to the fact that valence electrons in metals are not locked into rigid chemical bonds. When photons of light strike the surface of a metal, the valence electrons absorb these photons and move into higher energy levels. They remain in this excited state for a few moments, then move down to a lower energy level while emitting photons. These photons are emitted in random directions, which is why metals do not act as perfect mirrors. The emission of photons gives metals their characteristic shiny appearance.

?

1. What property of metals allows them to form metallic bonds? ___________________________

   ____

2. How does the mobility of valence electrons in metals relate to each of these properties?

   A. Conductivity: ____

      ____

   B. Ductility/malleability: ____

      ____

   C. Luster: ____

      ____

3. When iron rusts, iron atoms form ionic bonds with oxygen atoms to form iron oxide, Fe_2O_3. How do you think that affects the conductivity, malleability, and luster of iron? Explain.

   ____

   ____

   ____

   ____

   ____

Activity G: Reactivity

While chemical bonds are very stable, some are more stable than others. When molecules or atoms collide, chemical bonds may break and new chemical bonds may form. When new substances are formed, a **chemical reaction** has taken place.

As you have learned, elements vary in their tendency to give up or gain electrons. This variability affects the **reactivity** of elements, or their ability to take part in chemical reactions. Chemists have constructed an informal list called the **reactivity series**, shown at right. The most reactive elements are at the top of the list, while the least reactive are at the bottom.

Reactivity depends on ionization energy as well as on the number of valence electrons. The most reactive elements are the alkali metals (group 1A). These elements all have just one valence electron to lose when they form ions. Within the group, the largest elements (francium, cesium, and rubidium) have the greatest distance between the valence electron and the nucleus and therefore have the smallest ionization energy and the greatest reactivity within the group.

Element	Group	Row	Reactivity
Rubidium (Rb)	1A	5	Reacts with water and acids
Potassium (K)	1A	4	
Sodium (Na)	1A	3	
Lithium (Li)	1A	2	
Strontium (Sr)	2A	5	
Calcium (Ca)	2A	4	
Magnesium (Mg)	2A	3	Reacts with acids but not with water
Beryllium (Be)	2A	2	
Aluminum (Al)	3A	3	
Zinc (Zn)	Trans.	4	
Chromium (Cr)	Trans.	4	
Iron (Fe)	Trans.	4	
Cobalt (Co)	Trans.	4	
Nickel (Ni)	Trans.	4	
Tin (Sn)	3A	5	
Lead (Pb)	3A	6	
Copper (Cu)	Trans.	4	May react with very strong acids
Silver (Ag)	Trans.	5	
Mercury (Hg)	Trans.	6	
Gold (Au)	Trans.	6	
Platinum (Pt)	Trans.	6	

The alkali earth metals (group 2A) all have two valence electrons and are the second most reactive group of elements. As with the alkali metals, the largest atoms in this group are most reactive, with reactivity decreasing as you move up in the periodic table. Reactivity trends are less regular in the transition metals and other groups.

Elements with greater reactivity will react more vigorously and produce more heat in exothermic reactions than less reactive elements. More reactive elements will also displace less reactive elements in single replacement reactions. For example, iron (Fe), which is more reactive than copper (Cu), will replace copper in copper sulfate in the following single replacement reaction:

$$Fe + CuSO_4 \rightarrow Cu + FeSO_4$$

Copper, on the other hand, will replace silver in silver nitrate because copper is more reactive than silver:

$$Cu + 2AgNO_3 \rightarrow 2Ag + Cu(NO_3)_2$$

While the reactivity series is not a perfect predictor of the behavior of metals in reactions, it is correct in most cases.

1. What two factors are the most important in determining the reactivity of a metal? __________

____

2. Based on the reactivity chart, which of the following single replacement reactions will occur? (Circle your choices.)

$Na + RbCl \rightarrow Rb + NaCl$

$2Li + MgSO_4 \rightarrow Mg + Li_2SO_4$

$2Fe + Al_2O_3 \rightarrow 2Al + Fe_2O_3$

$Zn + 2AgNO_3 \rightarrow 2Ag + Zn(NO_3)_2$

$Cu + FeSO_4 \rightarrow Fe + CuSO_4$

$Mg + 2LiNO_3 \rightarrow 2Li + Mg(NO_3)_2$

$2Fe + 3Pb(NO_3)_2 \rightarrow 3Pb + 2Fe(NO_3)_3$

$2Al + 3CuCl_2 \rightarrow 3Cu + 2AlCl_3$

3. The elements cesium, francium, barium, and radium were not included on the reactivity series table on the previous page. Based on their positions in the periodic table, where would each of these elements fit on the chart?

____

____

____

4. **Wrap-up question:** At the very beginning of this two-part Guided Learning activity, you were asked a question about rubbing balloons on your head. What does rubbing balloons on your head have to do with the structure of the atom, ions, chemical bonds, and reactivity? Give examples to support your answer. Continue your answer on extra sheets if necessary.

____

____

____

____

____

____