Name: ______________________________________ Date: ________________________

Guided Learning: Chemical Bonding (Part 1)

Learning goals
After completing this activity, you will be able to …

- Describe the structure of atoms and the forces that hold them together.
- Determine electron configurations of elements.
- Relate the formation of ions to electron affinity and ionization energy.
- Define and describe ionic, covalent, and metallic bonds.
- Use a reactivity series chart to predict whether a given reaction will take place.

Vocabulary: chemical bond, chemical reaction, conductivity, covalent bond, ductility, electron affinity, electron configuration, electron dot diagram, electrostatic forces, energy level, ion, ionic bond, ionization energy, luster, malleability, metal, metallic bonds, metalloid, molecule, noble gases, nonmetal, octet rule, orbital, Pauli exclusion principle, periodic table, quantum number, reactivity, reactivity series, spin, strong nuclear force, subshell, valence, valence electron

Warm-up questions

1. You rub a balloon on your hair and it sticks to your head. Why does this happen?

____

____

2. Circle the situation(s) that you think will produce an attractive force. Explain your choice(s).

    A. Two positively charged balloons
    B. Two negatively charged balloons
    C. A positively charged balloon and a negatively charged balloon
    D. A positively charged balloon and a neutral wall

    Explain: ____

    ____

3. Why do you think some atoms form **chemical bonds**? _____________________________

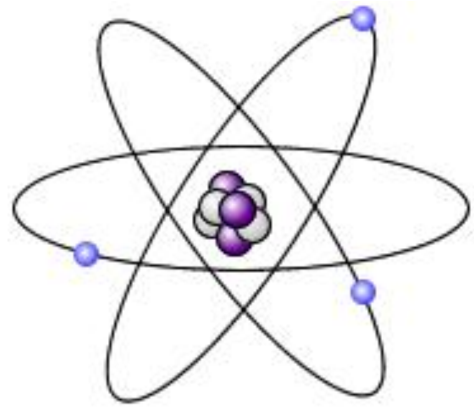
____

____

____

Activity A: Atomic structure

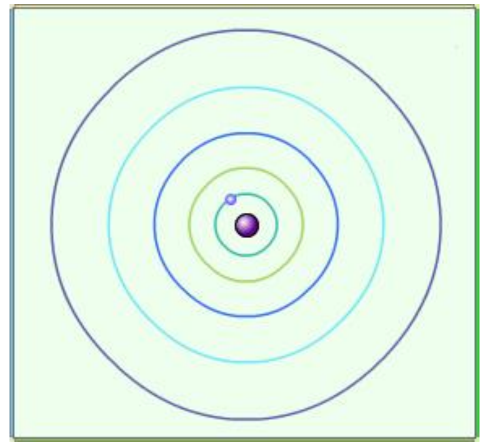
Atoms consist of arrangements of three particles: positively charged protons, neutral neutrons, and negatively charged electrons. The massive protons and neutrons are concentrated in the nucleus of the atom, while the tiny electrons exist outside of the nucleus. Conveniently, the electric charge of an electron is exactly opposite the charge of a proton. Therefore, an atom with an equal number of protons and electrons is neutral. A traditional "planetary" atomic model is shown at right.



The planetary atom: Protons are purple, neutrons are white, and electrons are blue.

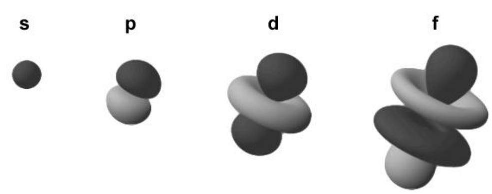
Atoms are held together by the interplay of two forces. The negatively charged electrons are held to the positively charged nucleus by **electrostatic forces**. Inside the nucleus, the protons and neutrons are bound together by the powerful **strong nuclear force**.

Although the planetary model was successful, physicists wondered what prevented the electrons from gradually spiraling into the nucleus. This problem was solved in 1913 by Niels Bohr, who hypothesized the electrons could only have certain specific energies. For example, in a hydrogen atom, the electron could only be located at specific distances from the nucleus, as shown at right. Because each orbit was associated with a certain energy, Bohr dubbed these orbits **energy levels**. They are also known as "shells."



The energy levels of hydrogen

In the years that followed, the model of the atom became more complex. For example, the electrons sometimes acted as particles and sometimes acted as waves. The shapes of some orbits were not perfect circles but strange donut and barbell shapes that indicated probability distributions rather than specific electron locations. Some examples of these probability distributions are shown in the bottom diagram at right.



Orbital shapes

When learning about atomic bonds, however, it is still quite useful to consider Bohr's model and assume electrons are particles that make circular orbits around the nucleus.

1. What might happen to an atomic nucleus if the strong force did not exist? _______________

____

____

2. Suppose protons have a charge of 1+ and electrons have a charge of 1-. What is the total charge of an atom with 23 protons and 28 electrons? ______________________________

3. **Looking ahead** Suppose an atom has electrons in several different energy levels. Which electrons do you think would be held most tightly, the electrons nearest the nucleus or the electrons farthest away? Explain your answer.

____

____

Activity B: Electron configurations

The arrangements of electrons around nuclei, or **electron configurations**, are complex. The state of each electron in an atom is described by four **quantum numbers**.

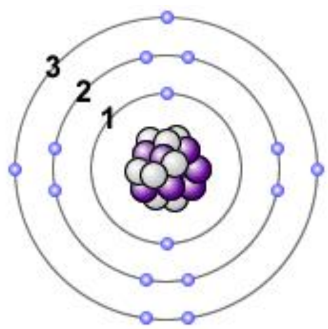
The first quantum number describes the energy level, or shell, an electron is in. The second quantum number describes the **subshell** of the electron. Subshells are given the names *s*, *p*, *d*, and *f*. The number of subshells a shell can hold is equal to the number of the shell. So shell 1 has an *s* subshell; shell 2 has *s* and *p* subshells; shell 3 has *s*, *p*, and *d* subshells; and shell 4 has *s*, *p*, *d*, and *f* subshells.

The third quantum number describes the **orbital** an electron occupies. Each orbital can hold two electrons. There is one orbital in the *s* subshell, three in the *p* subshell, five in the *d* subshell, and seven in the *f* subshell.

The fourth quantum number describes the **spin** of the electron, "up" or "down." Electrons do not actually spin—as far as we know—but they do exert a tiny magnetic field as if they *did* spin, so it is convenient to describe them this way. Up and down arrows are often used to represent electrons in their orbitals.

The **Pauli exclusion principle** states that no two electrons in an atom can be in precisely the same quantum state. Therefore, if two electrons occupy the same shell, subshell, and orbital, they must have opposite spins.

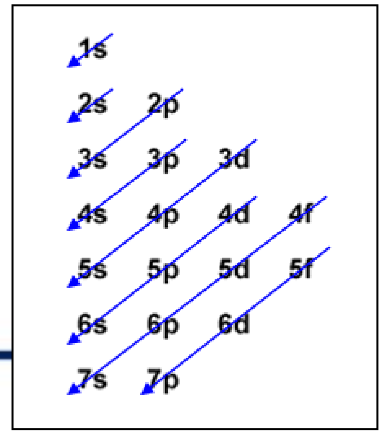
For example, consider neutral Phosphorus, which has an atomic number of 15. The 1*s* subshell is filled first. Because it is an *s* subshell, it has one orbital, so it contains two electrons. Next the 2*s* and 2*p* subshells are filled. The 2*s* subshell contains two electrons, while the 2*p* subshell contains six electrons, for a total of eight electrons in shell 2 and 10 electrons in shells 1 and 2. Next comes 3*s* (2 electrons) and 3*p*, which contains the remaining three electrons. In shorthand, the electron configuration of phosphorus is written $1s^2 2s^2 2p^6 3s^2 3p^3$. [Note: In the simplified diagram at right, each shell is shown as a single circle. Subshells are not marked.]



Phosphorus

The five electrons in the third shell are located farthest from the nucleus and are called **valence electrons**. Valence electrons are extremely important in chemistry because they participate in forming chemical bonds.

The remaining subshells are filled in a peculiar order that arises from the fact that there is energy overlap between shells. After 3*p* is filled, the next subshell to fill is 4*s* rather than 3*d*. This is followed by 3*d*, 4*p*, and 5*s*. Next in line are 4*d*, 5*p*, and 6*s*. The complete order of subshell filling is shown in the diagram at right. The pattern is known as the "diagonal rule." You can learn more about electron configurations with the *Electron Configurations* Gizmo™.



1. Write the electron configurations of each of the following elements. (Atomic numbers in parentheses.)

   Hydrogen (1) _______ Lithium (3) ____________ Carbon (6) ____________________

   Fluorine (9) _______________________ Sodium (11) _______________________

   Aluminum (13) _______________________ Chlorine (17) _______________________

   Challenge: Tin (50) ____

2. Fluorine and chlorine are found in the same column of the **periodic table**.

   A. How many valence electrons does fluorine have? _____ Chlorine? _____

   B. Why do you think fluorine and chlorine are found in the same column of the periodic table? ____

Activity C: Electron affinity, ionization energy, and ions
As you have learned, negatively charged electrons are held to the positively charged nucleus by electromagnetic forces. This force increases with the amount of charge and decreases with distance. Electrons in shells that are closer to the nucleus are therefore held much more tightly than the valence electrons that are farthest from the nucleus.

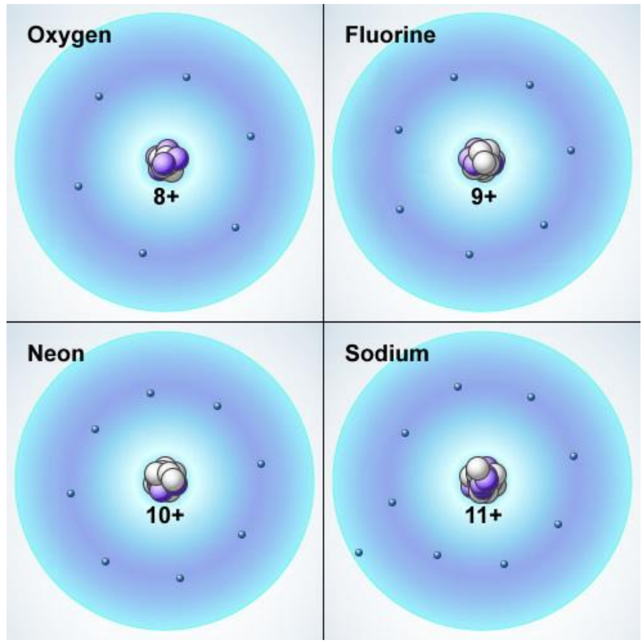
Predict: How do you think the attraction of electrons to the nucleus will change as a subshell is filled up? Explain your answer.

____

____

As a valence shell is filled, the total force binding these electrons to the nucleus increases, and the radius of the atom actually decreases. As a result, valence electrons are most tightly bound to the nucleus when the shell is nearly full. When a new shell is begun, however, its radius is greater than the radius of the previous shell, so the attraction between the valence electron and the nucleus is much less.

In the image at right, the attraction between valence electrons and the nucleus increases from oxygen to neon because there are more electrons in the 2*p* subshell and more protons in the nucleus. This causes the valence electrons to be pulled closer to the nucleus.



When an electron is added to neon, however, it occupies a new subshell, 3*s*, which is located significantly farther from the nucleus. This electron is held less tightly than the others. As a result, the radius of sodium (186 pm) is much greater than the radius of neon (38 pm).

Quick check: Which atom has a smaller radius? (Circle one)   Oxygen   Fluorine

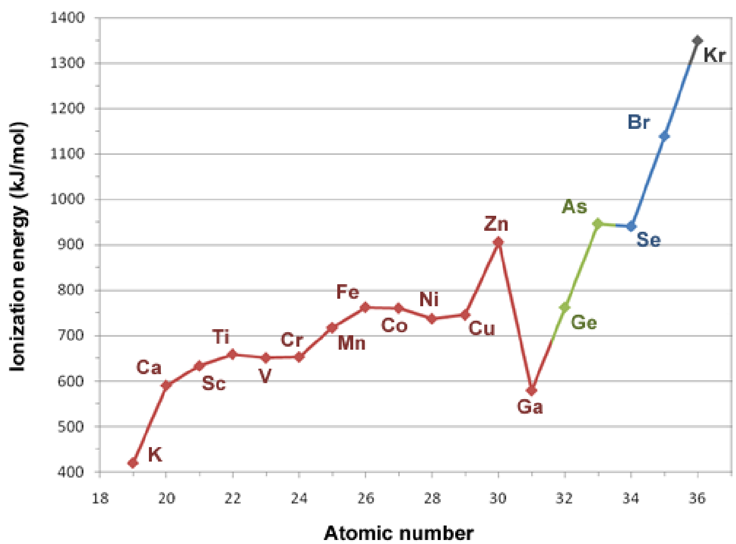
Explain: ____

____

The tendency of atoms to attract electrons is called the **electron affinity** of the atom. Electrons with nearly full valence shells tend to have a much greater electron affinity than atoms with valence subshells with only a few electrons. In addition, electrons in the *p* subshells tend to be held much more tightly than electrons in either the *s* or *d* subshells.

As atoms move and interact with other atoms, they tend to lose electrons or pick up extra electrons. When the number of electrons an atom holds is different from the number of protons, the atom becomes a charged **ion**. Positive ions (cations) have more protons than electrons, while negative ions (anions) have more electrons than protons. The **ionization energy** of an atom is the energy required to remove an electron.

The graph at right shows ionization energies for one row of the periodic table. Elements on the left side of the periodic table have low ionization energies and lose their valence electrons easily. These elements form positively charged ions and are called **metals**, shown in red on the graph. Elements on the right side of the periodic table have high electron affinities and accept electrons. These elements are **nonmetals**, shown in blue. A third group, called **metalloids** (green), exhibits intermediate characteristics.



The group on the far right of the periodic table, called the **noble gases**, consists of elements that have full *p* subshells and eight valence electrons. (The first noble gas, Helium, has two valence electrons.) Like nonmetals, these elements do not give up electrons easily. However, these elements do not tend to gain electrons either because their valence energy level is full. Therefore, the noble gases do not form ions and generally do not form chemical bonds.

The charge of an ion depends on the number of electrons an atom gains or loses. For example, neutral lithium has three protons and three electrons with an electron configuration of $1s^22s^1$. Lithium has one valence electron, located in the 2*s* subshell. This electron is easily lost because it is in the *s* subshell. When lithium looses this electron, it has three protons and two electrons, giving it an overall charge of 1+. The symbol for this ion is Li^{1+} or Li^+.

Neutral oxygen has eight protons and eight electrons with an electron configuration of $1s^22s^22p^4$. Oxygen has six valence electrons. Because the 2*s* subshell is full and the 2*p* subshell is nearly full, oxygen has a strong electron affinity and is much more likely to gain electrons than lose electrons. Oxygen can gain two more electrons before filling it's the *p* subshell, resulting in an ion with eight protons, ten electrons, and an overall charge of 2^-. This oxygen ion is written O^{2-}.

The columns of the periodic table contain groups of elements with the same number of valence electrons. Group 1 (hydrogen, lithium, sodium, potassium, etc.) are elements with one valence electron. Group 2 (beryllium, magnesium, calcium, etc.) are elements with two valence electrons. Elements in each group have similar chemical properties because they have the same number of valence electrons.

?

1. How is ionization energy related to electron affinity? ___________________________________

____

____

2. Why don't the noble gases tend to form ions? ____

____

____

3. Why is the electron affinity of sodium ($1s^22s^22p^63s^1$) lower than the electron affinity of fluorine ($1s^22s^22p^5$)? ____

____

____