

## Chapter 6 – Nomenclature of Inorganic Compounds

Nomenclature is the system used to name compounds. In reality, there are actually many systems for naming compounds and we will cover a few of them in this chapter. With one exception, that we address at the beginning of this chapter, the names of compounds are generated systematically, rather than arbitrarily. Nonetheless, the presence of multiple systems requires attention to detail. Naming compounds is exceptionally important in chemistry because in the name is the identity and composition of the material. For example, saying calcium chloride tells a chemist that not only does the compound contain calcium and chlorine, but that the calcium and chlorine exists as ions in the compound, with the calcium as cation and chlorine as anion, and that they appear in a one-to-two ratio.

### 6.1 Common and Systematic Names

In the early days, when alchemy existed, but chemistry had not yet begun, chemicals still existed and were worked with. At that time, elements as we know them did not exist, but air, earth, fire, and water did. Elements were known, but not as substances that could not be broken down, they were simply seen as materials that were particularly resistant to being broken down. As you know, the ancients knew gold, silver, copper, and sulfur (brimstone) existed because they appear in many old documents, including the Bible. Usually, they had no idea of the composition of the material, other times they knew that one material derived from a known second material. The result was a hodgepodge of names that ranged from descriptions of the appearance of the material, to its appearance, to what it was derived from, the location of its discovery, etc.

Your book provides a handful of examples (Table 6.1, p. 104) that you may have seen, here are a few with more information: laughing gas ( $\text{N}_2\text{O}$ ), muriatic acid ( $\text{HCl}$ , muriatic means “derived from salt”), flowers of zinc ( $\text{ZnO}$ , so called because the solid looked like tiny flowers when viewed under a microscope), oil of vitriol ( $\text{H}_2\text{SO}_4$ , because it interacted violently with other compounds, including skin), etc. The first compound is called baking soda by everyone outside of chemistry laboratories. The second is a name you probably have never heard before unless you worked at a swimming pool, where muriatic (hydrochloric) acid is used to adjust the pH of the water. The final

two compounds are never called their common/trivial names by anyone today, but are sometimes seen in old documents. Some common names are used by everyone, even chemists, for reasons that are obvious. For example, vitamin-D3 is 9,10-secocholesta-5,7,10(19)-triene-3 $\beta$ ,25-diol and dihydrogen oxide is water.

Moving forward through time, that compounds had components became obvious, but the exact compositions weren't known. This led to the first systematic method of naming compounds, which in the mid-20<sup>th</sup> century was replaced by the current system. We will use the newest system almost exclusively, but there are places where knowing the older system is helpful (e.g. baking soda = sodium bicarbonate = sodium hydrogen carbonate = NaHCO<sub>3</sub>).

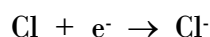
## 6.2 Elements and Ions

### Elements

The large majority of elements are represented by a single atom (e.g. He, Ca) whether they exist as free atoms or in a bulk solid, like a metal. A small number of elements exist as diatomic molecules (H<sub>2</sub>, N<sub>2</sub>, O<sub>2</sub>, F<sub>2</sub>, Cl<sub>2</sub>, Br<sub>2</sub>, and I<sub>2</sub>), but their existence as molecules is very important and you should memorize this list. Phosphorus (P<sub>4</sub>), sulfur (S<sub>8</sub>), and selenium (Se<sub>8</sub>) also exist as molecules. You should remember this because you will see phosphorus and sulfur molecules in the future, but the exact formula is optional for now.

### Ions (Part A)

Remember that in neutral atoms the number of protons in the nucleus equals the number of electrons in the space around the nucleus. Ions are charged particles and atoms become ions when electrons are added or subtracted. The number of protons never changes. (If it did, the atom would become a different element.) Elemental cations are positively charged ions that result from an atom losing one or more electrons. Elemental anions are negatively charged ions that result from an atom adding one or more electrons. For example,



Elemental ions such as these are frequently formed in reactions similar to these, but

complex ions (ions with more than one atom: e.g.  $\text{CO}_3^{2-}$  (carbonate) and  $\text{NO}_3^-$  (nitrate)) form by more complex pathways.

### Ions (Part B)

This is really just an introduction. There is a deep dive into this in Section 6.4.

### Cations

With only one exception ( $\text{NH}_4^+$ ), all of the cations you will see in CHM 211/212 are elemental cations. Cations usually form directly from metals (*vide infra*) and their name is simply the metal name. For example,  $\text{Na}^+$  is the sodium cation and when used in compounds, the cation is simply referred to as “sodium.” For example,  $\text{NaCl}$  is sodium chloride. When looking at names, one differentiates elemental sodium ( $\text{Na}$ ) from ionic sodium ( $\text{Na}^+$ ) by context. When sodium exists as the metal, it only binds to other sodium atoms, so the atoms carry a zero charge. In contrast, in compounds most elements have a greater tendency to attract electrons than sodium, so the sodium has a +1 charge because it gives an electron what becomes the anion.

There is one significant adjustment that must be made for some cations. While many metals, including all of those from Groups 1A and 2A form a single cation, some metals including nearly all of the transition metals, lanthanides, and actinides can form two or more different cations (zinc and cadmium are the only exceptions). For example, there are two common iron ions, +2 ( $\text{Fe}^{2+}$ ) and +3 ( $\text{Fe}^{3+}$ ). In this case, the charge, expressed as Roman numerals in parentheses is placed after the metal to indicate charge.



For elements in Groups 3A - 6A, some form only one ion, others two or more and there is no good generalization so these must be learned individually. For these elements, if you aren't sure, it is better to include the charge and not need it than to need it and not have provided it.

### Elemental Anions

Naming elemental anions is a little more complicated than cations. Here, all elements form only one anion. When naming an anion, part of the element's name is dropped and “-ide” is added. For example, elemental Cl is chlorine and the anion,  $\text{Cl}^-$ , is chloride. Here, the “-ine” in chloride is replaced with “-ide” to indicate the species is an anion. The challenge when naming

anions is that there is no well-defined place to cut off the parent element's name. Thus, oxygen becomes oxide ( $O^{2-}$ ). The construction of the anion name is somewhat idiosyncratic and is done largely by how the name sounds/is pronounced. For oxygen, cutting at the syllable break would give "oxyide," which would be difficult to say and sounds clunky, so it becomes oxide. Nitrogen would not sound good as "nitroide", so, like oxygen, it is truncated earlier to nitride ( $N^{3-}$ ). There are trends in how the endings are added to the elements, but no one rule. Table 6.2, p. 109, has the major elemental anions listed.

### Ions (Part C)

Determining the charges on ions follows a set of rules. A few of them appear here.

- 1) The maximum charge on a cation is its group number. E.g. for lithium in Group 1A, the charge is  $3+$ , for aluminum in Group IIIA, the charge is  $3+$ .
- 2) For main group elements (the "A" elements), if a second positive charge is possible, it is probably two smaller than the maximum positive charge. E.g. for tin (Sn)  $+2$  and  $+4$  exist.
- 4) The maximum negative charge on an anion is the group number  $-8$ . Thus, for chlorine in Group VIIA,  $7 - 8 = -1$ , for nitrogen  $5 - 8 = -3$ .
- 5) For transition metals, nearly all metals have a  $+2$  charge as one option. Many have  $+3$ .

Beyond that, like the main group elements the maximum charge equals the group number.

### 6.3 Writing Formulas from the Name of Ionic Compounds

In the last section, we covered writing the name of a compound from its formula, now we do the reverse process. Consider aluminum chloride. With a metal bound to a nonmetal, we automatically have half of the formula "Al." From Section 6.2, chloride is "Cl-." What is missing is the ratio of cations to anions. Aluminum is in Group IIIA and, absent any additional information, it is best to assume that the charge equals the group number, so  $3+$ . Since it is a compound, aluminum chloride is overall neutral (has no charge). This means that there must be three  $1-$  anions to offset the one  $3+$  cation, giving the formula  $AlCl_3$ .

Determining the formula can also be done via an equation. Using the same information from the last paragraph, we know that the formula for aluminum chloride is  $Al_xCl_y$ , where  $x$  and

$x$  and  $y$  are whole numbers. Thus,

$$x(3+) + y(1-) = 0$$

This is an equation with two unknowns, so it can't be solved, so the way to go forward is to assume either  $x$  or  $y$  equals 1. Pick the ion with the largest absolute charge and make its coefficient 1, then solve

$$1(3+) + y(1-) = 0$$

$$3+ + y(1-) = 0$$

$$y = 3$$

to give the formula  $\text{AlCl}_3$ .

While many compounds yield the formula directly by this method, it is also possible to have a more complicated situation, for example magnesium nitride. As previously, magnesium is in Group 2A, so it has a 2+ charge. Nitride is the nitrogen anion (Group 5A) so its charge is  $5 - 8 = 3-$ . Neither 2 nor 3 goes into the other a whole number of times, so the number of each ion is determined by the minimum number each required to balance the charges. For 2+ and 3-, that would be three of the magnesium ions and two of the nitride ions. Here  $3(2+) = 6+$  and  $2(3-) = 6-$ , which would yield a net zero charge for  $\text{Mg}_3\text{N}_2$ .

There is no formula for handling compounds like magnesium nitride, but it usually works to multiply the ion charges and then divide the product by the charge on each ion to determine how many of that ion appear in the compound. In the calculations, ignore the charges on the ions. For magnesium nitride that would be:

$$(2+)(3-) = 6-$$

$$\text{Mg: } (6-)/(2+) = 3$$

$$\text{N: } (6-)/(3-) = 2$$

This yields the formula we obtained earlier. There is a temptation to do calculation automatically, but it can lead to incorrect results. For example, if the metal cation is 2+ and the anion 4-, the salt would then have a formula of  $\text{M}_4\text{A}_2$  if you used this method. In fact, there is a common divisor of "2" which would simplify this formula to  $\text{M}_2\text{A}$ , which is what you would get if you tried to balance the charges as in the example of aluminum chloride

earlier.

#### 6.4 Naming Binary Compounds

The name of this section may confuse you because we touch on this briefly in Section 6.2, but we only scratched the surface of naming compounds. In general, naming compounds is more challenging than determining their formulas for a number of reasons. The primary one is that different types of compounds use different nomenclature systems. You have already seen that there are very old common names and systematic names for ionic compounds (salts). There are also acids and compounds between nonmetals. We will cover each type of compound in the coming pages.

##### Binary Compounds Containing a Metal and Nonmetal

At this point, please go back and reread Section 6.2. For ionic compounds (salts) composed of two elemental ions and the metal has only one possible charge, the name is just the metal name followed by the nonmetal name adjusted to include “-ide.” For example,  $\text{BaBr}_2$  is barium bromide. Table 6.2, p. 109, lists common elemental anions. When the metal has only more than one possible charge, the charge on the cation is provided in parentheses as a Roman numeral following the cation name. For example,  $\text{FeCl}_2$  is iron(II) chloride.

The book covers a second systematic method of naming ionic compounds that predates the system just described. For elements with two possible charges, the higher charge has the metal end in “-ic” while the lower charge ends in “-ous.” You are not responsible for knowing this system, but should be aware of its existence because much older science literature includes it and some of the terms are actually common in regular language. For example, you may see the terms ferric, ferrous, cupric, and/or cuprous in newspapers. If you ask your grandparents, they will likely remember that Crest<sup>®</sup> toothpaste used to advertise that it included stannous fluoride.

##### Binary Compounds Containing Two Nonmetals

The next major class of compounds with a distinct nomenclature system are binary inorganic molecules. These are composed of two different kinds of nonmetal atoms. Here, prefixes (mono-, di-, tri-, etc.) are used to indicate the number of each atom present. The name of the first

element is the one appearing to the left and/or below the second on the periodic table. (“Electronegativity” is the real selection criterion, but we won’t course that in this course. It maps well on the position of the elements on the periodic table.) The first name is the unaltered element’s name with a prefix to indicate how many atoms of that are present in the molecule (unless it is one atom, in which case the prefix is not used). The second element is altered to end in “-ide” as for salts, with the prefix added (this time including “mono-” for molecules with just one atom. A few examples illustrating these rules appear below.

CO – carbon monoxide

CO<sub>2</sub> – carbon dioxide

S<sub>2</sub>O<sub>3</sub> – disulfur trioxide

You should know the prefixes from one to ten. (margin table on p. 112)

## 6.5 Naming Compounds Containing Polyatomic Ions

While there are many salts that are simple, binary compounds, there are many more that include polyatomic ions. Polyatomic ions are just what the name suggests: the cation and/or anion is composed of more than one atom. In nearly all cases, there will be at least two elements in the ion (e.g. carbonate = CO<sub>3</sub><sup>2-</sup>), but there are a few single element polyatomic ions (e.g. azide = N<sub>3</sub><sup>-</sup>). Polyatomic ions are sometimes called “complex ions,” to differentiate them from the “simple” ions which are made of one atom.

It is likely that you have heard the names of lot of polyatomic ions from your experience around the house or work. For example, acetate, carbonate, bicarbonate (hydrogen carbonate), cyanide, hydroxide, nitrate, and phosphate are all polyatomic ions. Table 6.5 (p. 113) provides a list of 18 polyatomic ions, most of which you will see repeatedly in CHM 211/212. You should memorize their names, formulas, and charges.

Unlike the nomenclature of simple ions, naming polyatomic ions is not straightforward so you must memorize the names rather than trying to rationalize them in most cases. Thus, “cyanide” doesn’t tell you that it is an ion composed of carbon and nitrogen. However, sometimes there are hints relating to composition, e.g. permanganate has a manganese atom in it.

The one important situation where there is a systematic aspect to naming polyatomic ions are the HX anions, e.g.  $\text{HCO}_3^{2-}$ . In this case, it is hydrogen carbonate. When you see a hydrogen atom in the anion separate from a recognizable anion, the anion as a whole is named in this way. Another example,  $\text{H}_2\text{PO}_4^-$  is dihydrogen phosphate.

Finally, for the purposes of this course and CHM 211/212, there is only one polyatomic cation that matters,  $\text{NH}_4^+$ , ammonium. This cation is very important and you will see it many times in these courses.

## 6.6 Acids

Acids are a very important class of compounds and have their own nomenclature system. All acids will have “acid” explicitly in the name, so there is no ambiguity. There are two types of acids, binary/simple and complex, and the designations work the same way as for ionic compounds. It is important to remember that acids are molecular compounds, not ionic. On p. 116, there is a flowchart for naming binary compounds.

The binary acids of interest here are the halogen acids, HF, HCl, HBr, HI, and  $\text{H}_2\text{S}$ . Their names are based on the element to which the hydrogen atom binds: hydro[element]ic acid. Thus, HF is hydrofluoric acid, HCl is hydrochloric acid, and  $\text{H}_2\text{S}$  is hydrosulfuric acid. You won't see  $\text{H}_2\text{S}$  often, but remembering the name is important because it has a closely related acid that has a very similar name and which is very common.

Most acids involve hydrogen bound to a polyatomic ion (which forms a molecular compound). For example, nitric acid is  $\text{HNO}_3$ , which contains a hydrogen cation bound to a nitrate anion. We haven't gotten to bonding yet, but when this happens the compound that forms is a molecule, rather than an ionic compound.

All polyatomic acids encountered in this course, CHM 211 and 212 are “oxyacids” meaning that the acid will have one or more oxygen atoms in its formula,  $\text{HXO}_n$ , where  $n = 1 - 4$ . For many elements, X, there will be two associated oxyacids. Thus, there is nitric acid,  $\text{HNO}_3$ , and nitrous acid,  $\text{HNO}_2$ . Where there are two acids, the one with the larger number of oxygen atoms will end in “-ic acid” and the one with fewer will end in “-ous acid.” In nearly all cases, there is a one oxygen



atom differential between acids (i.e.  $\text{HNO}_2$  vs.  $\text{HNO}_3$ ). Table 6.8, p. 116, lists many common oxyacids. Memorize them.

There is a final, smaller grouping of oxyacids that come from Group 7A. They have 4 related formulas ( $\text{HXO}$ ,  $\text{HXO}_2$ ,  $\text{HXO}_3$ , and  $\text{HXO}_4$ , where  $\text{X} = \text{Cl, Br, I}$ ). The names are related to each other.

$\text{HClO}$  - hypochlorous acid

$\text{HClO}_2$  - chlorochlorous acid

$\text{HClO}_3$  - chloric acid

$\text{HClO}_4$  - perchloric acid

The brom and iodo acids are named analogously. Fluorine does not form an oxyacid.

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