Experiment 10

Chemical Names and Formulas: A Study Assignment

Performance Goal

10–1 Within the limits discussed in this exercise, and using a periodic table for reference, given the name (or formula) of any chemical species among the classifications below, write the formula (or name):

Elements in their stable form
Molecular binary compounds
Binary acids; oxyacids
Monatomic ions; polyatomic ions
Ionic compounds

INTRODUCTION

This study assignment presents a brief summary of the rules for writing formulas and naming substances commonly encountered in an introductory chemistry course. Basic definitions are stated, but theory relating to chemical bonding and the formation of ions is not considered. The purpose of this exercise is to practice writing formulas and names with help immediately available to clear up points that you may not understand. Hopefully you will *master* formula writing techniques during this laboratory period.

CHEMICAL OVERVIEW

Elements

This discussion will be limited to the more common elements listed in Figure 10–1. Given the name of one of these elements, you should be able to write its symbol, using a full periodic table for reference; given the symbol, you should be able to identify the element by name. This requires a certain amount of memorization, but

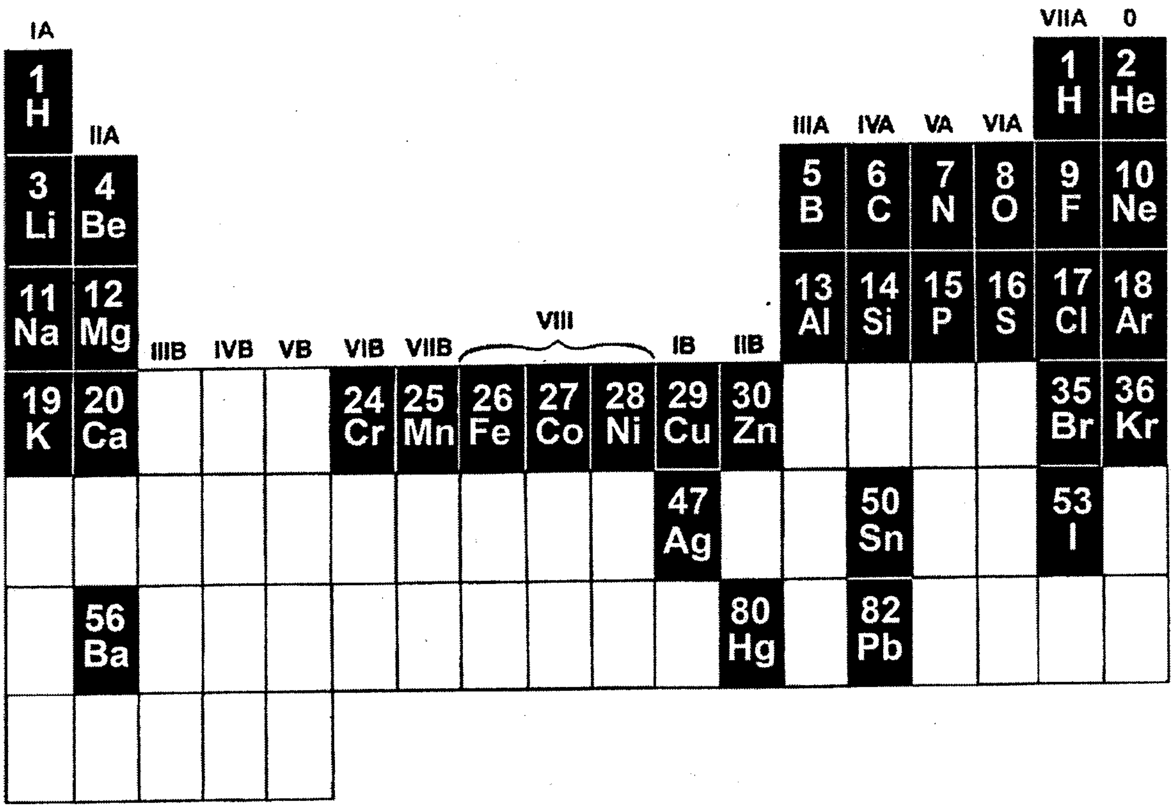

Figure 10–1. Partial periodic table showing the symbols and locations of the more common elements. The symbols above and the list that follows identify the elements you should be able to recognize or write, referring only to a complete periodic table. Associating the names and symbols with the table makes learning them much easier. The elemental names are:

aluminum	**bromine**	**chromium**	**iodine**	**magnesium**	**nitrogen**	**silver**
argon	**calcium**	**copper**	**iron**	**manganese**	**oxygen**	**sodium**
barium	**carbon**	**fluorine**	**krypton**	**mercury**	**phosphorus**	**sulfur**
beryllium	**chlorine**	**helium**	**lead**	**neon**	**potassium**	**tin**
boron	**cobalt**	**hydrogen**	**lithium**	**nickel**	**silicon**	**zinc**

the sheer memory work is lessened if you relate elemental names and symbols to the periodic table.

Generally the chemical formula of an element in its stable form at room conditions is simply the symbol of the element. Seven gaseous elements, however, are not stable as individual atoms; two atoms combine to form a *diatomic molecule* as the unit particle of each of those elements. These elements and their correct chemical formulas are nitrogen, N_2; oxygen, O_2; hydrogen, H_2; fluorine, F_2; chlorine, Cl_2; bromine, Br_2; and iodine, I_2. Always remember to include the subscript 2 when writing the formulas of these substances *as elements, uncombined with any other elements*. In particular, notice that *this has nothing to do with these elements as they exist in compounds*. This is evident in the formulas for water, H_2O, and dinitrogen trioxide, N_2O_3.

Molecular Binary Compounds

Compounds made up of atoms held together entirely by covalent bonds are called **molecular compounds.** When a compound consists of two kinds of elements, it is called a **binary compound.** Molecular binary compounds therefore consist of two elements held together by covalent bonds. These elements are generally *both non-metals*. You may use this identification feature to distinguish molecular binary compounds from ionic binary compounds that will be discussed shortly.

Molecular binary compounds are identified by names consisting of two words. The main part of the first word is simply the name of the element appearing first in the formula; the main part of the second word is the name of the element appearing second in the formula, modified by an *-ide* suffix. The other part of each word is a prefix indicating the number of atoms of that particular element in the molecule. This is illustrated in the name dinitrogen trioxide for N_2O_3, in which *di-* is the prefix for 2 and *tri-* is the prefix for 3. A list of prefixes for numbers from 1 to 10 is given in Table 10–1. When the molecule contains only one atom of an element, the prefix *mono-* is frequently omitted, unless the species named is one of two or more compounds formed from the same two elements, such as CO, carbon monoxide, as compared to CO_2, carbon dioxide.

Acids

Inorganic acids, and some organic acids, are compounds that yield a hydrogen ion, or proton, when they ionize. (A proton and a hydrogen ion are the same thing. A hydrogen atom consists simply of a proton and an electron. When the electron is removed, producing a hydrogen ion, the only thing left is the proton.) Formulas of such acids are written with the ionizable hydrogen appearing first. This feature can usually be used to identify a formula as that of an acid.

A **binary acid** consists of hydrogen and one other nonmetallic element, usually in water solution. A binary acid is named by surrounding the root of the nonmetal with the prefix *hydro-* and the suffix *-ic*. Thus HCl is hydrochloric acid, the *chlor*

Table 10–1. Prefixes Used in Naming Covalent Binary Compounds

mono-	= 1	hexa-	= 6
di-	= 2	hepta-	= 7
tri-	= 3	octa-	= 8
tetra-	= 4	nona-	= 9
penta-	= 5	deca-	= 10

coming from chlorine. The name hydrosulfuric acid suggests that the element other than hydrogen is sulfur. Its formula is H_2S.

Oxyacids contain oxygen as well as hydrogen and another nonmetal. The name of the most common oxyacid of each nonmetal is the root of the nonmetal followed by *-ic*. Thus H_2SO_4 is sulfuric acid, and the formula for the common oxyacid of chlorine, called chloric acid, is $HClO_3$. These names and formulas are somewhat similar to the names and formulas of the hydro-ic acids. Catch the distinction: *hydro-ic* acids have no oxygen, whereas *-ic* acids do contain oxygen.

There are six so-called *-ic* acids whose names and formulas you should memorize, as they constitute the base from which we will develop our approach to learning the names and formulas of a large number of chemical compounds. If you memorize these six acids, plus some prefixes and suffixes, you will be able to figure out all the other names and formulas without further memorization. The six acids are:

chloric, $HClO_3$
sulfuric, H_2SO_4
nitric, HNO_3
carbonic, H_2CO_3
phosphoric, H_3PO_4
acetic, $HC_2H_3O_2$

Acetic acid is the best known of a large group of organic acids that contain hydrogen but ionize only slightly in water. Organic chemists write the formulas for such acids differently, but for the purpose of this exercise we will follow the usual procedure of writing the ionizable hydrogen first.

The number of oxygen atoms may vary in oxyacids of the same nonmetal. Chlorine, for example, forms four oxyacids; $HClO_4$, $HClO_3$, $HClO_2$, and $HClO$. The names of these compounds are distinguished from each other by a series of prefixes and suffixes that are explained in Table 10–2. The key to the entire nomenclature

Table 10–2. Names of Oxyacids and Oxyanions of Chlorine (HCl included for comparison)

I Acid Name	*II* Acid Suffixes and Prefixes	*III* Acid Formula	*IV* Oxygens Compared to -ic Acid	*V* Ion Name	*VI* Ion Suffixes and Prefixes	*VII* Ion Formula
hydrochloric (binary acid)	*hydro-ic*	HCl	no oxygen	chloride	*-ide* named as monatomic anion	Cl^-
hypochlorous	*hypo-ous*	$HClO$	−2	hypochlorite	*hypo-ite*	ClO^-
chlorous	*-ous*	$HClO_2$	−1	chlorite	*-ite*	ClO_2^-
CHLORIC	*-IC*	$HClO_3$	SAME	CHLORATE	*-ATE*	ClO_3^-
perchloric	*per-ic*	$HClO_4$	+1	perchlorate	*per-ate*	ClO_4^-

system is the number of oxygen atoms *compared to the number in the -ic acid.* Study Table 10–2 to help you memorize these prefixes and suffixes and understand their use.

Nonmetals of the same chemical family frequently form acids that are similar in name and formula. Among the halogens, for example, HCl is hydrochloric acid, HF is hydrofluoric acid, HBr is hydrobromic acid and HI is hydroiodic acid. The similarities extend to oxyacids for bromine and iodine, but not for fluorine, which forms no oxyacids. We thus find that $HBrO_2$ is bromous acid, and HIO_4 is periodic acid.

Aside from the halogens, only sulfur and nitrogen form important oxyacids other than their well known *-ic* acids. In both cases it is the *-ous* acid that is formed, each with one less oxygen atom than is present in the *-ic* acid. Thus HNO_2 is the formula for nitrous acid, and sulfurous acid has the formula H_2SO_3. Selenium and tellurium, atomic numbers 34 and 52, in the same column of the periodic table as sulfur, form corresponding *-ous* acids.

Oxidation State: Oxidation Number

Chemists use a set of **oxidation numbers,** or consider the **oxidation state** of an element, in discussing oxidation–reduction reactions. These numbers are also part of the modern nomenclature system. The rules by which these numbers are assigned are as follows.

1. The oxidation number of any elemental substance is zero.
2. The oxidation number of a monatomic ion is the same as the charge on the ion.
3. The oxidation number of combined oxygen is -2, except in peroxides (-1) and superoxides ($-\frac{1}{2}$). We will not encounter peroxides or superoxides in this assignment.
4. The oxidation number of combined hydrogen is $+1$, except in hydrides (-1).
5. In any molecular or ionic species, the sum of the oxidation numbers of all atoms in the species is equal to the charge on the species.

The manner in which these rules are applied will be discussed as the need arises.

Monatomic Ions

A monatomic ion is a single atom that has acquired an electrical charge by gaining or losing one, two, or three electrons. Its formula is the symbol of the element followed by a superscript indicating the charge. For example, the formula of a calcium ion is Ca^{2+}, and for a chloride ion, Cl^-. It is important that the charge be indicated for an ion. Without that charge, the formula would be that of an electrically neutral atom from which the ion was formed, a very different species with very different chemical properties. Ions with a negative charge are called **anions;** ions with a positive charge are called **cations.**

The nonmetals in Groups VA, VIA, and VIIA form monatomic anions by gaining electrons. Ions from Group VA elements have a 3− charge, as in N^{3-}; from Group VIA, a 2− charge, as in O^{2-}; and from Group VIIA, a 1− charge, as in F^{-}. The name of a monatomic anion is simply the name of the element, modified by an *-ide* suffix, as in nitride, oxide, or fluoride.

Metals in Groups IA, IIA, and IIIA form cations with charges of 1+, 2+ and 3+, respectively. Many metals in the B groups of the periodic table form two monatomic ions that differ in charge. The best example is iron, which yields the Fe^{2+} and Fe^{3+} ions. These ions are distinguished by adding the oxidation state, or charge, to the name of the element. Accordingly, Fe^{2+} is the iron(II) ion, and Fe^{3+} is the iron(III) ion. Notice how these names are written; the oxidation state is written in Roman numerals *and enclosed in parentheses* immediately after the name of the element, with no space between the name and the parentheses. **Caution:** Students often neglect to enclose the oxidation state in parentheses; the name is not correctly written if the parentheses are missing. The names of iron(II) and iron(III) ions are pronounced "iron two" and "iron three" respectively.

Notice that oxidation states in the names of monatomic ions are used only to distinguish between ions of the same element that have different charges. Oxidation numbers are not commonly used if a metal forms only one kind of ion. The two monatomic ions of copper are an exception to this. The copper(II) ion, Cu^{2+}, is so much more common than the copper(I) ion, Cu^{+}, that the name "copper ion" is understood to apply to Cu^{2+}. Copper(I) must be used to identify the Cu^{+}; and you are always correct if you use copper(II) for Cu^{2+}.

The cations formed by mercury require special comment. The mercury(II) ion, Hg^{2+}, is a typical monatomic ion. There is also a mercury(I) ion, but it is diatomic. Its formula is Hg_2^{2+}. The mercury(I) name is logical if you realize that *each atom* is contributing a 1+ charge to the diatomic ion.

Figure 10–2 locates in a periodic table the monatomic ions you should be able to recognize on sight, or write if given the name of the ion.

Polyatomic Anions Derived from the Total Ionization of Oxyacids

When an oxyacid ionizes, the resulting anion has more than one atom; it is a *polyatomic* anion. These are *oxyanions,* so called because they contain oxygen. Names of oxyanions are related to the acid from which they come; the prefix or suffix of the acid is replaced by a prefix or suffix for the anion. The system is illustrated for chlorine in Table 10–2, page 108. Memorize these prefixes and suffixes, and you will be able to apply them to a large number of compounds, including many you may never have heard of before.

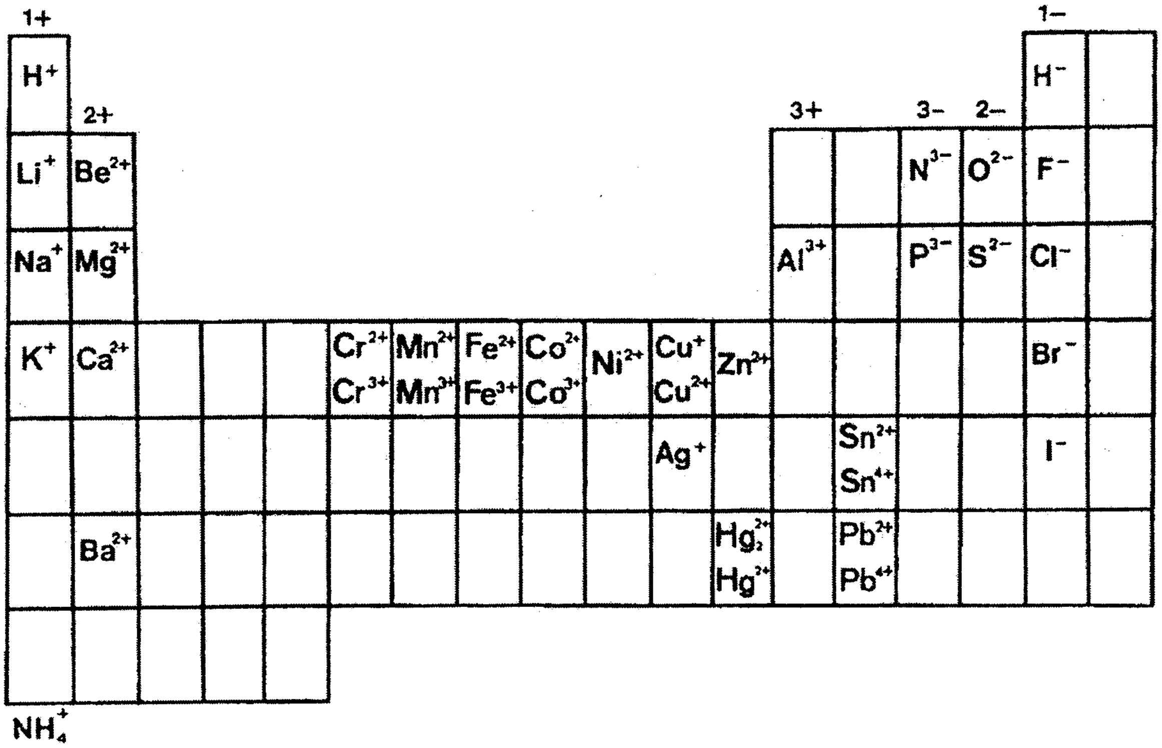

Figure 10–2. Partial periodic table of common ions. Notes: (1) Tin (Sn) and lead (Pb) form monatomic ions in a +2 oxidation state. In their +4 oxidation states they are more accurately described as being covalently bonded but such compounds are frequently named as if they were ionic compounds. (2) Hg_2^{2+} is a diatomic elemental ion. Its name is mercury(I), indicating a +1 charge from each atom in the diatomic ion. (3) Ammonium ion, NH_4^+, is included as the only other common polyatomic cation, thereby completing this table as a minimum list of the cations you should be able to recall simply by referring to a full periodic table.

The negative charge on an ion from the total ionization of an oxyacid is equal to the number of hydrogen atoms in the neutral acid molecule. Chloric acid, with one hydrogen, produces an oxyanion with a single negative charge, ClO_3^-; sulfuric acid, with two hydrogens, yields the double negative sulfate ion, SO_4^{2-}; and removal of three hydrogens from phosphoric acids yields an ion with a 3− charge, PO_4^{3-}.

Oxyanions Derived from the Stepwise Ionization of Polyprotic Acids

When an acid containing two or more hydrogen atoms ionizes, it loses the hydrogen ions one by one. There are, therefore, intermediate ions that contain hydrogen. The stepwise ionization of sulfuric acid may be represented by

$$H_2SO_4 \xrightarrow{-H^+} HSO_4^- \xrightarrow{-H^+} SO_4^{2-}$$

The HSO_4^- ion can be thought of as a sulfate ion with a hydrogen attached. It is given the logical name, hydrogen sulfate ion. When triprotic phosphoric acid, H_3PO_4, ionizes, there are two intermediate ions, $H_2PO_4^-$ and HPO_4^{2-}. The first of these is the phosphate ion with two hydrogens attached, so it is called the dihydrogen phosphate ion, which distinguishes it from HPO_4^{2-}, the hydrogen (or monohydrogen) phosphate ion. Intermediate ions from other polyprotic acids are named in a similar manner.

Other Polyatomic Ions

There are two other polyatomic ions that are so common you should recognize them instantly. These are the ammonium ion, NH_4^+, and the hydroxide ion, OH^-. Many other polyatomic ions exist, but it is not necessary that they be memorized at this time unless your instructor directs you to do so. Some of them are in Tables 10–3 and 10–4, which include most of the ions you are apt to encounter in a beginning chemistry course.

Table 10–3. Common Cations

Ionic Charge: +1		*Ionic Charge: +2*		*Ionic Charge: +3*	
Alkali Metals: Group IA		*Alkaline Earths: Group IIA*		*Group IIIA*	
Li^+	Lithium	Be^{2+}	Beryllium	Al^{3+}	Aluminum
Na^+	Sodium	Mg^{2+}	Magnesium	Ga^{3+}	Gallium
K^+	Potassium	Ca^{2+}	Calcium	*Transition Elements*	
Rb^+	Rubidium	Sr^{2+}	Strontium	Cr^{3+}	Chromium(III)
Cs^+	Cesium	Ba^{2+}	Barium	Mn^{3+}	Manganese(III)
Transition Elements		*Transition Elements*		Fe^{3+}	Iron(III)
Cu^+	Copper(I)	Cr^{2+}	Chromium(II)	Co^{3+}	Cobalt(III)
Ag^+	Silver	Mn^{2+}	Manganese(II)		
Polyatomic Ions		Fe^{2+}	Iron(II)		
NH_4^+	Ammonium	Co^{2+}	Cobalt(II)		
Others		Ni^{2+}	Nickel(II)		
H^+	Hydrogen	Cu^{2+}	Copper(II)		
or		Zn^{2+}	Zinc		
H_3O^+	Hydronium	Cd^{2+}	Cadmium		
		Hg_2^{2+}	Mercury(I)		
		Hg^{2+}	Mercury(II)		
		Others			
		Sn^{2+}	Tin(II)		
		Pb^{2+}	Lead(II)		

Table 10–4. Common Anions

Ionic Charge: −1				*Ionic Charge: −2*		*Ionic Charge: −3*	
Halogens: Group VIIA		*Oxyanions*		*Group VIA*		*Group VA*	
F^-	Fluoride	ClO_4^-	Perchlorate	O^{2-}	Oxide	N^{3-}	Nitride
Cl^-	Chloride	ClO_3^-	Chlorate	S^{2-}	Sulfide	P^{3-}	Phosphide
Br^-	Bromide	ClO_2^-	Chlorite				
I^-	Iodide	ClO^-	Hypochlorite	*Oxyanions*		*Oxyanion*	
Acidic Anions		BrO_3^-	Bromate	CO_3^{2-}	Carbonate	PO_4^{3-}	Phosphate
HCO_3^-	Hydrogen carbonate	BrO_2^-	Bromite	SO_4^{2-}	Sulfate		
HS^-	Hydrogen sulfide	BrO^-	Hypobromite	SO_3^{2-}	Sulfite		
HSO_4^-	Hydrogen sulfate	IO_4^-	Periodate	$C_2O_4^{2-}$	Oxalate		
HSO_3^-	Hydrogen sulfite	IO_3^-	Iodate	CrO_4^{2-}	Chromate		
$H_2PO_4^-$	Dihydrogen phosphate	NO_3^-	Nitrate	$Cr_2O_7^{2-}$	Dichromate		
Other Anions		NO_2^-	Nitrite	*Acidic Anion*			
SCN^-	Thiocyanate	OH^-	Hydroxide	HPO_4^{2-}	Monohydrogen phosphate		
CN^-	Cyanide	$C_2H_3O_2^-$	Acetate	*Diatomic*			
H^-	Hydride	MnO_4^-	Permanganate	O_2^{2-}	Peroxide		

Ionic Compounds

Two rules govern the nomenclature of ionic compounds:

1. The name of an ionic compound is the name of the positive ion followed by the name of the negative ion.
2. The formula of an ionic compound is the formula of the positive ion followed by the formula of the negative ion, each taken as many times as may be necessary to bring the total charge to zero.

To name an ionic compound when given the formula, you need only to recognize the ions present. You must be familiar with the number of oxygen atoms in the various oxyanions, as well as the rules by which the anions are named. For a compound having a cation from a metal that forms two different monatomic ions, you must apply the oxidation-state rules to determine which of those ions is present. If the compound is $FeCl_2$, for example, you must recognize that the chloride ion has a 1− charge. There are two chloride ions present, so that the total negative charge in the formula unit is 2−. The sum of all the oxidation numbers in the formula must be zero, which means the 2+ charge must come from the iron ion, and the compound must therefore be iron(II) chloride. Similar reasoning would lead to the conclusion that $FeCl_3$ is iron(III) chloride.

In writing the formulas of compounds in which a polyatomic ion appears more than once, the entire ion is enclosed in parentheses, followed by a subscript indicating the number of ions in the formula unit. For example, the formula of calcium nitrate is $Ca(NO_3)_2$. This is the only time parentheses are used. Specifically, they are not used when a polyatomic ion appears only once in the formula, as in calcium sulfate, $CaSO_4$. Nor is the symbol of a monatomic ion enclosed in parentheses just because it happens to have two letters, as in calcium bromide, $CaBr_2$.

Write the formulas or names of the elements or compounds in the blanks provided.

A. ELEMENTS

Iron	Na
Calcium	Cl_2
Nitrogen	Cu
Bromine	Mg
Potassium	Ni

B. MOLECULAR COMPOUNDS

Carbon dioxide CO_2	CBr_4 Carbon tetrabromide
Dinitrogen tetraoxide N_2O_4	CO Carbon monoxide
Iodine chloride ICl	P_2O_5 Diphosphorus pentaOxide
Sulfur trioxide SO_3	SiS_2 Silicon disulfide
Silicon tetrachloride $SiCl_4$	S_2F_6 Disulfur hexaFloride
Bromine hexafluoride BrF_6	B_4Se_6 TetraBoron hexaselenide
Nitrogen triiodide NI_3	$Si(CO_3)_2$ Silicon dicarbonate
Silicon dicarbide SiC_2	$N(NO_3)_3$ Nitrogen trinitrate

C. ACIDS

Hydrochloric acid Chloride HCl	HClO Hypochlorite Hypochlorous acid
Sulfuric acid sulfurate H_2SO_4	HI Iodide HydroIodic acid
Nitrous acid Nitrite HNO_2	HNO_3 Nitrate Nitric Acid
Carbonic acid Carbonate H_2CO_3	$HClO_4$ Perchlorate Perchloric Acid
Phosphoric acid phosphate H_3PO_4	$H_2SO_4^{-2}$ Sulfate sulfic acid
Hydrobromic acid HBr	HF Floride Hydrofluoric acid
Nitric acid HNO_3	HIO_3 Iodate Iodic Acid
Sulfurous acid H_2SO_3	H_2CO_3 Carbonate Carbonic Acid

D. COMPOUNDS

Sodium nitrate $Na^{+1} NO_3^{-1}$ $NaNO_3$	**K_2SO_4** $^{2-}$ [illegible]
Calcium fluoride $Ca^{+2} F^{-1}$ CaF_2	**Na_3PO_4** ^{3} [illegible]
Sodium carbonate $Na^{+1} CO_3^{-2}$ Na_2CO_3	**$Pb(NO_3)_2$** Lead (II) Nitrate
Potassium bromide $K^{+1} Br^{-1}$ KBr	**$FeCl_3$** Iron (III) chloride
Iron (III) sulfide $Fe^{+3} S^{-2}$ Fe_2S_3	**$CaSO_4$** [illegible] sulfate
Magnesium chloride $Mg^{+2} Cl^{-1}$ $MgCl_2$	**$Ca(OH)_2$** [illegible] Hydroxide
Ammonium sulfate $NH_4^{+1} SO_4^{-2}$ $(NH_4)_2SO_4$	**$Al_2(SO_4)_3$** Aluminum sulfate
Copper (II) sulfate $Cu^{+2} SO_4^{-2}$ $CuSO_4$	**$HgCO_3$** Mercury (II) Carbonate
Barium hydroxide $Ba^{+2} OH^{-1}$ $BaOH_2$	**K_2O** Potassium Oxide
Silver bromide $Ag^{+1} Br^{-1}$ AgBr	**$(NH_4)_2CO_3$** [illegible] Carbonate
Mercury (II) sulfate $Hg^{+2} SO_4^{-2}$ $HgSO_4$	**$PbSO_4$** Lead (II) sulfate
Potassium nitrate $K^{+1} NO_3^{-1}$ KNO_3	**FeO** Iron (II) oxide
Calcium chlorate $Ca^{+2} ClO_3^{-1}$ $Ca(ClO_3)_2$	**CaI_2** [illegible] Iodide
Iron (II) hydroxide $Fe^{+2} OH^{-1}$ $Fe(OH)_2$	**NH_4Br** Ammonium Bromide
Copper(II) phosphate $Cu^{+2} PO_4^{-3}$ $Cu_3(PO_4)_2$	**$BaCl_2$** Barium Chloride
Copper (II) phosphide $Cu^{+2} P^{-3}$ Cu_3P_2	**PbS** Lead (II) sulfide
Aluminum sulfite [illegible] $Al_2(SO_3)_3$	**$FePO_4$** Iron (III) Phosphate

D. COMPOUNDS CONTINUED

Magnesium oxide $Mg^{+2}O^{-2}$ MgO	**Ag_2SO_4** Silver Sulfate
Lead (II) iodide PbI_2	**$Co(OH)_2$** Cobalt (II) Hydroxide
Ammonium carbonate $(NH_4)_2CO_3$	**Cu_2O** Copper (I) Oxide
Aluminum oxide Al_2O_3	**K_3PO_4** Potassium Phosphate
Potassium perchlorate $KClO_4$	**Ba_3As_2** Barium Arsenide
Lithium sulfate Li_2SO_4	**$KClO_4$** Potassium Perchlorate
Lead (IV) phosphate $Pb_3(PO_4)_4$	**$Pb(NO_3)_4$** Lead (IV) Nitrate

E. MIXED COMPOUNDS

Aluminum nitride $Al^{+3}N^{-}$ AlN	**$W(NO_3)_6$** Tungsten nitrate
Strontium phosphate $Sr_3(PO_4)_2$	**$Ni(NO_3)_2$** Nickel (II) Nitrate
Barium iodide BaI_2	**$ZnSO_4$** Zinc Sulfate
Tin (II) sulfate $SnSO_4$	**Cr_2O_3** Chromium (III) Oxide
Phosphorus pentachloride PCl_5	**H_2SO_3** Sulfurous Acid
Iron (III) sulfide Fe_2S_3	**$(NH_4)_2S$** Ammonium Sulfide
Dichlorine hexabromide Cl_2Br_6	**$I_3(NO_3)_4$** Triiodine tetranitrate
Ammonium phosphate $(NH_4)_3PO_4$	**$MgSO_4$** Magnesium Sulfate
Mercury (II) phosphide Hg_3P_2	**Li_2O** Lithium Oxide

E. MIXED COMPOUNDS

Aluminum sulfite $Al_2(SO_3)_3$	**AlI_3** Aluminum Iodide
Magnesium nitrate $Mg(NO_3)_2$	**$C(NO_3)_4$** Carbon tetranitrate
Cesium carbide Cs_4C	**$Bi_2(SO_4)_3$** Bismuth (III) sulfate
Barium nitride Ba_3N_2	**SnO_2** Tin (IV) oxide
Zinc chloride $ZnCl_2$	**NO_2** Nitrogen dioxide
Gold (III) oxide Au_2O_3	**$Mn_2(SO_4)_3$** Maganese (III) Sulfate