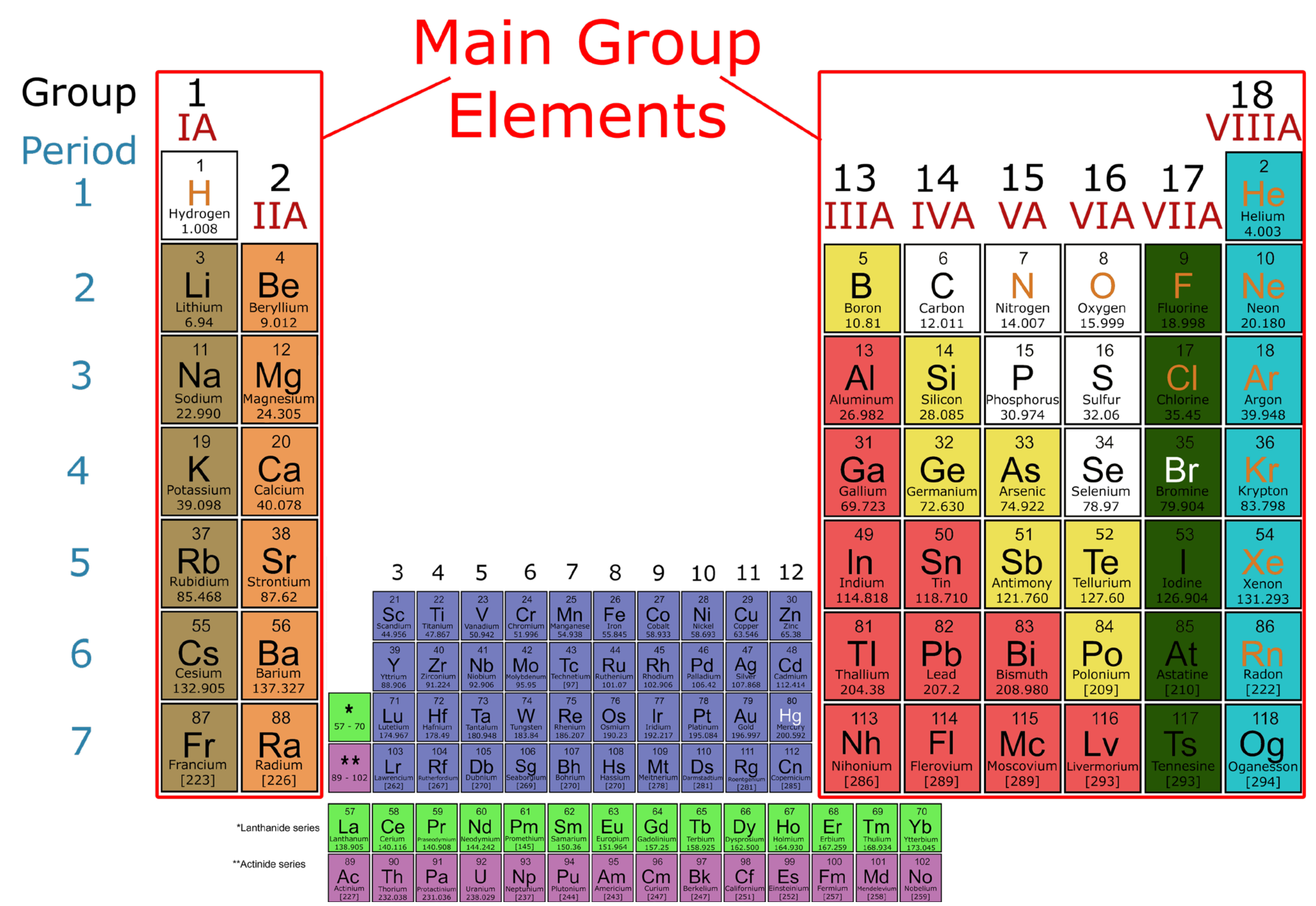
5.3 MAIN GROUP ELEMENTS

Now, when we study chemical bonding, there will be parts of the Periodic Table that are better at following the rules than other parts, because there are many exceptions to the rules with the Periodic Table. However, the groups of elements called the "main group elements" are quite good at following the rules and they are the ones that we will use when studying bonding. The main group elements are those in Groups 1,2 and 13–18. Now, the main group elements are not a made-up category for this course. No, they are a recognized category of the Periodic Table because chemists have long noted that they are the elements that follow the rules. The elements we leave out from the main group are the transition metals, Groups 3–12 (including elements 57–70 and 89 –102). See **Figure 5.3.1** for details.

Figure 5.3.1

Main Group Elements
Here we have the main group elements—Groups 1, 2, 13, 14, 15, 16, 17 and 18. It may seem like we are excluding a lot of elements by focusing "only" on those 8 groups, but the main group elements comprise almost all the known matter in the universe. Therefore, they are the ones that are best studied and most likely to be involved with chemical reactions. So, just to reinforce, in our introductory course, when we talk about chemical bonds, we will always be talking about bonding that occurs between atoms of elements in the main group.



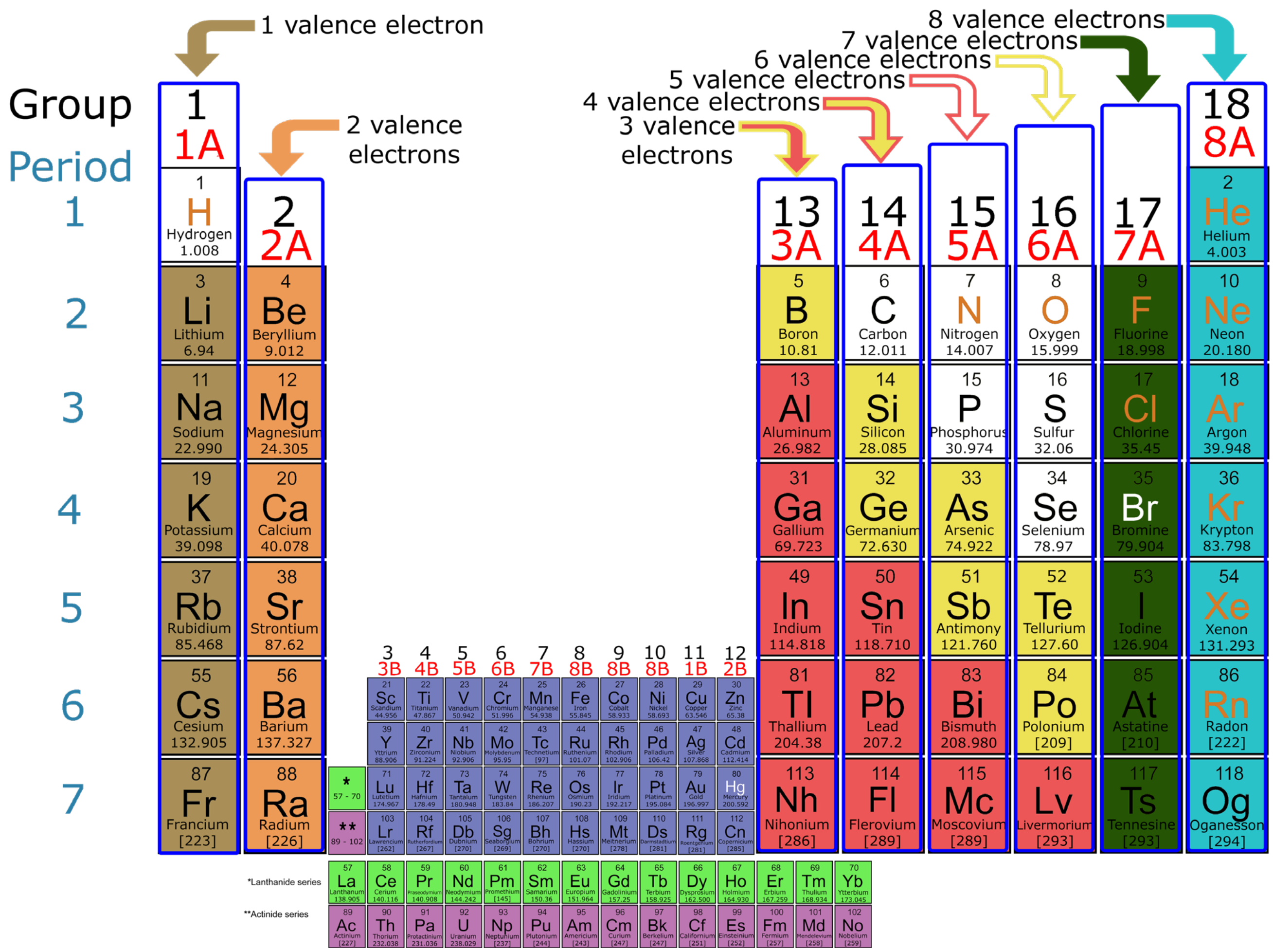
I don't want to beat this horse to death, but it is so important to understand that the chemical properties of an element are governed to a large extent by its electrons and their shell configurations. Is there 1 electron in the outer shell, 2 electrons, etc.? Will this atom participate in a chemical reaction, or not? If it does, how do I predict how many atoms of one element I need to react with how many atoms of another element? All of this kind of information is knowable based upon the periodicity of the Table, but only as much as the rules are followed. We stick with the main group elements for our chemical bonding discussion because the main group elements follow the rules, especially when it comes to electron configurations and valence electrons. The valence electrons are key when learning how atoms bond.

Remember that one of our rules from the last chapter is that elements in the same group have the same number of electrons in their outer shells. Another way to say this using our new term is that the Periodic Table groups are organized according to valence electrons. Recall that when we discussed this in the last chapter, I said, "generally" the elements in the same group have the same number of electrons in their outer shells. Well, when we remove the transition metals and just focus on the main group elements, they always follow that rule. Another rule that the main group elements follow is that they tend to show the strongest "periodic properties" in the Periodic Table (if you need a refresher, take a look at **Figure 4.5.2**). Because the main group elements follow the rules, especially regarding valence electrons, they are the most predictable and teach us the most about bonding in a preparatory course.

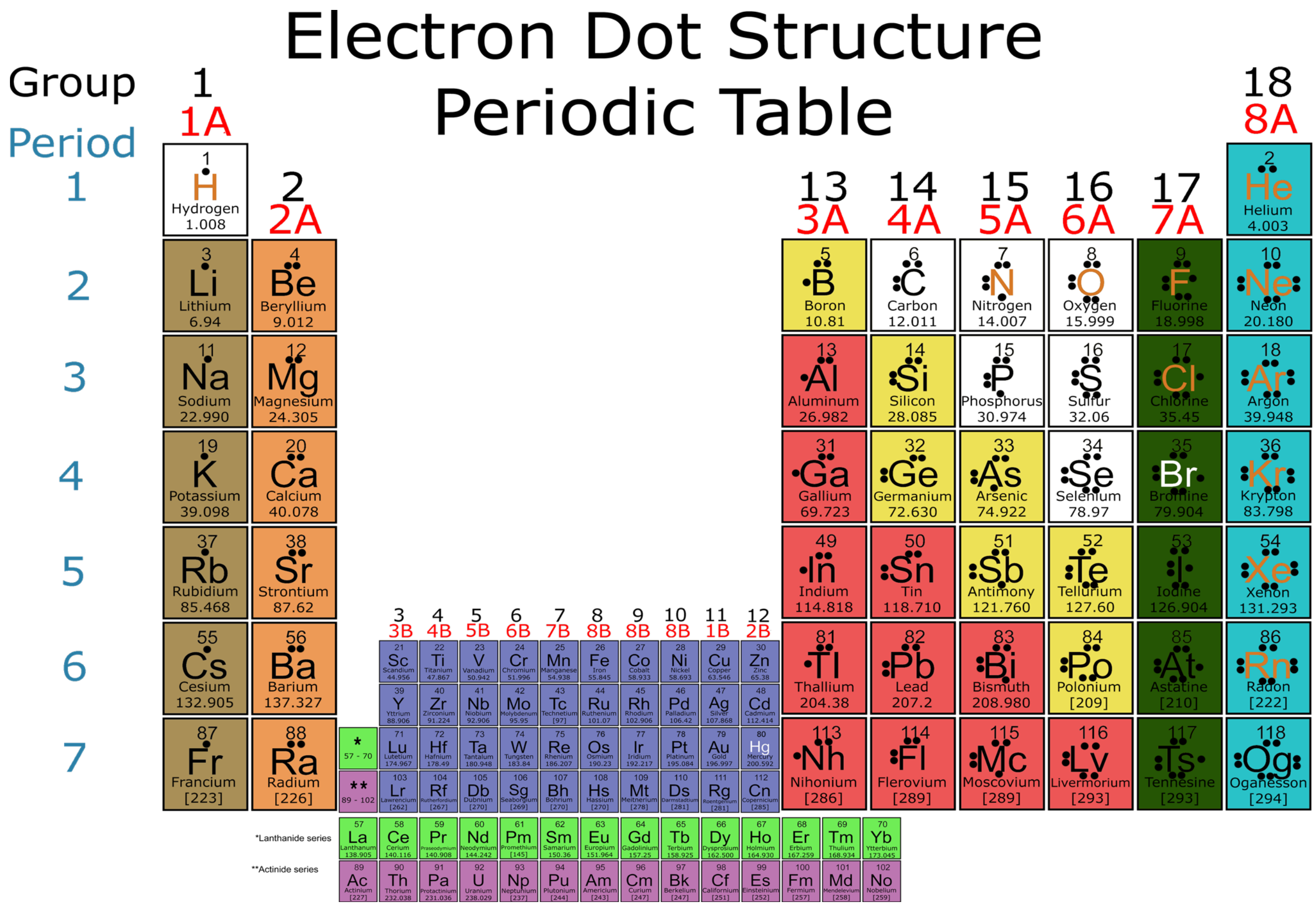
Figure 5.3.2

Helpful Electron Information from the 8-Group System

Last chapter, in **Figure 4.3.2**, we talked about the 8-Group and the 18-Group system, and I said that we'd have a little more to say about the 8-group system. Well, here it is! The 8-Group system provides easy-to-remember information about the valence electron configurations of an element. The reason that we exclude Groups 3—12 is that they tend to not follow the electron bonding rules as tightly as the main group elements. The main group elements—Groups 1/1A, 2/2A, 13/3A, 14/4A, 15/5A, 16/6A, 17/7A and 18/8A –all follow the rules. Group 1/1A elements have 1 valence electron, Group 2/2A elements have 2 valence electrons, Group 13/3A elements have 3 valence electrons, Group 14/4A elements have 4 valence electrons, Group 15/5A elements have 5 valence electrons, Group 16/6A elements have 6 valence electrons, Group 17/7A elements have 7 valence electrons and Group 18/8A elements have 8 valence electrons. This rule is followed for all elements of the group. This information is very valuable as we start to learn more and more about bond formation.



Knowing how many valence electrons an element has allows us to accurately predict if two atoms will, or will not, react with each other and, if they do, we can also figure out in what ratios they will do so.



5.4 ATOMIC BONDING, VALENCE ELECTRONS AND THE OCTET RULE (OR "ATOMS LOVE STABILITY")

Bonding is all about electrons and, more specifically, valence electrons. We have mentioned many times that the chemical properties of an atom/element are in large part controlled by an atom's electrons, and the way that chemical bonds occur reflect that. What is it about electrons that controls bonding? Why do atoms bond to one another? Well, it has to do with all atoms "wanting" to be in the most stable state that they can possibly be. A "stable state" is when the outer shell of the atom has 8 electrons in it and when an atom doesn't have 8 valence electrons, it looks for at least one other atom that also does not have 8 electrons in its outer shell so they can form a bond, get 8 electrons and be more stable. This "desire" to be more stable with 8 electrons in the outer shell, or the property of forming a bond so that each atom involved in the bond has 8 electrons in the outer shell, is called the **octet rule**.

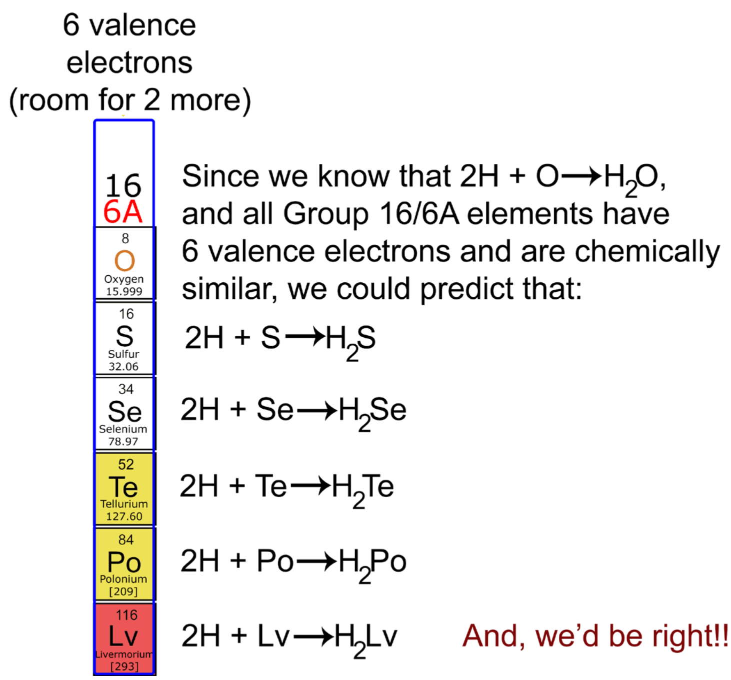
There are a couple of different ways that the octet rule is fulfilled in a bond, and we will learn about the two most common ways next chapter. But, for now, it is enough to know about the octet rule—atoms are more stable when they have 8 valence electrons.

- Atoms in the same group have the same number of valence electrons; therefore, they have similar chemical properties.
- Elements in the same group, then, would be expected to participate in similar chemical reactions in similar ways.

After you finish **Figure 5.4.2**, we will move onto bonding models and let the importance of valence electrons age a little bit until Chapter 7.

Figure 5.4.2

Predictability of Chemical Properties of Main Group Elements Makes Chemistry a LOT Easier
The periodicity of the main group elements helps make chemistry a whole lot easier. If you know a chemical reaction takes place for one of the elements in a group, because the other elements in that group contain the same number of valence electrons, chances are very, very good that the other elements in the same group will participate in a chemical reaction in the same way as the one you know. Let's look at this practically with water for a second. Water is H_2O—two atoms of hydrogen participate in a chemical reaction with one atom of oxygen and bond together to form H_2O (we'll look at the Bohr model for water in **Figure 5.5.1**). Since oxygen and hydrogen are both main group elements, we can expect them to follow the rules. Looking at the valence electron structures, oxygen is a Group 16/6A element, with 6 valence electrons. There are 5 other Group 16/6a elements—sulfur, selenium, tellurium, polonium and livermorium. If I were to ask you to predict if polonium would react with hydrogen and, if you think it does, how many atoms of hydrogen would it take, you would easily be able to answer that question even if you knew nothing about polonium. You'd be able to answer because you know about the Periodic Table of the Elements. You would figure that, since oxygen and polonium are both Group 16/6A elements, they both have 6 valence electrons, leaving both with "2 open slots" for electrons in their outer shell; therefore, you could feel comfortable concluding that polonium, like oxygen (and all of the other elements in Group 16/6A),would react with two atoms of hydrogen to form H_2Po. You can do this, despite knowing not one thing about polonium or its chemical reactions, because you do know the Periodic Table of the Elements and that its periodic properties mean that main group elements in the same group are chemically very similar.

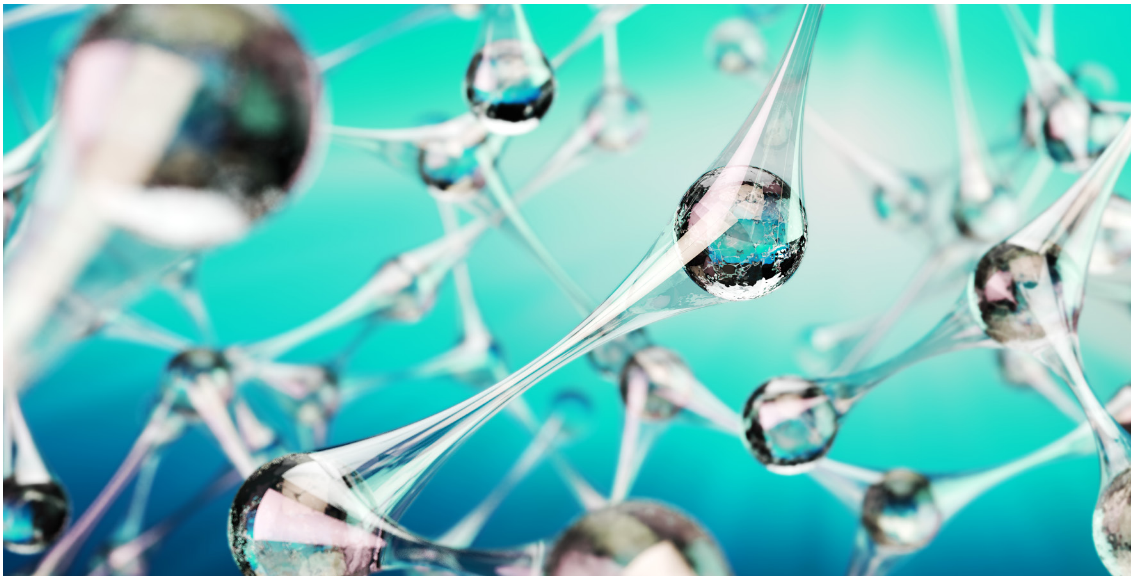


5.5 BONDING MODELS

There are several different ways that a bond between atoms can be displayed and before we go any further, I want to review the different models that are used so that you are familiar with them since we will be using them quite a bit from here on out. As you might imagine, if we used the Bohr models to draw out what molecules look like, then it would take forever to draw a very large molecule. Even drawing out the molecular structure of water is a stretch! See **Figure 5.5.1** for details. So, there are different ways to draw molecules and their bonding relationships to one another. Each model is a little different and has its own advantages and disadvantages, but each one is good for its use. There is no "perfect model," so when you are deciding to draw a molecule, you can choose which model you think is best for your purpose.

Chemical Bonding 2

6



6.0 CHAPTER PREVIEW

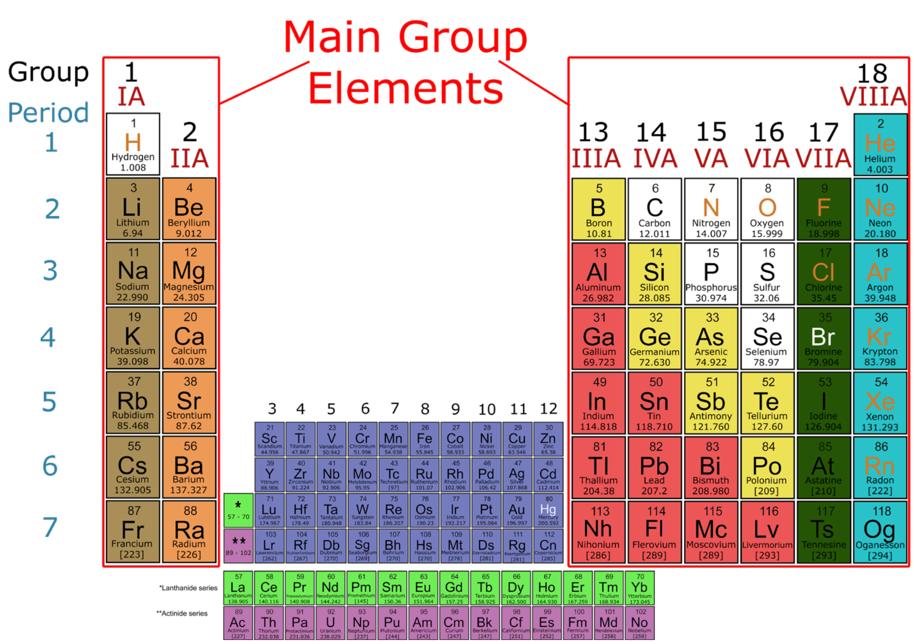
In this chapter, we will:

- Briefly review important concepts from Chapter 5.
- Study two types of bonds—ionic and covalent.

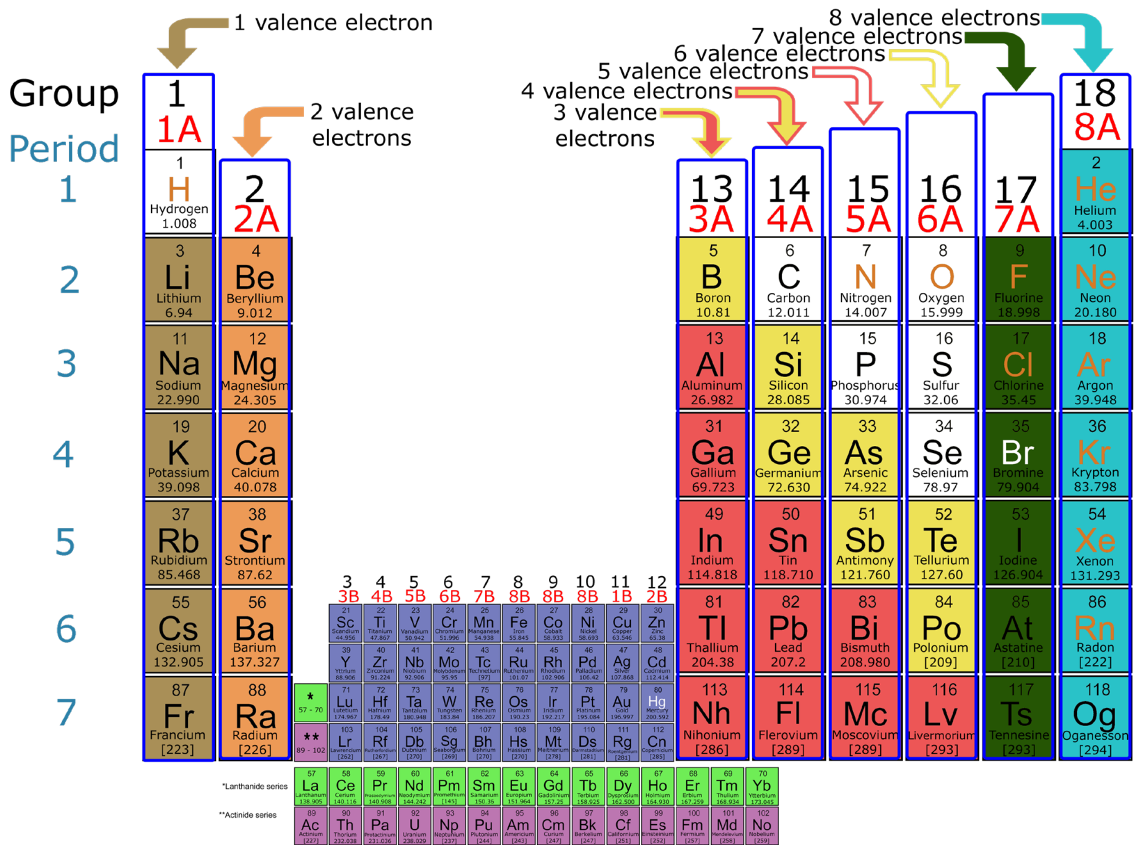
6.1 INTRODUCTION

Let's reinforce a couple of concepts from the last chapter, and then get deeper into chemical bonding. From Chapter 5:

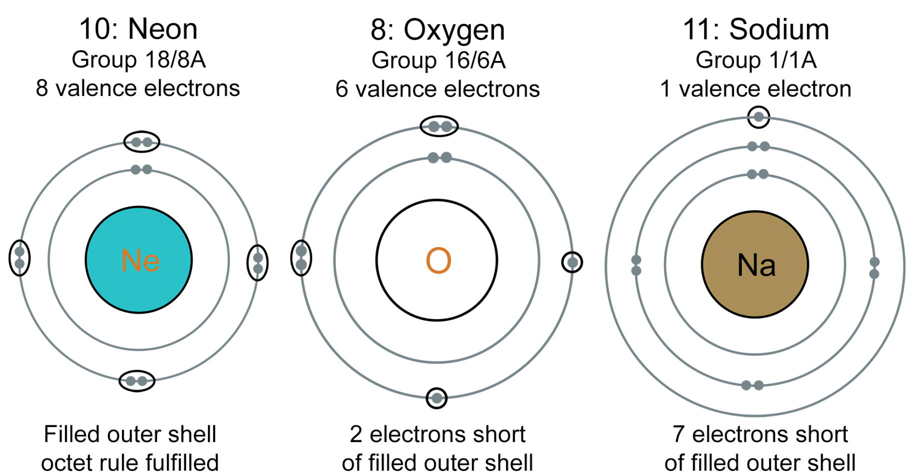
- We are focusing our bond discussion on the main group elements, Groups 1/1A, 2/2A, 13/3A, 14/4A, 15/5A, 16/6A, 17/7A and 18/8A.



- Valence electrons are critical to the discussion of chemical bonds, because they are what largely control the bonding characteristics of an atom.



- Atoms love to be stable, which means that they want to get 8 electrons in their outer shells to fulfill the octet rule (or, if hydrogen, 2 electrons in its outer shell to fulfill the duet rule).



6.6 COVALENT OR IONIC, HOW DO YOU KNOW?

This is a very technical subject and since this is an introductory course, we will take advantage of the periodicity of the Periodic Table to answer this question for our purposes. When speaking about the main group elements, the ones in Groups 1/1A and 2/2A—the metals—easily form ionic bonds with the elements in Group 17/7A—the halogens—and most of the elements color-coded as "non-metals;" specifically, hydrogen (H), carbon (C), nitrogen (N), oxygen (O), phosphorous (P), sulfur (S) and selenium (Se). For covalent bonds, the Groups 13/3A—17/7A elements readily form covalent bonds with other elements from those groups, as well as with hydrogen from Group 1. Remember the Group 18/8A elements (the Noble gases) are generally inert.

Figure 6.6.1

Covalent Bonds by the Periodic Table
We can take advantage of the predictability of the Periodic Table and make the following general statements about covalent bonds and be correct most of the time. Generally, the elements in Groups 14/4A through 17/7A, and hydrogen, will form covalent bonds with one another. Note the "**" notation in the middle of the graphic. We are constantly walking the line between "over-simplifying" and "being too technical" for this preparatory class. The "covalent bonding rule" described below is far more of a guideline for the area I circled in black than for the rest of the elements inside the blue rectangle. In our problems for the end of chapter questions and exams, I will always tell you what you need to know in order to answer the question, and usually I will stick with elements that follow the rules when asking you questions.

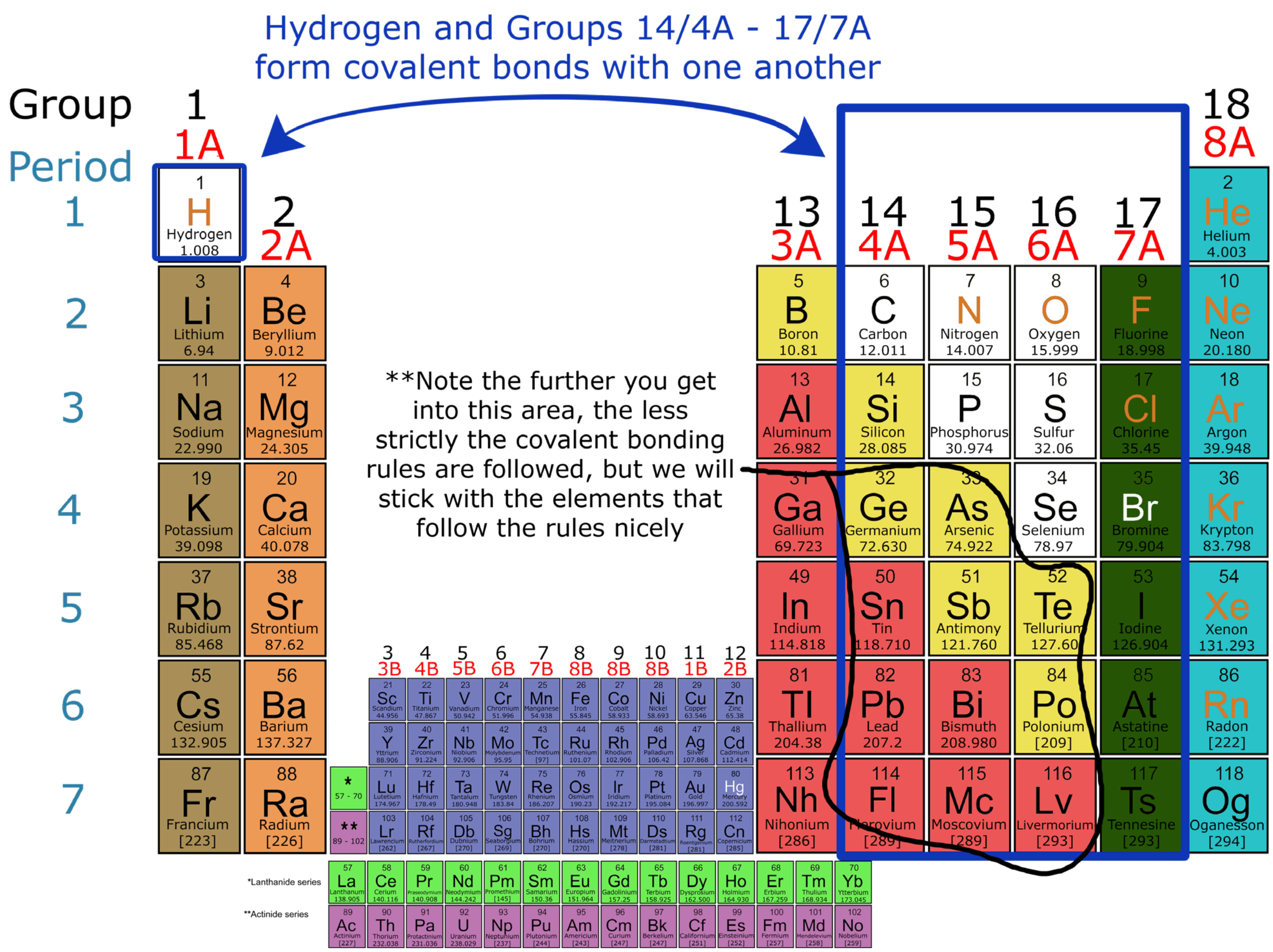
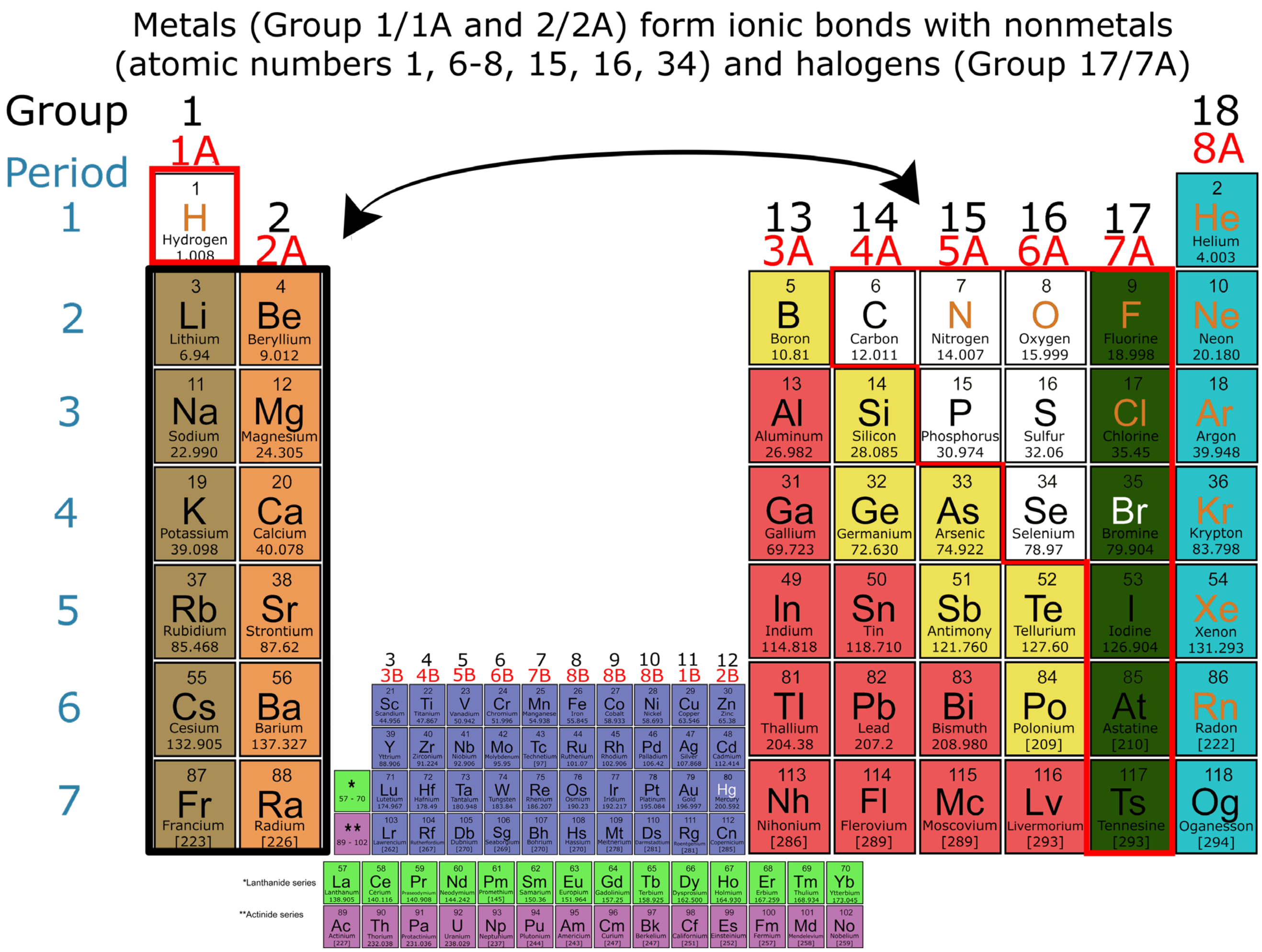


Figure 6.6.2

Ionic Bonds by the Periodic Table
The same Table predictability allows us to understand which elements will form ionic bonds, too. Generally, the metals in Groups 1/1A and 2/2A will form ionic bonds with the non-metals (the elements whose boxes are color-coded white) and the halogens (Group 17/7A).



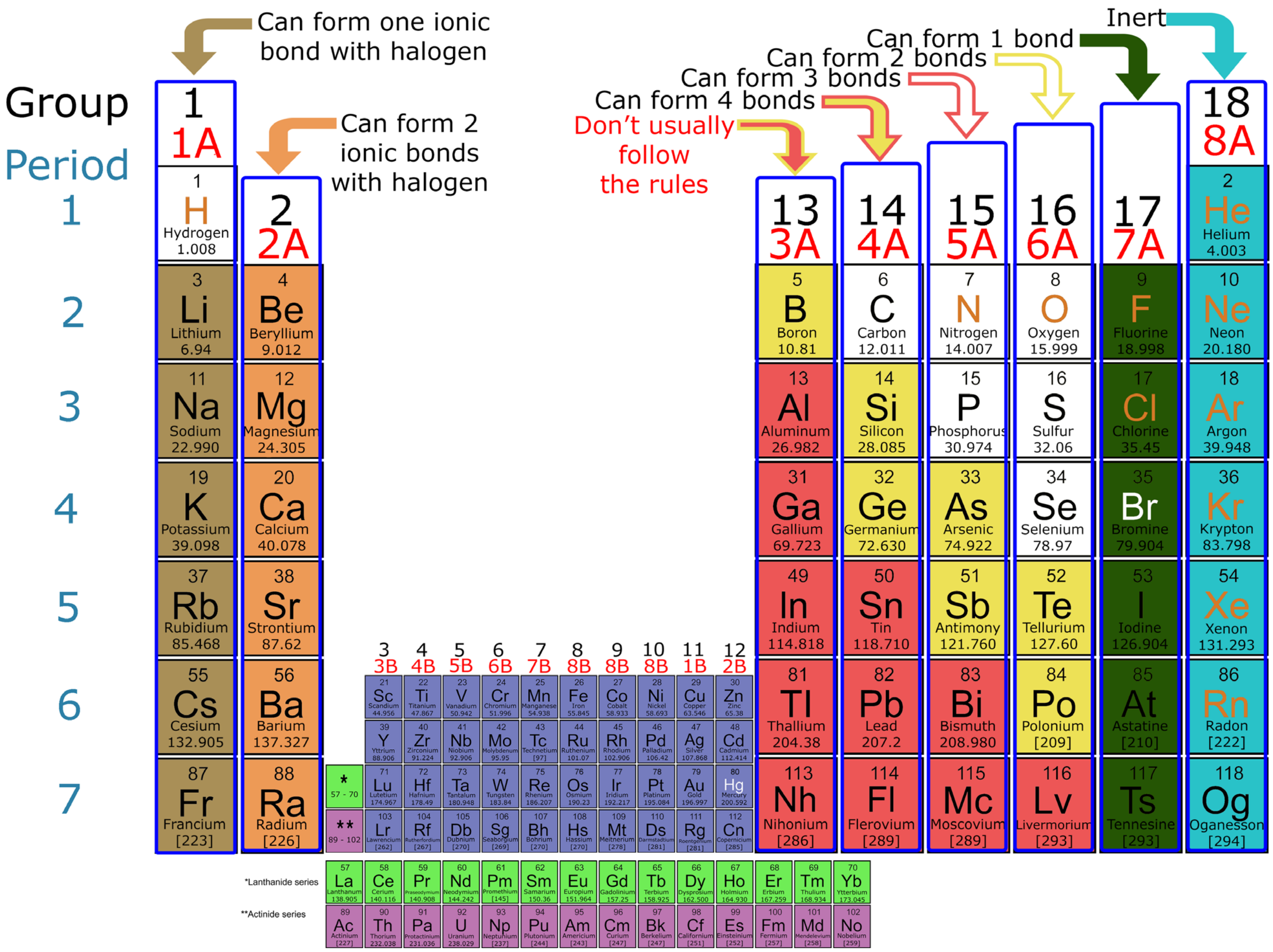
6.7 NUMBER OF BONDS, HOW DO YOU KNOW?

How many bonds can an element form? Let's find out! Now, like knowing whether an atom can form an ionic or covalent bond, this is very complicated, so I want to keep it straightforward for this class but still hold true to general chemistry principles. We will continue to stick with the main group elements and limit this "number of bonds" discussion to the elements in Rows 1–4. And, even then, it is important to realize that we are still talking "in general" properties so if you want to look for exceptions to this rule, you will be able to find them. However, the elements often follow this rule and so it is helpful to learn it for the foundation it provides for the future.

With ionic bonds, the alkali metals (Group 1/1A) and alkaline earth metals (Group 2/2A) can form the number of bonds equal to the Group they are in. So, lithium, sodium and potassium, being in Group 1/1A, can each form one ionic bond with one non-metal anion (atomic numbers 6-8, 15, 16, 34 and the halogens). We saw an example of this in **Figure 6.5.1** with Na and Cl bonding to form NaCl. The Group 2/2A elements can each form two bonds with anions of atomic numbers 6–8, 15, 16, 34 and the halogens (as we saw in **Figure 6.5.2**).

Figure 6.7.1

Predicting the Number of Bonds Possible
We can use the periodicity of the Table to make a generalized rule about the number of bonds formed for the main group elements in Periods 1–4, while noting that the elements in Periods 3 and 4 don't follow it quite as well as the ones in Periods 1 and 2. For the two groups of metals on the left side of the Table, all form ionic bonds. Indeed, one of the strongest properties of the Group 1 and 2 metals is that they always lose electrons and become cations to form ionic bonds. For example, the Group 1 elements lose one electron to become the 1^{+} cation in the bond, the halogen atom's element accepts the electron to become the anion, and they form the ionic bond. Likewise, the Group 2/2A element loses two electrons to become a 2^{+} cation and two atoms from Group 17/7A each pick up one of the electrons to become a 1^{-} anion and then they form the ionic bond. The way that ionic bonds form between the Groups 1/1A and 2/2A metals and carbon, nitrogen, oxygen, phosphorous, sulfur and selenium is a bit more complicated, and if I ask you any ionic bonding questions at the end of the chapter or on a test using these elements, I will make sure to give you the information you need to answer the question. For covalent bonds, to determine the number of bonds the elements in Groups 13/3A–17/7A can form, just subtract the number of valence electrons from 8. For example, carbon is atomic number 6, which means that it has 2 electrons in its inner shell and 4 valence electrons. Subtract 4 from 8 and you get 4, so carbon can form a total of 4 covalent bonds. Sulfur is atomic number 16 in Group 16/6A, Row 3, so it has 16 electrons; 2 in the inner shell, 8 in the second shell and 6 in the outer shell. So, 8 – 6 = 2, so sulfur can form 2 covalent bonds.



For covalent bonds, the octet rule allows us to formulate an easy-to-remember rule of thumb to determine the number bonds—subtract the number of valence electrons from 8 and that will give you the number of covalent bonds the element can form. Double bonds require 2 electrons, and triple bonds 3, to complete the bond. **Figure 6.7.2** has more details regarding double and triple bonds and how the electrons participate in those.

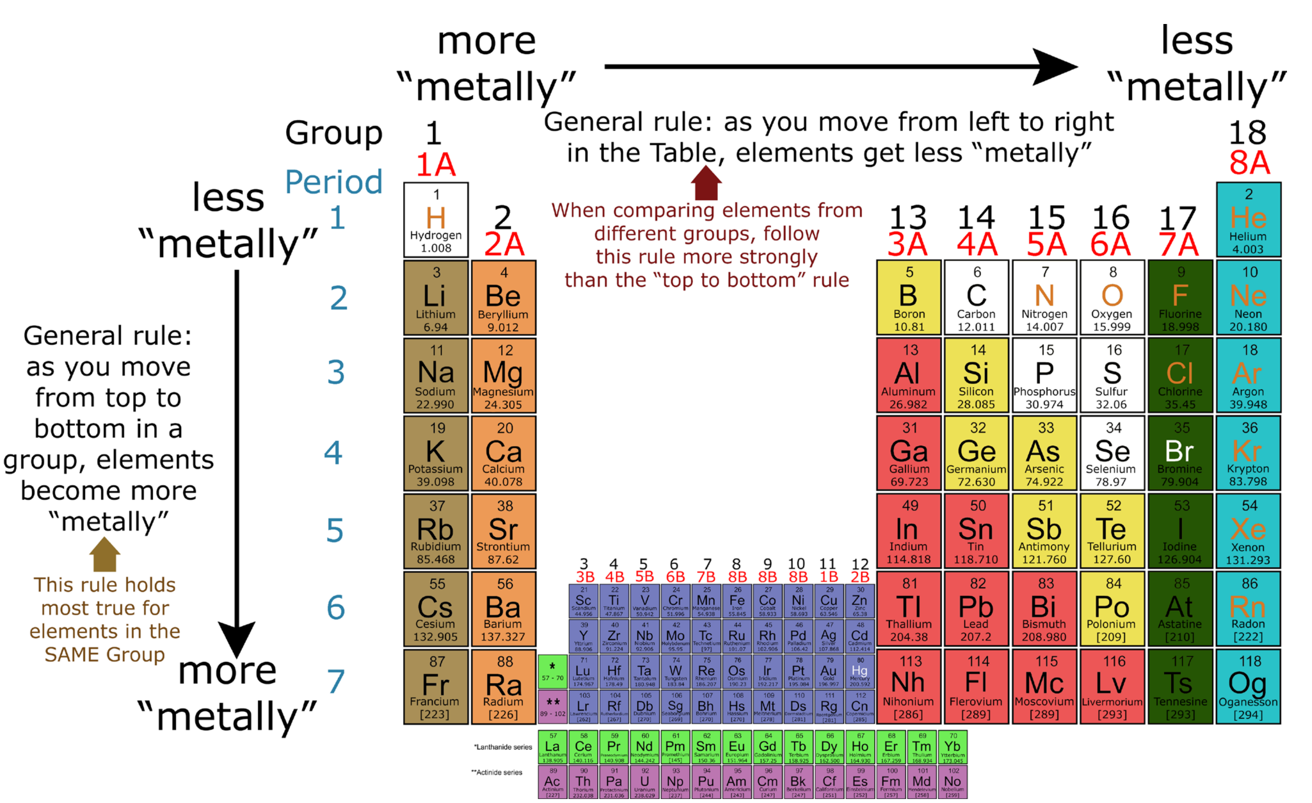
Note that there can be deviation from the "most metallic element is written first" rule as it pertains to the list of 10. Usually, the naming rules are followed, but not always and sometimes you just need to go with the flow so I just want you to be aware that the rules aren't ALWYAS followed strictly. For example, while hydrogen (H) is a non-metal, it's also in Group 1, all the way to the left of the Table. According to our rule of metallic nature, since it's way over to the left in the Table, H should be the "most metallic" of those 10 elements in the list of 10. So, if we were to follow the rule of writing the most metallic element first, it seems like hydrogen should be written before carbon, phosphorus and nitrogen. But it isn't. The reason is mainly historical, meaning that they are sometimes listed "out of order" because that's how chemists have always done it. Not a great reason, I know, but that is how it is done.

Before we move on from this subject let's do a couple more examples.

1. Write the chemical formula for the molecular compound containing 1 selenium atom for every 2 oxygen atoms.
   - To figure this out, look at the Periodic Table and see that neither Se nor O are in Group 1/1A or 2/2A AND both are color-coded as "non-metals."
   - Selenium is not in the list of 10 but oxygen is; therefore, we need to determine, based upon our knowledge of the Periodic Table, which is more metallic, Se or O.
     - Selenium and oxygen are both in the same group and selenium is further down in the column; therefore, it is more metallic and is written first in the chemical formula.

   SeO_2 is correct.

Figure 8.4.1 (continued)



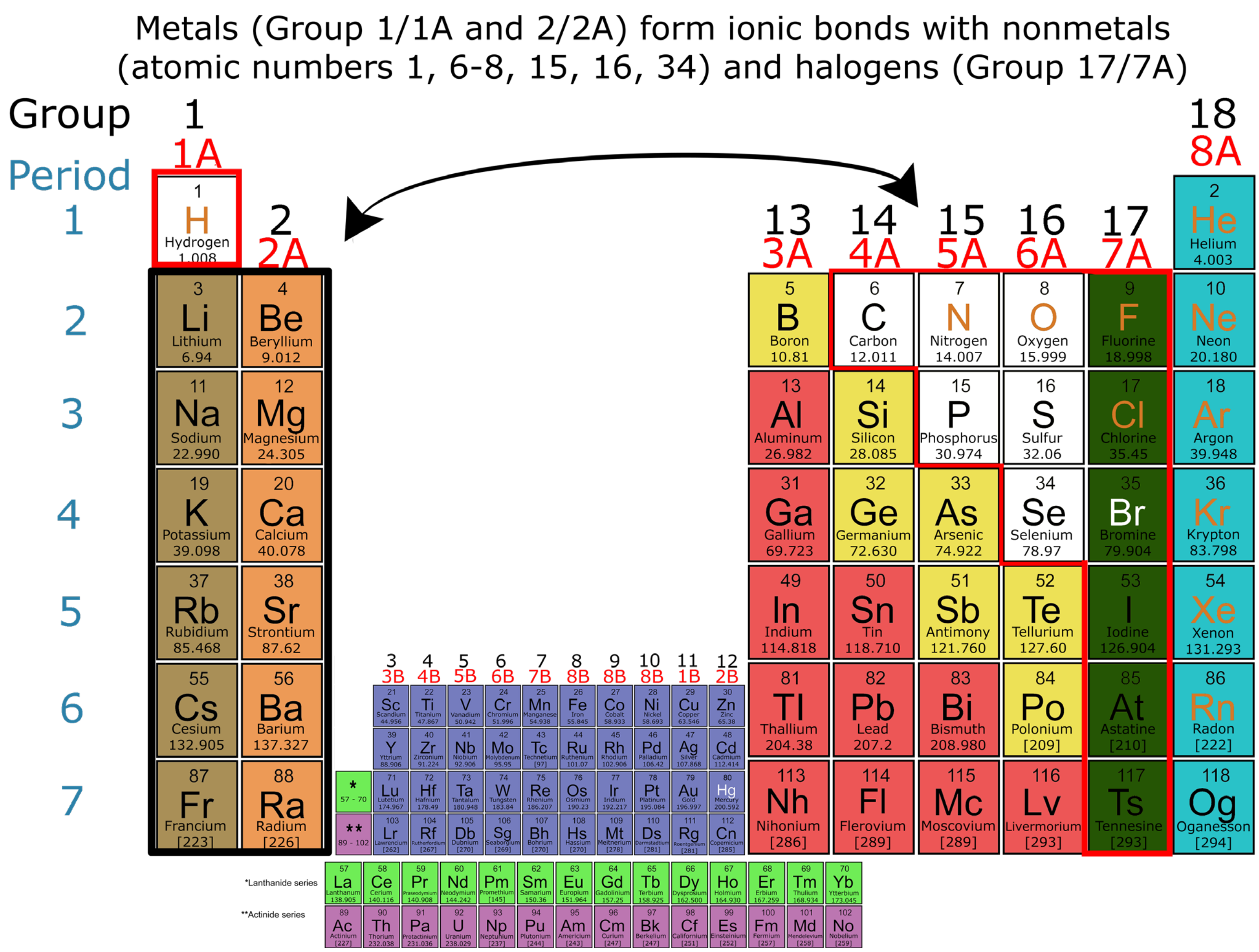
8.5 NAMING COMPOUNDS—IONIC

When naming a chemical compound, first determine if it is an ionic or molecular compound. Fortunately, this is pretty easy—if the compound contains an element from Group 1/1A or 2/2A, then it's an ionic compound. If it doesn't contain a Group 1/1A or 2/2A element, then for the purpose of our course, it is a molecular compound. I say "for the purpose of our course" because it gets a little more complicated when dealing with the non-Main Group elements and some of the Main Group elements. But what we are learning will set a good foundation for your chemistry education going into the future.

When naming either an ionic or molecular compound, we follow the same rules that we did for writing the order of the elements in a chemical formula; however, since ionic compounds all contain a metal, this means that the metal is always written first in an ionic compound. And, since we are learning how to name binary elements only, by default the other element in the ionic compound is written second.

Figure 8.5.1

Ionic Bonds and the Periodic Table
The elements outlined with black and red are the ones that form ionic bonds. The elements in black are metals, and any time a metal is in a compound, the bonds are ionic, so the Group 1/1A and 2/2A elements form ionic bonds with the elements outlined in red. This ionic bonding information will be handy for the discussion on naming compounds held together by ionic or covalent bonds. Refer to it as needed during this discussion.



8.6 NAMING COMPOUNDS—MOLECULAR

Remember that we defined the term "molecular compound" a little earlier, but here it is again because we are going to use it quite a bit. Naming molecules that are held together by covalent bonds gets really long-winded if we keep needing to read, "molecules held together by covalent bonds." So, instead of writing that out, we'll use "molecular compounds."

Figure 8.6.1

Covalent Bonds and the Periodic Table

The Periodic Table can also instruct as to which elements are most likely to form covalent bonds. The elements outlined in blue form covalent bonds with one another. As we discussed a little in Chapter 6, as you get toward the bottom of Groups 14/4A–16/6A, the elements don't follow the covalent bonding rules as closely as the other elements do; therefore, we will stick with the elements that do follow the rules.