Chapter 4 – Nomenclature

> I have mentioned this before (and will continue) but now is the time to start **LEARNING** *Table 4.2 on page 129* even if we have not yet gotten to it. <

The Periodic Table – Review from Chapter 2

The elements are arranged in rows and columns on the Periodic Table according to similarities in their properties. We are going to talk about the development of the modern Periodic Table, because there are important features that you need to understand now in order to be able to think about what this all means and then into formulas and names as we work through the chapter.

Section 4-1 – The Origin of the Periodic Table

There are a large number of people who have contributed (and are contributing) to the development of te Period Table. We are only going to talk about a few of them.

1) Sec 4-1.1 – The Origin of the Periodic Table

{Note: Much of this first part is NOT in your book. I think it is important that we go beyond the simple in our thinking about the periodic table, because it did not just appear to one person magically or overnight, but developed over time as people understood more and more, and they tried to explain things from their understanding.}

History of the Development of the Periodic Table

Similarities between the chemical and physical properties of the elements were known early in the nineteenth century. In 1817, Johann W. Döbereiner showed in a number of instances that when three elements with similar properties were listed in order of increasing atomic mass, the middle element had properties that were an approximate average of the other two. His groups of "triads" from top to bottom include:

Li	Ca	N	As	S	Cl
Na	Sr	P	Sb	Se	Br
K	Ba	As	Bi	Te	I

Each triad is a part of a family (or group) of elements in the modern periodic table.

In 1866, John Newlands established the "Law of Octaves." Being a lover of music, he theorized that when the elements were arranged in order of increasing atomic mass, every eighth element had similar chemical and physical properties, just like every eighth note on the musical scale has a similar sound! For the lighter elements, this periodicity is surprisingly valid, especially when the noble gases and the transition elements are not considered. In 1866, scientists were unaware of the noble gases.

Musical Scale	Do	Re	Mi	Fa	So	La	Ti
Element Scale	Li	Be	B	C	N	O	F
	Na	Mg	Al	Si	P	S	Cl

It remained until 1869 for two scientists, Dmitri Mendeleev from Russia and Lothar Meyer from Germany, to independently establish a periodic table similar to our present-day arrangement of the elements. Mendeleev showed that with the elements arranged in order of increasing atomic mass, their *chemical* properties occur periodically. When Meyer arranged the elements in order of increasing atomic mass, he found that their *physical* properties recur periodically. The two tables, however, were virtually identical. Because he published his table in 1869 (Meyer did not publish until 1870) and because his table included "blanks" for yet-to-be-discovered elements to fit, Mendeleev is considered the "father of the modern periodic table."

There were places where Mendeleev's table did not seem to match the data, particularly with atomic masses and where the elements should be placed. In some cases he answered the question with his blanks, and in others he simply stated that the atomic mass data must be wrong. It was a few more years before part of this was taken care of.

In 1913, H.G.J. Moseley's study of the X-ray spectra of the elements refined the periodic table to its current status: **when the elements are arranged in order of increasing atomic *number* (ie., the number of protons in the nucleus), certain chemical and physical properties repeat periodically.** This is what we now refer to as the **PERIODIC LAW.**

2) <u>Sec 4-1.2 – Metals and Nonmetals</u>

a) <u>Metals</u> – are elements that have a high luster when clean and a high electrical conductivity. They are <u>ductile</u> (can be drawn into a wire) and <u>malleable</u> (can be beaten into thin sheets). Most of the elements are metals. Metals also have the characteristic that they lose electrons when they form compounds.

b) <u>Nonmetals</u> – are elements that are non lustrous and are poor conductors of electricity, and have more variability in their properties. Nonmetals will gain electrons when they form many compounds, they also have the ability to share electrons.

c) <u>"Semi metals" or Metalloids</u> – are elements with the properties of both metals and nonmetals. (They are semiconductors of electricity.) Although they may be confused about their properties and we have a special name for them, they are ultimately either metals or nonmetals.

Metals do not all behave the same way. Some (like platinum, gold and silver) are very resistant to the chemical reactions such as corrosion and rust. These kinds of metals are sometimes referred to as the *noble* metals. Others, like lithium, sodium and potassium, have to be stored carefully (often under oil) because the react readily and violently with air and water, and are generally called *active* metals. We will see some of these differences later when we talk about reactions and activity.

<u>Section 4-2 – Using the Periodic Table</u>

3) <u>Sec 4-2.1 – Periods</u>

<u>Periods</u> – The horizontal rows of the periodic table are called periods and we will learn

that there are progressions as one moves across a row in the Periodic Table. We can notice at this point that the atomic numbers (the numbers of protons in the nuclei) increase as we move across the period and, generally, the atomic masses increase.

4) Sec 4.-2.2 – Groups

Groups or families – The vertical columns are called groups or families. Elements in a given group have similar chemical and physical properties.

a) Representative Elements – Groups 1A through 8A (groups 1,2, 13-18) are called the *representative elements*. Sometimes these are referred to as the Group A elements. These elements behave as we would expect them to if we were to be able to completely predict behavior. These groups are easy to identify because the groups are usually shown to "stick up" above the others.

   Four of the groups of representative elements have special names:
   i) Group 1A (or 1) elements are called the **alkali metals**.
   ii) Group 2A (or 2) elements are called the **alkaline earth metals**.
   iii) Group 7A (or 17) elements are called the **halogens**.
   iv) The elements in group 8A (or 18) are the **noble gases**. The noble gases will not form ions. It is thought that they have ideal electronic structures.

   The other groups of representative elements are not given special names.

b) Transition Elements – The other elements on the main part of the table (so-called Group B elements or groups 3 through 12) are called the *transition elements or metals*. These behave less predictably than the representative elements (although we will learn more details about their behaviors as time goes on). These are easy to identify because they are "in the valley" as you look at the Periodic Table. In this group are such elements as iron and chromium, as well as gold, silver and copper.

c) Inner Transition Elements – The elements at the bottom of the Period Table are most often called the *inner transition elements or metals*. They behave like the transition elements but their columns do not have numbers when we talk about groups. Although the groups do not have specific name, their parts of the periods do. The first row of inner transition elements are the **lanthanides** and are sometimes referred to as the *rare earth* elements or metals, and the second row are the **actinides**. All of the inner transition elements are radioactive.

d) Metals and Nonmetals – On the Periodic Table in our room, the separation between metals and nonmetals is the red stair-step line. In your book it is a thick, black line. All Periodic Tables somehow indicate this separation.

5) Sec 4-2.3 – Physical States and the Periodic Table

The majority of the elements on the Periodic Table are solids under standard conditions which most often is considered to be room conditions. Room temperature, which is defined in chemistry as 25 °C, is the standard reference temperature used to describe physical state. On the Periodic Table in our room the symbols for the solids are colored **BLACK**. Other elements are gases at room conditions and these are **RED**. There are two elements, Mercury and Bromine, that are liquid at room conditions and these are **BLUE**.

Note: Some versions of the Periodic Table also list Cesium as a liquid but, strictly speaking, it melts at about 30 °C and so is not a liquid at room temperature.

There are seven (7) elements that appear in nature as two atoms when they are elements. These are referred to as the *diatomic elements* or *diatomic molecules*. They are: Hydrogen - H_2, Nitrogen - N_2, Oxygen - O_2, Fluorine - F_2, Chlorine - Cl_2, Bromine - Br_2, and Iodine - I_2. I often refer to these as the "magnificent 7" – in part because, except for hydrogen, they form a 7 as you look at how they are located on the Periodic Table. (This does not mean that they have to be in pairs in compounds but we will get to that.)

Two other non-metals also have special forms. Phosphorus is often viewed as having an elemental formula of P_4 and Sulfur is S_8. At times chemists will also use these two in their single atom forms.

Section 4-3 – Naming and Writing Formulas of Metal-Nonmetal Binary Compounds

Compounds are pure substances that differ from elements because they contain more than one kind of atom. Compounds are formed when atoms of two or more different elements combine chemically. In any compound, the elements are always present in the same proportion by mass (Law of Definite Proportions).

There are roughly two kinds of chemical compounds – organic and inorganic – although some call any compound a molecule. We are going to refer to these as *molecular* and *ionic* compounds. This alternative naming helps us to recognize the kinds of bonds that are present and the units that make them up.

Organic compounds are mainly made up of carbon, hydrogen, and oxygen. They are the compounds of life. The smallest unit is the *molecule*. All other compounds are called inorganic compounds. Their smallest unit is called a *formula unit*.

Some compounds have common names that you just have to learn, while others are more straight-forward and predictable. We will start with inorganic compounds.

6) Sec 4-3.1 – Metals with Ions of Only One Charge

Although noble gases do not form ions, the elements on either side of the noble gases will either gain or lose electrons to have the same number of electrons as a noble gas.
This explanation does not hold for all the representative or the transition metal ions. We will explore these later in Chapter 8.

The representative metals form specific ions.

Group 1A (1) alkali metals form **+1** ions. *Note: These do not include hydrogen. Although it is in the same group, it is not a metal.*
Group 2A (2) alkaline earth metals form **+2** ions.
Group 3A (13) metals form **+3** ions. (Some more than one ion.)

When nonmetals are present in ionic compounds, they also form only one negative charge. {As you look at the Periodic Table in our classroom do not be confused by the extra numbers by the symbols - nonmetals only form one **ion**.}

Group 5A (15) non-metals form **–3** ions.
Group 6A (16) non-metals form **–2** ions.
Group 7A (17) halogens form **–1** ions.

Positive and negative ions do not exist separately. They are always found together in ordinary matter. Likewise, a cation cannot form (lose one or more electrons) unless there is a nonmetal present (to gain the one or more electrons).

When metals form only one ion, the name of the ion is just the name of the metal plus the word **ion**. So sodium (Na) loses 1 electron and becomes sodium ion (Na^+), and calcium (Ca) loses 2 electrons and becomes calcium ion (Ca^{2+}).

<u>Example</u> – How many electrons does one atom of the following lose to form the ion and then how many electrons does the ion contain?

1) K^+ 2) Ba^{2+} 3) Al^{3+}

Nonmetals, on the other hand, are different. The names of the anions that are formed from the nonmetals differ from that of the neutral nonmetal. The suffix **-ide** is added to the name of the nonmetal when it is an anion. (Example: Cl^- is called **chloride** instead of chlorine when it gains 1 electron, and O^{2-} is oxide and not oxygen when it gains 2 electrons. We talked about this in an earlier chapter.)

<u>Example</u> – How many electrons does one atom of the following gain to form the ion and then how many does the ion contain, and what is the name of each of the ions?

1) I^- 2) S^{2-} 3) N^{3-}

When the ions come together to form compounds, chemists just combine the names of the ions (of course, dropping the "ion" from the cation). Thus sodium ion and chloride come together to form sodium chloride or NaCl. Likewise, calcium ion and oxide form calcium oxide or CaO.

*It will be important for you to remember that the **ide** ending is associated with simple nonmetals when they are in compounds!*

<u>Example</u> – What is the name of the following compounds?

1) KI 2) MgS 3) AlN

4) CaF_2 5) Na_2O 6) Al_2O_3

However, it is a different task when one writes a formula from a name because we must first determine the number of atoms of each element present in the formula.

7) <u>Ionic Compounds from Ions with Unequal Charges</u>

Compounds are neutral, and ionic compounds are no exceptions. When the ions come together the charges must balance in the compound.

Thus, in example 4 above it takes 2 fluorides (F^-) to balance the charge of the 1 calcium ion (Ca^{2+}). Likewise, to balance charges, the formula for the compound formed from aluminum ion (Al^{+3}) and oxide (O^{2-}) is Al_2O_3.

There are all sorts of methods to determine a formula. One is to look at the common multiple between the charges on the ions. That is Al is +3 and O is –2, so the multiple is 6. This means that we need to account for 6 electrons. Thus, 2 Al will give up 6 electrons and it will take 3 O to accept them all.

An easier method to find a formula if the charges of the ions are known is to use a *cross-over* or *crisscross* method. This method always works. The only check that one needs to do is to make sure that the subscripts are in the lowest ratio that they can be.

<u>Example</u> – For calcium oxide one combines calcium ion (Ca^{2+}) with oxide (O^{2-}), the cross-over method would give Ca_2O_2, but this needs to be reduced to CaO.

A <u>Formula Unit</u> is the lowest whole-number ratio of ions in an ionic compound.

When we name these compounds we do not need to worry about the numbers of each of the ions present, only their names.

8) <u>Sec 4-3.2 – Metals with Ions of More than One Charge</u>

Except for Groups 1A, 2A, aluminum and some of the other elements, other metals can form more than one cation. What this means is that a name like iron chloride does not mean much since iron can have two ions, Fe^{2+} and Fe^{3+}.

To distinguish between compounds containing either of these so that everyone knows what we are talking about, we use the **Stock** method. In the ***Stock method*** or ***systematic naming system***, a Roman numeral in parenthesis is used after the name of the element to indicate the numerical value of the charge. (Example: Fe^{2+} is named iron(II) and Fe^{3+} is named iron(III).)

The Roman numeral always tells the charge on the metal ion! It never indicates how many atoms of the metal are in the formula.

If you have the name tin(II) fluoride, this means Sn^{2+} combines with F^- which gives SnF_2. And, tin(IV) fluoride is SnF_4.

9) If you have a formula you can come up with the proper name of a compound just as we did earlier. I think the most convenient means is to “uncross” the formula to be able to name the compound. This gives us a way to determine the charge of the metal ion so that we can properly name it. (Your book suggests using an “algebraic equation” method. Both really do the same thing.)

<u>Example</u> – The element Manganese can have 5 (yes, 5) possible ions. If you have the formula MnO to name, you can uncross the formula and you would get Mn^{+1} and O^-. What you should *immediately* notice is that this is not correct! It is not correct because oxide is ***ALWAYS*** a –2 ion. What this means is that the manganese must be a +2 ion for the charges to balance. Knowing this, we would call MnO manganese(II) oxide. (*Remember*

that formula units always the smallest ratio for the ions present, so the formula may have been reduced when it was first written.)

What would you name the following compounds?

1) Mn_3O_4 2) Mn_2O_3

3) MnO_2 4) Mn_2O_7

10) <u>Sec 4-3.3 – The Classical Method for Metals with More than One Charge</u>

An older, less preferred, **classical method** of naming these ions uses a root word and suffixes. If the source of the symbol for the metal is the classical (Latin) name, then that is used as the root for the name, otherwise the English root is used. An ***-ous*** ending is used for the ion with the lower charge and an ***-ic*** ending is used for the ion with the higher charge. (Example: Fe^{+2} is *ferrous* and Fe^{+3} is *ferric*.) **Table 4-1** on page 127 lists the more common metals that have both classical and Stock names for their ions.

We need to recognize this classical naming system because often the most commonly used names for some compounds are these older names. Pharmaceutical and drug companies, and chemical suppliers continue to use these classical names.

<u>Example</u> – what are the classical names for SnF_2 and SnF_4?

<u>Section 4-4 – Naming and Writing Formulas of Compounds with Polyatomic Ions</u>

When two or more atoms are covalently bonded in a group (electrons are shared) and the group is ionic (either lost or gained electrons), the group is called a ***polyatomic ion***. A list of some of the more common polyatomic ions is shown in **Table 4-2** on page 129 in your book. Notice that all but one of the ions (ammonium, $NH_4^{\ +}$) are anions.

11) <u>Sec 4-4.1 – Oxyanions</u>

The list in Table 4.2 may seem imposing but there are some ways to make it easier. In many cases, *the anions are made up of oxygen and one other element*. These ions are referred to as ***oxyanions***.

When two oxyanions have the same elements (such as $SO_3^{\ 2-}$ and $SO_4^{\ 2-}$) as you have seen by now, they must have different names so that we can tell them apart. The anion with the smaller number of oxygens uses the root of the *element* plus the ending ***-ite***. The one with the larger number of oxygens uses the root of the element plus ***-ate***. Therefore, $SO_3^{\ 2-}$ is called *sulfite* and $SO_4^{\ 2-}$ is *sulfate*.

There are some of the ions that have more than two possibilities in Table 4-2. For example there are four that contain chlorine. The middle two have the -*ite* and -*ate* endings for *chlorite* and *chlorate*, respectively. The one with one less oxygen than chlorite has the prefix *hypo* meaning “under” for hypochlorite, and the one with one more oxygen than chlorate has the prefix *per* from “hyper” for perchlorate.

You will also note that there are a couple that contain two elements but are a bit different. Two such are CN^- and OH^-. Both of these use the –*ide* ending, so the former is *cyanide ion* and the latter is *hydroxide ion*.

Most ionic compounds are also referred to as salts. *A **salt** is an ionic compound formed by the combination of a cation with an anion.*

Although polyatomic ions have more than one atom in them, the whole ion is treated as one thing just like the simple ions.

12) Sec 4-4.2 – Naming and Writing the Formulas of Salts with Polyatomic Ions

In naming and writing the formulas of compounds containing polyatomic ions we follow the same procedures as with metal–nonmetal compounds.

For $Cr_2(SO_4)_3$ the name is *chromium(III) sulfate*

And when we go from names to formulas: ammonium phosphate is $(NH_4)_3PO_4$ and we remember to use parentheses when we need more than one polyatomic ion in the formula.

Remember:

a) Inorganic or ionic compounds are made up of ions.
b) The charges on the cations and anions in the compound must balance.
c) The formulas for inorganic or ionic compounds are written with subscripts to indicate the numbers of each ion present in the formula to get the charges to balance.
d) The names of inorganic or ionic compounds are written just by naming the ions present.
e) In either the formula or name the metal comes first.
f) If a metal can have more than one possible ion, the name of the ion is the name of the metal followed by the ion charge written in Roman numerals in parentheses.
g) In some cases, an older naming system might be used where the ion with the lesser charge has a name ending in *-ous* while the higher charge ion ends in *-ic*.
h) Polyatomic ions are treated like other cations and anions in formulas and names.
i) Polyatomic ions have their own naming characteristics.

Section 4-5 – Naming Nonmetal-Nonmetal Binary Compounds

13) Sec 4-5.1 – Writing the Formulas of Binary Molecular Compounds

When a metal is combined with a nonmetal it is easy to decide which name or symbol comes first – it is always the metal.

When we have two nonmetals, we write the symbol of the element that is closest to being a metal first – that is, the nonmetal that is closest to the metal-nonmetal border (farther down or farther to the left). So we get CO_2 instead of O_2C, but OF_2 rather than F_2O.

When hydrogen as a nonmetal is combined with nonmetals it is written first. When there is a metal present with the hydrogen and other nonmetals it is written second.

(*Note: In organic compounds containing C, H, and other elements, the C is written first, followed by H and then other elements often in alphabetical order.*)

14) Sec 4-5.2 – Naming Binary Molecular Compounds

Two characteristics of binary molecular compounds affect how they are named. First, because they are composed of molecules, the ionic charges are not used in writing the formulas for these compounds. Second, when two nonmetallic elements combine, they often do so in more than one way. (Example: CO and CO_2)

Prefixes are used to show how many atoms of each element are present in each molecule of a binary molecular compound. The prefix is normally omitted if there is just a single atom of the first element in the name.

Prefixes Used: mono (1), di (2), tri (3), tetra (4), penta (5), hexa (6), etc. (See Table 4-3 for more.)

Binary Molecular Formulas and Names – These are really very simple in that both are "what you see is what you get." If you have a formula, what elements are present and how much of each? If you have a name, what information does it give you?

Examples

1) Name these compounds:
   a) CBr_4 b) N_2O_5 c) BCl_3

2) Write formulas for these compounds:
   a) nitrogen dioxide b) carbon tetrachloride c) diphosphorus trioxide

There are some exceptions to the "rules" we have discussed. It is often easier to learn the exceptions than to try to modify the rules. Sometimes formulas and names can be done multiple ways, and all are correct.

15) Sec 4-5.3 – Naming Alkanes

There is a special class of binary molecular compounds that contain only carbon and hydrogen, and these compounds are referred to as ***hydrocarbons***. These compounds are very important in our lives. They serve as the fuels that power our cars, heat our homes, and cook our food. One specific class of hydrocarbons is distinguished by its relative number of carbons to hydrogens. These hydrocarbons are called **ALKANES**. Alkanes have the general formula C_nH_{2n+2}. In other words, for the number of carbon atoms in the molecule, there are double that number of hydrogen atoms plus two. Alkanes make up some of the best-known fuels that we use.

As you can see from the following list (similar to Table 4.5 on page 134), the simple straight-chained alkanes are not named the way we have named other molecular compounds, they have their own naming system. (This is, in part, because they fall into the broader class of compounds in Organic Chemistry which we will learn more about in Chapter 17.) However, they do follow predictable patterns like other compound names.

Name	Formula
Methane	CH_4
Ethane	C_2H_6
Propane	C_3H_8
Butane	C_4H_{10}
Pentane	C_5H_{12}
Hexane	C_6H_{14}
Heptane	C_7H_{16}
Octane	C_8H_{18}
Nonane	C_9H_{20}
Decane	$C_{10}H_{22}$

The first four have their own distinctive names but after that they are named for the prefix (Table 4.3) that tells the number of carbon atoms in the chain and the *-ane* ending which describes the particular carbon-hydrogen relationship.

We will get to this in much more detail when we get to Chapter 17.

16) A summary of Ionic and Molecular Compounds

Molecular Compounds – Compounds that are composed of molecules.

1) They tend to have relatively low melting and boiling points. (That is, most are either gases or liquids at room temperature.)

2) Most are composed of two or more nonmetallic elements.

3) They are held together by molecular or covalent bonds.

Ionic Compounds – are composed of positive and negative ions. These compounds do not exist in any unit or group (do not exist as molecules). Ionic compounds are electrically neutral although they are composed of ions.

1) Most ionic compounds are crystalline solids at room temperature.

2) They are usually formed from at least a metallic and a nonmetallic element.

3) They usually have high melting and boiling points.

4) The are held together by ionic bonds.

Chemical Formulas show the kinds and numbers of atoms in the smallest representative particle or unit of the substance.

1) A Molecular Formula shows the number and kinds of atoms present in a molecule of the compound.

   Example – Give the molecular formula of butane if it contains 4 carbon atoms and 10 hydrogen atoms.

2) A Formula Unit is the lowest whole-number ratio of ions in an ionic compound.

   Example – Give the simplest formula unit for calcium chloride if it contains one Ca^{+2} ion for each two Cl^{-1} ions.

   Example – Give the simplest formula unit for ferric bromide if it contains 1 iron atom and 3 bromine atoms.

Section 4-6 – Naming Acids

When hydrogen is combined with a simple anion it forms a compound that is molecular rather than ionic. When dissolved in water, however, the molecule is ionized to form H^+ ions and the anion. Most of the hydrogen compounds formed from the representative nonmetals and the polyatomic anions behave the same way, at least to some extent. *This common property of forming H^+ in aqueous solution is a property of a class of compounds called **acids**.* Acids are important enough to earn their own nomenclature.

16) Sec 4-6.1 – Binary Acids

The acids formed from the simple anions are composed of hydrogen plus one other element, so they are called **binary acids**.

When the anion ends in *-ide*, the acid name begins with the prefix ***hydro-***. The stem of the ion has the suffix ***-ic*** and it is followed by the word ***acid***. (Example: HCl is named as hydrochloric acid when it is a water solution.) This holds true for the polyatomic ions that also end in *-ide*.

When these acids are not in water they are just named as compounds of the anion. (Example: HCl is named as hydrogen chloride.)

Some hydrogen compounds of anions are not generally considered to be binary acids: H_2O, NH_3, CH_4, and PH_3.

17) Sec 4-6.2 – Oxyacids

The acids formed from most of the polyatomic anions in Table 4-2 are known as **oxyacids** because they are formed from oxyanions. Note also that there have to be enough H^+ ions to balance the charge of the anion.

When the polyatomic anion (X) ends in -ite, the acid name is the stem of the anion followed with the suffix ***-ous***, followed by the word acid. (Example: H_2SO_3 is named as sulfurous acid.)

When the anion (X) ends in -ate, the acid name is the stem of the anion followed with the suffix ***-ic***, followed by the word acid. (Example: HNO_3 is named as nitric acid.)

18) A Summary of Acid Naming

HCl —> chloride —> --chlor-- —> *hydro*chlor*ic* acid
HClO —> hypochlorite —> hypochlor-- —> hypochlor*ous acid*
$HClO_2$ —> chlorite —> chlor- —> chlor*ous acid*
$HClO_3$ —> chlorate —> chlor- —> chlor*ic acid*
$HClO_4$ —> perchlorate —> perchlor-- —> perchlor*ic acid*

19) Example of Acidss

1) Name these compounds as compounds and as acids:
   a) HF b) H_2SO_4 c) HNO_3

2) Write formulas for the following acids:
   a) chromic acid b) hydrobromic acid c) chlorous acid

20) Writing Formulas from Names – It is important for all that we will be doing from now on that you can write formulas from names and, of course, also names from formulas.

You should review the following:

Sodium Sulfide	Na_2S
Stannous Phosphate	$Sn_3(PO_4)_2$
Carbonic Acid (carbonate)	H_2CO_3
Calcium Iodide	CaI_2
Iron(II) Oxide	Fe_2O_3
Sodium Sulfate	Na_2SO_4
Copper(II) nitrate	$Cu(NO_3)_2$
Carbon Tetrachloride	CCl_4
Diphosphorus Pentoxide	P_2O_5