Introduction

Part of learning any new field is learning how the people in that field communicate with each other. If you were going to become a plumber, you'd familiarize yourself with the names of the tools of the trade and the techniques you'd use. You'd also have to learn what all those funny-shaped pieces of metal under the sink are called. To learn **chemistry**, you need to know the names of the basic tools chemists use: elements and chemical compounds.

Chemical formulas

In almost any field, whether it's scientific or not, there is a language that people familiar with the field understand. For a chemist, part of this language is the **chemical formula**.

The chemical formula of a compound uses atomic symbols with **subscripts** to indicate the amount of each type of element in a compound. For example, the compound water has a formula written this way: $\mathbf{H_2O}$. Thus, water has one part oxygen to two parts hydrogen. Ammonium sulfide has a formula written like this: $\mathbf{(NH_4)_2S}$. Thus, ammonium sulfide has **two** parts nitrogen, **eight** parts hydrogen, and **one** part sulfur.

Subscripts modify **only** what is directly in front of them. If they immediately follow parenthesis, the subscripts modify what is in the parenthesis - as in ammonium sulfide. We will discuss how to write chemical formulas for chemical compounds and how to name chemical compounds in the next few note packs.

Two types of compounds

Naming basic chemical compounds is fairly simple. You name chemical compounds in a **systematic way**. This means that the amount of memorization you have to do is minimal. To start off our discussion of naming compounds, the first thing to realize is that there are two large categories we can put most compounds into. Each category is named using its own conventions, so by looking at a chemical name you can determine something about how the compound is put together.

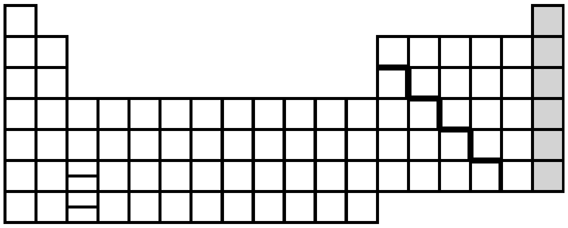
We can categorize chemical compounds into **ionic compounds** and **molecular compounds**. These categories are based on the nature of the forces that hold the compound together.

Ionic compounds and ionic formulas

Ionic compounds are made of particles called **ions**. These ions are formed when **atoms gain or lose electrons**. Since electrons are negatively charged, an ion formed by the loss of electrons carries a positive charge, while an atom formed by a gain of electrons has a negative charge. The positively charged ions are called **cations**, while the negatively charged ions are called **anions**.

Type of ion	***Charge***	***Formed by***
cation	positive (+)	loss of electrons by an atom
anion	negative (-)	gain of electrons by an atom

But why do atoms lose or gain electrons in the first place? We'll learn later that normal chemical reactions proceed in such a way that the reacting atom ends up with a stable configuration of electrons. The atoms on the periodic table that are most stable (that is, they don't react with anything) are the noble gases located on the far right-hand side. These are group VIIIA, 0, or 18 depending on which group numbering system your periodic tables uses.

	• The noble gases are shaded gray in the periodic table to the left. • The noble gases normally do not react with other elements. • Noble gases have very stable arrangements of their electrons.
Illustration 1: The "noble gases" on the periodic table	

If the arrangement of the electrons in the noble gases account for their stability, then it's logical to suggest that **atoms form ions to get an arrangement of electrons like a noble gas**.

According to this hypothesis, then, the **sodium atom** would be expected to lose an electron to get the same number of electrons as the nearest noble gas, **neon**. This would form a **sodium ion**. The **magnesium atom** (one over from sodium), would lose two electrons, forming the **magnesium ion**. We abbreviate these two ions as Na^+ and Mg^{2+}. Experimentally, it has been verified that sodium normally forms an ion with a +1 charge and magnesium normally forms an ion with a +2 charge.

What about elements like fluorine? The nearest noble gas to **fluorine** is again **neon**. Fluorine, though, has to **gain** an electron to match neon. We'd expect that the **fluorine atom** would gain an electron to become the **fluoride ion**, with a charge of -1. The nearest noble gas to oxygen is also neon. To match neon, the oxygen atom must gain **two** electrons to form the oxide ion, with a charge of -2. We abbreviate these two ions as F^- and O^{2-}. Experimentally, it has been verified that in many ionic compounds, fluorine and oxygen form the two ions above.

In fact, you can tell what kinds of ions many elements are likely to form using simple rules-of-thumb derived from the above. If you use the US system of group numbers (IA-VIIIA), you can use the group number to tell what charge an ion will have:

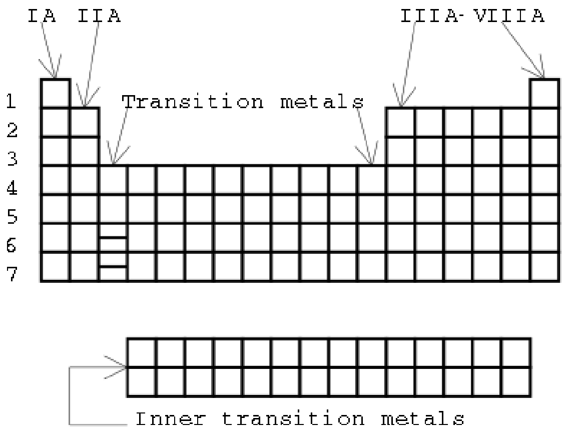


Illustration 2: Groups vs. kinds of ions formed

- The atoms in groups IA and IIA can form **cations** with a charge equal to their group number.
- The atoms in groups VA VIA and VIIA can form **anions** with a charge equal to -(8-group number). Remember that 8 is the group number of the noble gases in this numbering system.
- The transition metals form **cations**, but they have different charges under different conditions. More on that later.
- The noble gases, group VIIIA, don't tend to form ions at all.

Metals tend to form cations, and nonmetals tend to form anions. The noble gases (though they're considered nonmetals) don't readily form ions.

As you might expect, ions of opposite charge attract each other, and it's these attractions (called **electrostatic interactions**) that cause **ionic compounds** to form. These oppositely charged ions arrange themselves in such a way that each anion is surrounded by cations and vice versa. This arrangement is called a **crystal**. A crystal of sodium chloride is a bunch of sodium ions and chloride ions packed together in a regular fashion!

In water, which is **not** an ionic compound (it's molecular), an oxygen atom is attached to two hydrogen atoms and no others. In an ionic compound like sodium chloride, the sodium ions are attracted to whatever chloride ions happen to be around. We can talk about the chemical formula of an ionic compound, which is simply t**he smallest ratio of one type of ion to the other(s) in the compound**. (This is also called the **ionic formula**.) Since all stable chemical compounds must be electrically **neutral**, it's easy to come up with the chemical formulas of ionic compounds. Simply arrange the ions so that their charges all balance out to zero. For example, here are a few ions and the ionic compounds they form.

Cation (+)	***Anion (-)***	***Ionic formula***
Sodium, Na^+	Bromide, Br^-	$NaBr$
Calcium, Ca^{2+}	Bromide, Br^-	$CaBr_2$
Strontium, Sr^{2+}	Oxide, O^{2-}	SrO
Potassium, K^+	Sulfide, S^{2-}	K_2S
Calcium, Ca^{2+}	Nitride, N^{3-}	Ca_3N_2

Molecular compounds and molecular formulas

Molecular compounds are formed when atoms share electrons with other atoms. We won't go into detail on why atoms share electrons just yet, but the reason is similar to the reason that ionic compounds gain and lose electrons. In molecular compounds, unlike ionic compounds, atoms are attracted to only certain other nearby atoms - the ones they're sharing electrons with. These atoms that are hooked together ("**bonded**") by electron sharing form units called **molecules**. Each molecule in a molecular compound contains a fixed number of atoms. A water **molecule** contains two hydrogen atoms and one oxygen atom. Its chemical formula is H_2O. Since the chemical formula of a molecular species indicates exactly how many atoms are in each molecule, it is also called the **molecular formula**. Here are a few example molecular compounds.

Name of molecule	***Molecular formula***
Water (dihydrogen monoxide)	H_2O
Carbon dioxide	CO_2
Carbon tetrafluoride	CF_4
Phosphorus pentachloride	PCl_5
Dinitrogen tetrafluoride	N_2F_4

We'll learn more about how molecular compounds are put together near the end of the course. For now, we'll concern ourselves only with naming and writing formulas for very simple molecular compounds.

Molecular compounds vs. ionic compounds - how to tell them apart

Given a set of atoms that you know are present in a chemical compound, how can you tell them apart? There are a few rules of thumb that let you quickly identify whether a compound will be molecular or ionic if you know what atoms it contains.

1. A compound containing a *metal bonded with nonmetal(s)* will usually be ionic. A good example of a compound of this type is sodium chloride, NaCl.
2. A compound containing *several nonmetals bonded together* will usually be molecular. A good example of a compound of this type is water, H_2O.
3. A compound containing a *metalloid (semiconductor) bound to a nonmetal* will usually be molecular.

Remember that these are **rules of thumb**. There are exceptions to the rules above, and we will make note of them as we come to them.

Summary

In this note pack, we discussed what a chemical formula was. We then looked at the chemical formulas of typical ionic and molecular compounds. In the case of ionic compounds, we looked in some detail at how to write the proper formulas (with correct charges) of ions and how to use those ions to write correct ionic formulas. In the case of molecular compounds, we briefly discussed how these compounds are formed and gave some examples. In future note packs, we will name both ionic and molecular compounds.