Physical Science 2020

(This fourth quarter is an Intro to Chemistry)

LEWISBURG HIGH SCHOOL

John Estes (Instructor)

Week 15 and 16 Chapter 17 and 18

(Spring Semister 2020)

April 20 – May 1, 2020

Please Only Take ONE Packet Per Student

Review the Packet and Answer the (10 question) Multiple Choice Questions at the End and a 3 sentence reflective essay. Submit the number with the letter answer and 3 sentence essay to my email at john.estes@dcsms.org or drop answered packet at the school. All other questions and examples are simply for **review** .

Hint: The 10 questions were pulled from the notes and the 3 sentences only need to relate to the subject matter in the chapter review.

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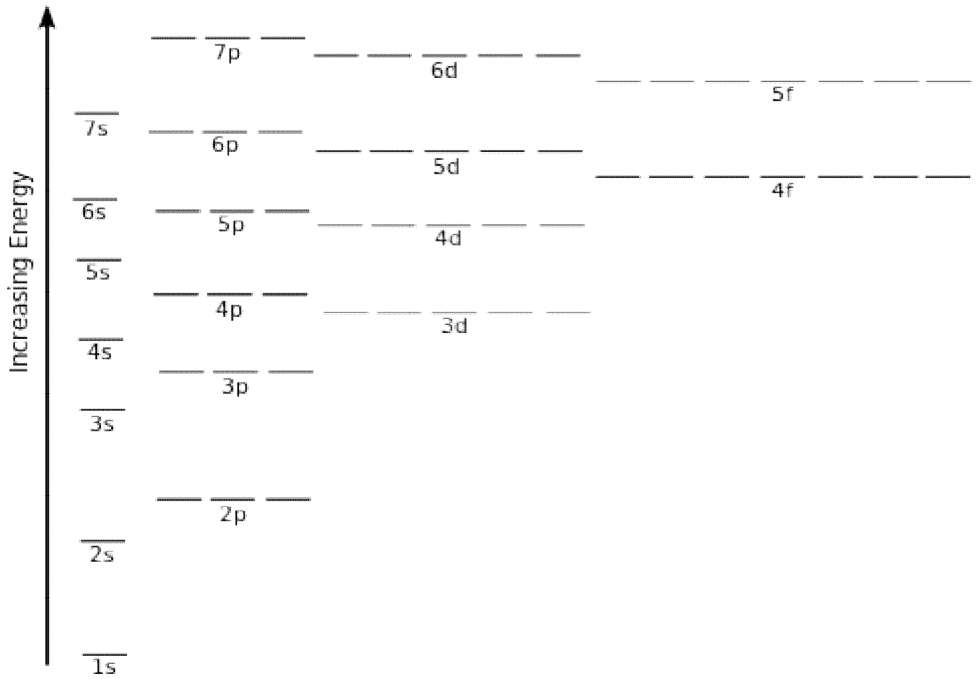
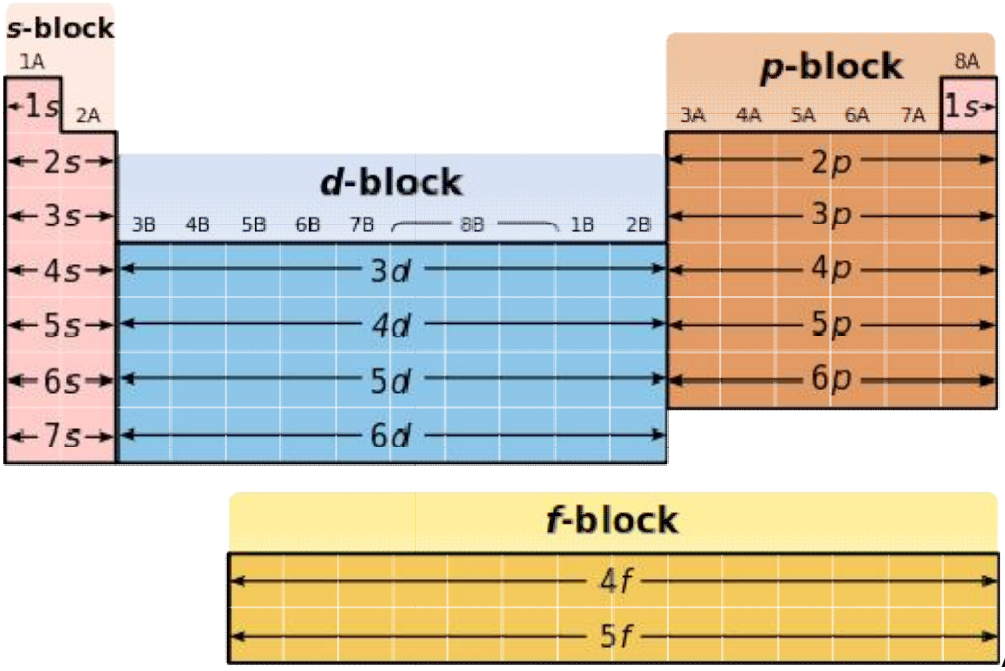
High School Chemistry/Families on the Periodic Table

With the introduction of electron configurations, we began to get a deeper understanding of the Periodic Table. An understanding of these electron configurations will prove to be invaluable as we look at bonding and chemical reactions. The orbital representation method for representing electron configuration is shown below. The orbital representation was learned in an earlier chapter but like many of the skills you learn in chemistry, it will be used a great deal in this chapter and in several chapters later in the course.



In this lesson, we will focus on the connection between the electron configuration and the main group elements of the Periodic Table. We will need to remember the sub-level filling groups in the Periodic Table. Keep the following figure in mind. We will use it for the next two chapters.



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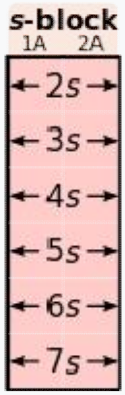
Alkali Metals Have One Electron in Their Outer Energy Level

Elements Ending with s^1 = Alkali

In the Periodic Table, the elements are arranged in order of increasing atomic number. In previous material we learned that the atomic number is the number of protons in the nucleus of an atom. For a neutral atom, the number of protons is equal to the number of electrons. Therefore, for neutral atoms, the Periodic Table is also arranged in order of increasing number of electrons. Take a look now at the first **group** or column in the Periodic Table. It is the one marked "1A" in the Period Table figure above. The groups or families are the vertical rows of elements. The first group has seven elements representing the seven periods of the Periodic Table. Remember that a **period** in the Periodic Table is a horizontal row. Group 1A is the only group with seven elements in it.

Table 9.1: Electron Configurations for Group 1A Metals

Element	Atomic Number	Electron Configuration
Lithium (Li)	3	
Sodium (Na)	11	
Potassium (K)	19	
Rubidium (Rb)	37	
Cesium (Cs)	55	
Francium (Fr)	87	



What do you notice about all of the elements in Group 1? They all have s^1 as the outermost energy level electron configuration. The whole number in front of the "*s*" tells you what period the element is in. For example sodium, Na, has the electron configuration $1s^2 2s^2 2p^6 3s^1$, so it is in period 3. It is the first element of this period.

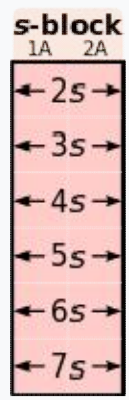
This group of elements is called the **alkali metals**. They get their name from ancient Arabic (al kali) because "scientists" of the time found that the ashes of the vegetation they were burning contained a large amount of sodium and potassium. In Arabic, *al kali* means *ashes*. We know today that all alkali metals have electronic configurations ending in s^1. You might want to note that while hydrogen is often placed in group 1, it is not considered an alkali metal. The reason for this will be discussed later.

Alkaline Earth Elements Have Two Electrons in Their Outer Energy Level

Elements Ending with s^2 = Alkaline Earth

Taking a look at Group 2A in Table 9.2, we can use the same analysis we used with group 1 to see if we can find a similar trend. It is the second vertical group in the Periodic Table and it contains only six elements.

Table 9.2: Electron Configurations for Group 2A Metals

Element	Atomic Number	Electron Configuration	
Beryllium (Be)	4		
Magnesium (Mg)	12		
Calcium (Ca)	20		
Strontium (Sr)	38		
Barium (Ba)	56		
Radium (Ra)	88		

What do you notice about all of the elements in group 2A? They all have an outermost energy level electron configuration of s^2. The whole number in front of the "*s*" tells you what period the element is in. For example, magnesium, Mg, has the electron configuration $1s^2 2s^2 2p^6 3s^2$, so it is in period 3 and is the second element in that period. Remember that the *s* sublevel may hold two electrons, so in Group 2A, the *s* orbital has been filled.

Elements in this group are given the name **alkaline earth metals**. They get their name because early "scientists" found that all of the alkaline earths were found in the earth's crust. Alkaline earth metals, although not as reactive as the alkali metals, are still highly reactive. All alkaline earth metals have electron configurations ending in s^2.

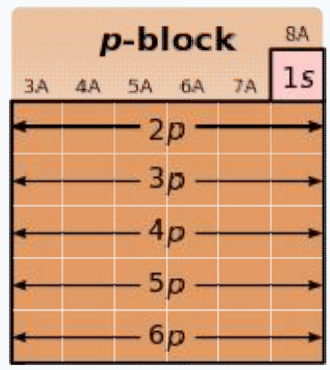
Noble Gases Have 8 Electrons in Their Outer Energy Level

Elements Ending with s^2p^6 = Noble Gases

The first person to isolate a noble gas was Henry Cavendish, who isolated argon in the late 1700s. The noble gases were actually considered inert gases until the 1960s when a compound was formed between xenon and fluorine which changed the way chemists viewed the "inert" gases. In the English language, inert means to be lifeless or motionless; in the chemical world, inert means *does not react*. Later, the name "**noble gas**" replaced "inert gas" for the name of Group 8A.

When we write the electron configurations for these elements, we see the same general trend that was observed with groups 1A and 2A; that is, similar electron configurations within the group.

Table 9.3: Electron Configurations for Group 8A Gases

Element	Atomic Number	Electron Configuration	
Helium (He)	2		
Neon (Ne)	10		
Argon (Ar)	18		
Krypton (Kr)	36		
Xenon (Xe)	54		
Radon (Rn)	86		

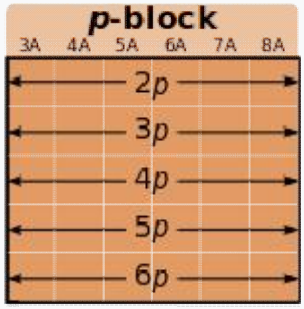
Aside from helium, He, all of the noble gases have outer energy level electron configurations that are the same, ns^2np^6, where n is the period number. So Argon, Ar, is in period 3, is a noble gas, and would therefore have an outer energy level electron configuration of $3s^2 3p^6$. Notice that both the s and p sublevels are filled. Helium has an electron configuration that might fit into Group 2A. However, the chemical reactivity of helium, because it has a full first energy level, is similar to that of the noble gases.

Halogens Have 7 Electrons in Their Outer Energy Level

Elements Ending with s^2p^5 = Halogens

The **halogens** are an interesting group. Halogens are members of Group 7A, which is also referred to as 17. It is the only group in the Periodic Table that contains all of the states of matter at room temperature. Fluorine, F_2, is a gas, as is chlorine, Cl_2. Bromine, Br_2, is a liquid and iodine, I_2, and astatine, At_2, are both solids. What else is neat about Group 7A is that it houses four (4) of the seven (7) diatomic compounds. Remember the diatomics are H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2. Notice that the latter four are Group 17 elements. The word halogen comes from the Greek meaning *salt forming*. French chemists discovered that the majority of halogen ions will form salts when combined with metals. We all know some of these already: LiF, NaCl, KBr, and NaI.

Taking a look at Group 7A in the figure, we can find the same pattern of similar electron configurations as found with group 1A, 2A, and 8A. It is the 17th group in the Periodic Table and it contains only five elements.

Table 9.4: Electron Configurations for Group 7A Elements			
Element	**Atomic Number**	**Electron Configuration**	
Fluorine (F)	9		
Chlorine (Cl)	17		
Bromine (Br)	35		
Iodine (I)	53		
Astatine (At)	85		

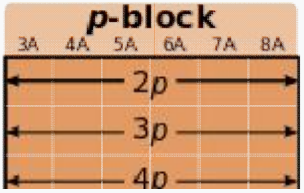
What is the general trend for the elements in Group 7A? They all have, as the outermost energy level electron configuration, ns^2np^5, where n is the period number. You should also note that these elements are one group away from the noble gases (the ones that generally don't react!) and the outermost electron configuration of the halogens is one away from being filled. For example, chlorine (Cl) has the electron configuration [Ne] $3s^23p^5$ so it is in period 3, the seventh element in the **main group elements**. The main group elements, as you recall, are equivalent to the s + p blocks of the Periodic Table (or the pink and orange groups in the diagram above).

The Oxygen Family Has 6 Electrons in the Outer Energy Level

Elements Ending with s^2p^4 = the Oxygen Family

Oxygen and the other elements in Group 6A have a similar trend in their electron configurations. Oxygen is the only gas in the group; all others are in the solid state at room temperature. Oxygen was first named by Antoine Lavoisier in the late 1700s but really the planet has had oxygen around since plants were first on the earth.

Taking a look at Group 6A in the figure below, we find the same pattern in electron configurations that we found with the other groups. Oxygen and its family members are in the 16th group in the Periodic Table. In Group 16, there are, again, only five elements.

Table 9.5: Electron Configurations for Group 6A Elements			
Element	**Atomic Number**	**Electron Configuration**	
Oxygen (O)	8		

Sulfur (S)	16		
Selenium (Se)	34		
Tellurium (Te)	52		
Polonium (Po)	84		

When we examine the electron configurations of the Group 6A elements, we see that all of these elements have the outer energy level electron configuration of ns^2np^4. We will see that this similar electron configuration gives all elements in the group similar properties for bonding.

These elements are two groups away from the noble gases and the outermost electron configuration is two away from being filled. Sulfur, for example, has the electron configuration $1s^22s^22p^63s^23p^4$ so it is in period 3. Sulfur is the sixth element in the main group elements. We know it is the sixth element across the period of the main group elements because there are 6 electrons in the outermost energy level.

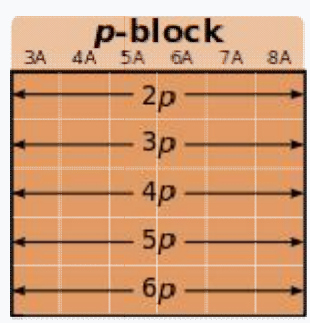
The Nitrogen Family Has 5 Electrons in the Outer Energy Level

Elements Ending with s^2p^3 = the Nitrogen Family

Just as we saw with Group 6A, Group 5A has a similar oddity in its group. Nitrogen is the only gas in the group with all other members in the solid state at room temperature. Nitrogen was first discovered by the Scottish chemist Rutherford in the late 1700s. The air is mostly made of nitrogen. Nitrogen has properties that are different in some ways from its group members. As we will learn in later lessons, the electron configuration for nitrogen provides the ability to form very strong triple bonds.

Nitrogen and its family members belong in the 15th group in the periodic table. In Group 15, there are also only five elements.

Table 9.6: Electron Configurations for Group 5A Elements

Element	Atomic Number	Electron Configuration	
Nitrogen (N)	7		
Phosphorous (P)	15		
Arsenic (As)	33		

Antimony (Sb)	51		
Bismuth (Bi)	83		

What is the general trend for the elements in group 5A? They all have, as the electron configuration in the outermost energy level, ns^2np^3, where n is the period number. These elements are three groups away from the noble gases and the outermost energy level electron configuration is three away from having a completed outer energy level. In other words, the p sublevel in the Group 15 elements is half full. Arsenic, for example, has the electron configuration $1s^22s^22p^63s^23p^64s23d^{10}4p^3$ so it is in period 4, the fifth element in the main group elements. We know it is the fifth element across the period of the main group elements because there are 5 electrons in the outermost energy level.

Lesson Summary

- Families in the periodic table are the vertical columns and are also referred to as groups.
- Group 1A elements are the alkali metals and all have one electron in the outermost energy level because their electron configuration ends in s^1.
- Group 2A elements are the alkaline earth metals and all have two electrons in the outermost energy level because their electron configuration ends in s^2.
- Group 5A elements all have five electrons in the outermost energy level because their electron configuration ends in s^2p^3.
- Group 6A elements all have six electrons in the outermost energy level because their electron configuration ends in s^2p^4.
- Group 7A elements are the halogens and all have seven electrons in the outermost energy level because their electron configuration ends in s^2p^5.
- Group 8A elements are the noble gases and all have eight electrons in the outermost energy level because their electron configuration ends in s^2p^6.
- Elements in group 8A have the most stable electron configuration in the outermost shell because the sublevels are completely filled with electrons.

Review Questions

1. If an element is said to have an outermost electronic configuration of ns^2np^3, it is in what group in the periodic table?

(a) Group 3A

(b) Group 4A

2. What is the general electronic configuration for the Group 8A elements? (Note: when we wish to indicate an electron configuration without specifying the exact energy level, we use the letter "n" to represent any energy level number. That is, ns^2np^3 represents any of the following; $2s^22p^3$, $3s^23p^3$, $4s^24p^3$, and so on.)

(a) ns^2np^6

(b) ns^2np^5

(c) ns^2np^1

(d) ns^2

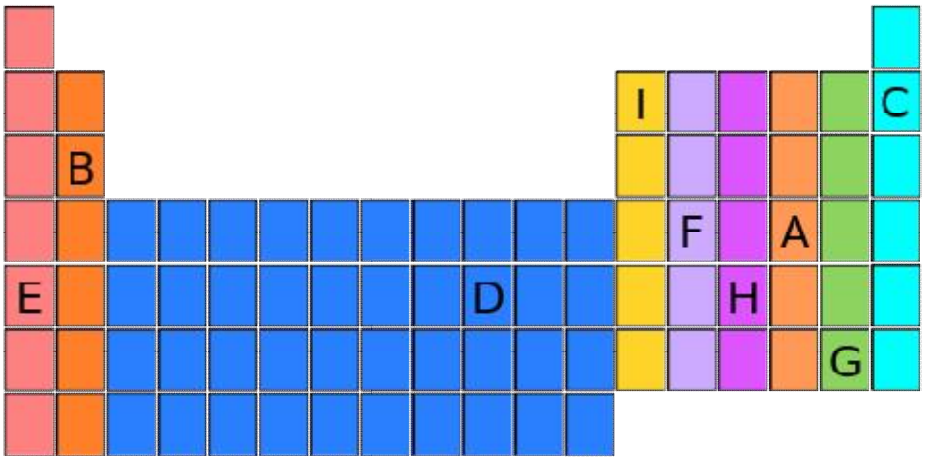
3. The group 2 elements are given what name?

(a) alkali metals
(b) alkaline earth metals
(c) halogens
(d) noble gases

4. Using the diagram below, identify:

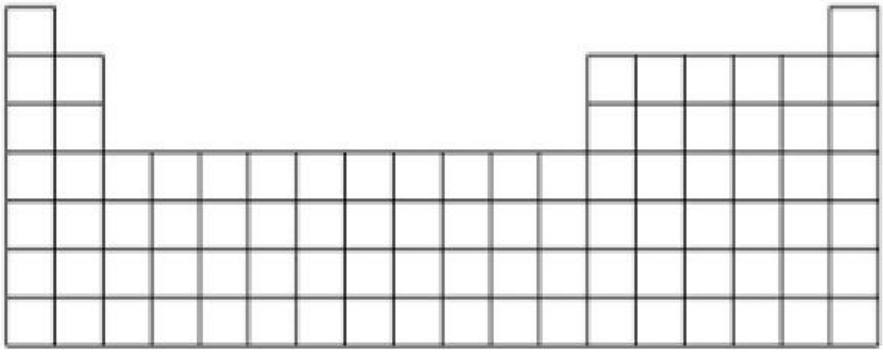
The alkali metal by giving the letter that indicates where the element would be located and

(f)



5. In the periodic table, name the element whose outermost electronic configuration is found below. Where possible, give the name of the group.

(a) $5s^2$ Non Metals
(b) $4s^2 3d^{10} 4p^1$ Alkaline Earth Metals
(c) $3s^2 3p^3$ Nobel Gases
(d) $5s^2 4d^{10} 5p^2$ Transition Metals
(e) $3s^1$ Alkali
(f) $1s^2$
(g) $6s^2 5d^{10} 6p^5$ Halogen
(h) $4s^2 4p^4$
(i) $2s^2 4p^1$ Metals



Questions on the Structure of the Periodic Table

Indicate the letter of the best answer. (Email the question number with the letter answer and 3 sentence essay to John Estes at the School Email Address)

(This is a open note assignment.)

1. On the far left of the Periodic Table is what the first suborbital?
   a. S b. P c. A
2. What has to be equal for the charge of an atom to be neutral?
   a. Protons and Electrons b. Neutrons and Electrons c. Neutrons and Protons
3. What is the name of group 1 elements?
   a. Lanthanides b. Nobel Gases c. Alkali Metals
4. What is the name of the group 2 elements
   a. Alkali Metals b. Alkaline Earth c. Nobel Gases
5. What element of the Nobel Gases does NOT have the ns^2np^6 configuration?
   a. Neon b. Helium c. Argon
6. Who was the first person to isolate a Noble Gas?
   a. Henry Cavendish b. Elon Musk c. Neil Bohr
7. What type of compound does Chlorine by itself tend to make?
   a. Monoatomic b. Diatomic c. Triatomic
8. What are the two atomic suborbitals in the Main Group Elements?
   a. D and F b. A and B c. S and P
9. What element has the electron configuration to form strong triple bonds?
   a. Hydrogen b. Nitrogen c. Chlorine
10. How many outermost electrons are in the Nobel Gas energy level?
    a. 4 b. 6 c. 8

- After you finish the questions leave a three sentence or more essay on what you know, what you wanted to know, or what you learned. Thank you

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Ionic Bonding (chem.libretexts.org)

Ionic bonding is the complete transfer of **valence** electron(s) between atoms and is a type of chemical bond that generates two oppositely charged ions. It is observed because metals with **few** electrons in its outer-most orbital. By losing those electrons, these metals can achieve noble-gas configuration and satisfy the **octet rule**. Similarly, nonmetals that have close to 8 electrons in its valence shell tend to readily accept electrons to achieve its **noble** gas configuration.

Introduction

In ionic bonding, electrons are transferred from one **atom** to another resulting in the formation of positive and negative ions. The electrostatic attractions between the positive and negative ions hold the compound together. The predicted overall energy of the ionic bonding process, which includes the **ionization energy** of the metal and **electron affinity** of the nonmetal, is usually positive, indicating that the reaction is endothermic and unfavorable. However, this reaction is highly favorable because of their electrostatic attraction. At the most ideal inter-atomic distance, attraction between these particles releases enough **energy** to facilitate the reaction. Most ionic compounds tend to dissociate in polar solvents because they are often polar. This phenomenon is due to the opposite charges on each ions.

At a simple level, a lot of importance is attached to the electronic structures of **noble** gases like neon or argon which have eight electrons in their outer energy levels (or two in the case of helium). These **noble** gas structures are thought of as being in some way a "desirable" thing for an atom to have. One may well have been left with the **strong** impression that when other atoms react, they try to organize things such that their outer levels are either completely full or completely empty.

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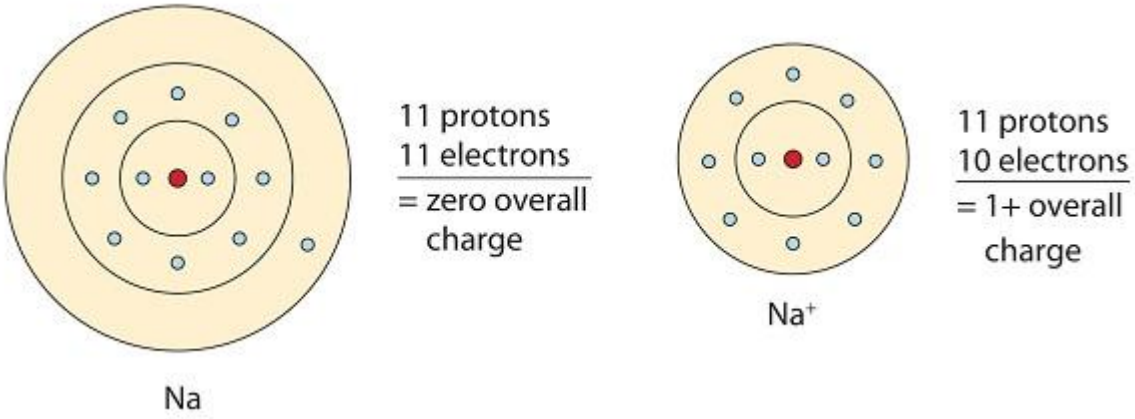
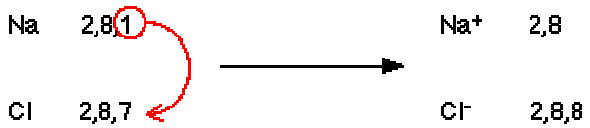
Sodium Chloride:

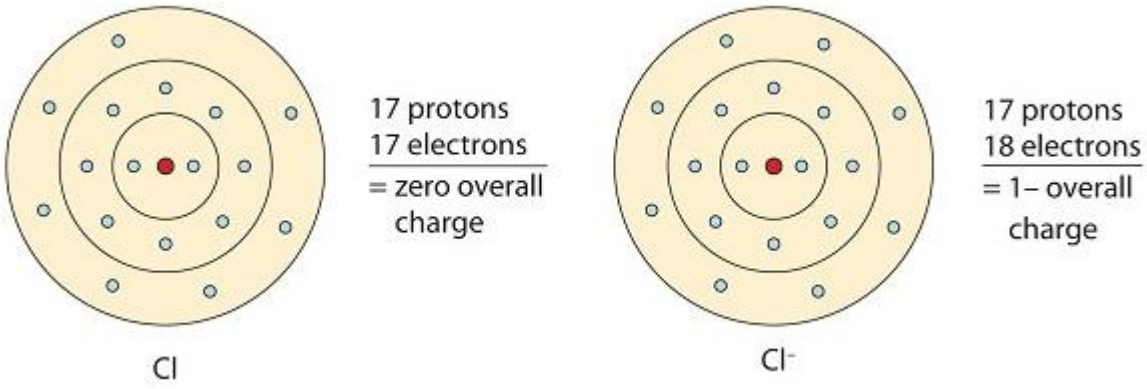
- Sodium (2,8,1) has 1 electron more than a **stable** noble gas structure (2,8). If it gave away that electron it would become more **stable**.
- Chlorine (2,8,7) has 1 electron short of a stable **noble** gas structure (2,8,8). If it could gain an electron from somewhere it too would become more **stable**.

The answer is obvious. If a sodium atom **gives** an electron to a chlorine atom, both become more **stable**.





This sodium molecule **donates** the lone electron in its valence orbital in order to achieve **octet** configuration. This creates a positively charged cation due to the loss of electron.



The sodium has **lost** an electron, so it no longer has equal numbers of electrons and protons. Because it has one more proton than electron, it has a charge of 1+. If electrons are lost from an atom, positive ions are formed. Positive ions are sometimes called **cations**.

The chlorine has **gained** an electron, so it now has one more electron than proton. It therefore has a charge of 1-. If electrons are gained by an atom, negative ions are formed. A negative ion is sometimes called an **anion**.

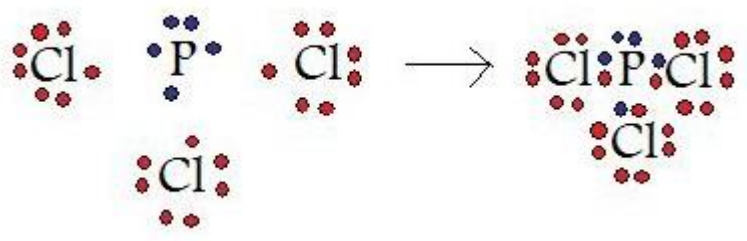
The nature of ionic bonding

The sodium ions and chloride ions are held together by the **strong** electrostatic attractions between the positive and negative charges. You need **one** sodium atom to provide the extra electron for **one** chlorine atom, so they combine together 1:1. The formula is therefore NaCl.

Covalent Bonding

Covalent bonding is the **sharing** of electrons between atoms. This type of bonding occurs between two atoms of the **same** element or of elements close to each other in the periodic table. This bonding occurs primarily between nonmetals; however, it can also be observed between **nonmetals** and **metals**.

If atoms have **similar** electronegativities (the same affinity for electrons), covalent bonds are most likely to occur. Because both atoms have the same **affinity** for electrons and neither has a tendency to donate them, they share electrons in order to achieve **octet** configuration and become more stable. In addition, the ionization energy of the atom is too large and the electron affinity of the atom is too small for ionic bonding to occur. For example: carbon does not form ionic bonds because it has 4 valence electrons, half of an **octet**. To form ionic bonds, Carbon molecules must either gain or lose 4 electrons. This is highly unfavorable; therefore, **carbon** molecules share their 4 valence electrons through single, double, and triple bonds so that each atom can achieve **noble** gas configurations. Covalent bonds include interactions of the sigma and pi orbitals; therefore, covalent bonds lead to formation of single, double, triple, and **quadruple** bonds.



In this example, a phosphorous atom is **sharing** its three unpaired electrons with three chlorine atoms. In the end product, all four of these molecules have 8 **valence** electrons and satisfy the octet rule.

Covalent Bonding in Organic Chemistry

Ionic and covalent bonds are the two extremes of bonding. **Polar** covalent is the intermediate type of bonding between the two extremes. Some **ionic** bonds contain covalent characteristics and some covalent bonds are partially ionic. For example, most carbon-based compounds are covalently bonded but can also be **partially** ionic. Polarity is a measure of the separation of charge in a compound. A compound's polarity is dependent on the **symmetry** of the compound and on differences in electronegativity between **atoms**. Polarity occurs when the electron pushing elements, found on the left side of the periodic table, exchanges electrons with the electron pulling elements, on the **right** side of the table. This creates a spectrum of polarity, with ionic (polar) at one extreme, covalent (nonpolar) at another, and **polar** covalent in the middle.

Both of these bonds are important in **organic** chemistry. Ionic bonds are important because they allow the synthesis of specific organic compounds. Scientists can manipulate **ionic** properties and these interactions in order to **form** desired products. Covalent bonds are especially important since most carbon molecules interact primarily through **covalent** bonding. Covalent bonding allows

molecules to share electrons with other molecules, creating long chains of compounds and allowing more complexity in life.

1. Are these compounds ionic or covalent?

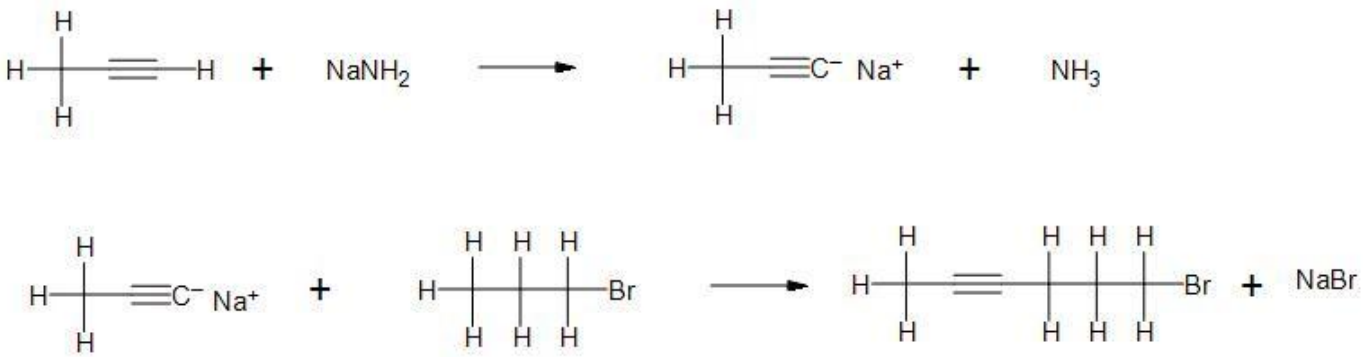
CH_4 Fe_2O_3 KNO_3 $\boxed{I_2}$ C_4 H_2O $BeCl_2$

2. In the following reactions, indicate whether the reactants and products are ionic or covalently bonded.

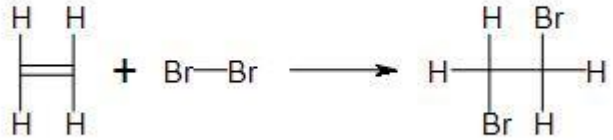
a)

$$Na_2CO_3 + CaCl_2 \longrightarrow CaCO_3 + 2\,NaCl$$

b) Clarification: What is the nature of the bond between sodium and amide? What kind of bond forms between the anion carbon chain and sodium?



c)



Metalic Bonding

A third major type of chemical bonding is **metallic bonding**. Whereas ionic bonds join metals to non-metals and covalent bonds join non-metals to each other, *metallic bonding joins a* **bulk** *of metal atoms*. A metallic

substance may be a pure element (e.g. aluminum foil, copper wires), or it may be a mixture of two or more elements in an **alloy** (e.g. brass instruments, "white gold" jewelry). Metals tend to have high melting points and boiling points suggesting strong bonds between the atoms. Even a soft metal like sodium (melting point 97.8°C) melts at a considerably higher temperature than the element (**neon**) which precedes it in the Periodic Table. However, unlike ionic compounds, metals are usually malleable rather than brittle, suggesting that they do not form a rigid lattice structure of oppositely charged ions; neither do **metals** form bonded molecules like covalent compounds, however. A different model of bonding is necessary to explain the properties of metallic substances. In the 1900's, Paul Drüde came up with the "**sea of electrons**" metallic bonding theory by modeling metals as a mixture of atomic cores (atomic cores = positive nuclei + inner shell of electrons) and valence electrons.

Electron Sea Model

Consider **sodium** metal as an example. **Sodium** has the electronic structure $1s^2 2s^2 2p^6 3s^1$. When **sodium** atoms come together, the electron in the 3s atomic orbital of one sodium atom can **share** space with the corresponding electron on a neighboring atom to form a bond - in much the same sort of way that a covalent bond is formed. The difference, however, is that each **sodium** atom is being touched by eight other sodium atoms - and the sharing occurs between the central atom and the 3s orbitals on *all* of the eight other atoms. Each of these eight is in turn being touched by eight sodium atoms, which in turn are touched by **eight** atoms - and so on and so on, until you have taken in all the atoms in that lump of **sodium**. *All* of the 3s electrons on all of the atoms are shared in *nondirectional* bonds which extend over the whole piece of metal. The electrons can move freely within the lump of metal, and so each electron becomes detached from its parent atom. The electrons are said to be **delocalized**. The metal is held together by the strong forces of attraction between the positive nuclei and the delocalized electrons (Figure 9.5.19.5.1).

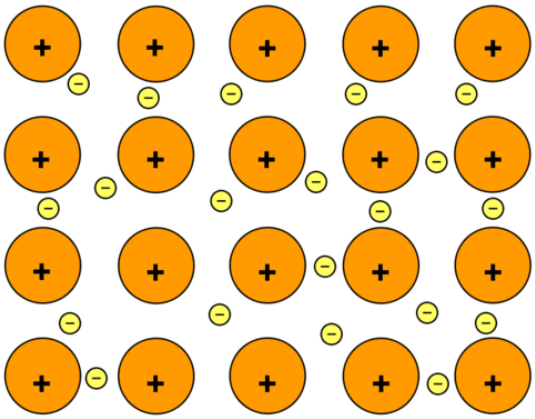


***Figure** 9.5.19.5.1**:** Metallic Bonding: The Electron **Sea** Model: Positive atomic nuclei (orange circles) surrounded by a **sea** of delocalized electrons (yellow circles).*

This is sometimes described as "an array of positive ions in a **sea** of electrons". If you are going to use this view, beware! Is a metal made up of atoms or ions? It is made of **atoms.** Each positive center in the diagram represents all the rest of the atom apart from the outer electron, but that electron has not been lost - it may no longer have an attachment to a particular **atom**, but it's still there in the structure. Sodium **metal** is therefore written as NaNa, not Na+Na+.

In a molten **metal**, the metallic bonding is still present, although the ordered structure has been broken down. The metallic bond is not fully broken until the **metal** boils. That means that boiling point is actually a better guide to the strength of the metallic bond than melting point is. On melting, the bond is loosened, not broken. The strength of a metallic bond depends on three things:

1. The **number** of electrons that become delocalized from the metal
2. The **charge** of the cation (metal).
3. The **size** of the cation.

A strong metallic bond will be the result of more delocalized electrons, which causes the effective **nuclear** charge on electrons on the cation to increase, in effect making the size of the cation smaller. Metallic bonds are strong and require a great deal of **energy** to break, and therefore *metals have high melting and boiling points***.** A metallic bonding theory must explain how so much bonding can occur with such few electrons (since metals are located on the **left** side of the periodic table and do not have many electrons in their **valence** shells). The theory must also account for all of a metal's unique chemical and physical properties.

EXAMPLE 9.5.19.5.1: METALLIC BONDING IN MAGNESIUM

Use the **sea** of electrons model to explain why magnesium has a higher melting point (650 °C) than sodium (97.79 °C).

Solution

If you work through the same argument above for sodium with magnesium, you end up with stronger bonds and hence a **higher** melting point.

Magnesium has the **outer** electronic structure $3s^2$. Both of these electrons become delocalized, so the "sea" has *twice* the electron **density** as it does in sodium. The remaining "ions" also have twice the charge (if you are going to use this particular view of the metal bond) and so there will be more attraction between "**ions**" and "**sea**".

More realistically, each magnesium atom has 12 **protons** in the nucleus compared with sodium's 11. In both cases, the nucleus is screened from the delocalized electrons by the same number of **inner** electrons - the 10 electrons in the $1s^2\ 2s^2\ 2p^6$ orbitals. That means that there will be a **net** pull from the magnesium nucleus of 2+, but only 1+ from the sodium nucleus.

So not only will there be a **greater** number of delocalized electrons in magnesium, but there will also be a greater attraction for them from the magnesium nuclei. Magnesium atoms also have a slightly **smaller** radius than sodium atoms, and so the delocalized electrons are closer to the nuclei. Each magnesium atom also has twelve near neighbors rather than sodium's **eight**. Both of these factors increase the strength of the bond still further.

Note: Transition metals tend to have particularly **high** melting points and boiling points. The reason is that they can involve the 3d electrons in the delocalization as well as the 4s. The **more** electrons you can involve, the stronger the attractions tend to be.

Questions On Bonding and Elements

Indicate the letter of the best answer. (Email the question number with the letter answer and 3 sentence essay to John Estes at the School Email Address)

(This is a open note assignment.)

1. What is completely transferred in ionic bonding between atoms?
   a. Inner electrons b. valence electrons c. outer protons
2. What set of elements have eight outer most valence electrons in their atomic structure?
   a. Transition metals b. Halides c. Noble gases
3. When a sodium atom reacts to form a ionic bond it tends to do what with it's valence electron?
   a. Gain b. Lose c. Share
4. Bonds between atoms that are the same (other than metals) tend to be what bond?
   a. Metallic b. Ionic c. Covalent
5. How many valence does carbon have?
   a. One b. Four c. Eight
6. What is the other extreme of the spectrum of polarity from ionic polar?
   a. Metallic Pairing b. Covalent Nonpolar c. Hydrogen Bonding
7. How many electrons are involved in resolving the octet rule?
   a. Four b. Six c. Eight
8. What characteristics of metal bonds helps explain the properties of that metal?
   a. Ionization energy b. electron affinity c. sea of electrons
9. Metals are located on what side of the Periodic Table?
   a. Left b. Right c. Bottom
10. How many protons are on each atom of Magnesium?
    a. 6 b. 8 c. 12

- After you finish the questions leave a three sentence or more essay on what you know, what you wanted to know, or what you learned. Thank you