EXPERIMENT 9

Chemical Names and Formulas: A Study Assignment

Performance Goal

9–1 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 9.1. 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 the sheer memory work is reduced if you relate elemental names and symbols to the periodic table.

IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII	VIII	VIII	IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
1 H																1 H	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca				24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn			33 As	34 Se	35 Br	36 Kr
	38 Sr									47 Ag			50 Sn	51 Sb		53 I	
	56 Ba									79 Au	80 Hg		82 Pb	83 Bi			

Figure 9.1
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	**bismuth**	**cobalt**	**hydrogen**	**magnesium**	**oxygen**	**sodium**
antimony	**boron**	**chromium**	**iodine**	**manganese**	**phosphorus**	**strontium**
argon	**bromine**	**copper**	**iron**	**mercury**	**potassium**	**sulfur**
arsenic	**calcium**	**fluorine**	**krypton**	**neon**	**selenium**	**tin**
barium	**carbon**	**gold**	**lead**	**nickel**	**silicon**	**zinc**
beryllium	**chlorine**	**helium**	**lithium**	**nitrogen**	**silver**	

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 nonmetals.* You may use this identification

Weiner_5209_ch09, 12/29/8, 17:31, page: 109

Table 9.1 *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

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 9.1. 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* 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, because 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 9.2. The key to the entire nomenclature system is the number of oxygen atoms *compared to the number in the -ic acid.* Study Table 9-2 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.

Table 9.2 *Names of Oxyacids and Oxyanions of Chlorine (HCl included for comparison)*

I	*II*	*III*	*IV*	*V*	*VI*	*VII*
Acid Name	*Acid Suffixes and Prefixes*	*Acid Formula*	*Oxygens Compared to -ic Acid*	*Ion Name*	*Ion Suffixes and Prefixes*	*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^-

Weiner_5209_ch09, 12/29/8, 17:31, page: 111

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.
6. In a compound, the oxidation numbers of all atoms add up to zero.

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.

1+	2+											3+		3–	2–	1–	
H^+																H^-	
Li^+	Be^{2+}													N^{3-}	O^{2-}	F^-	
Na^+	Mg^{2+}											Al^{3+}		P^{3-}	S^{2-}	Cl^-	
K^+	Ca^{2+}				Cr^{2+} Cr^{3+}	Mn^{2+} Mn^{3+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+}	Cu^+ Cu^{2+}	Zn^{2+}			As^{3-}	Se^{2-}	Br^-	
	Sr^{2+}									Ag^+			Sn^{2+} Sn^{4+}	Sb^{3+}		I^-	
	Ba^{2+}										Hg_2^{2+} Hg^{2+}		Pb^{2+} Pb^{4+}	Bi^{3+}			

NH_4^+

Figure 9.2

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.

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 used if a metal forms only one kind of ion.

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 9.2 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 9.2, page 110. 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.

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 listed in Tables 9.3 and 9.4, which include most of the ions you are apt to encounter in a beginning chemistry course.

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 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.

Table 9.3 *Common Cations*

Ionic Charge: +1		*Ionic Charge: +2*		*Ionic Charge: +3*	
Alkali Metals: Group IA		*Alkali Earths: Group IIA*		*Group IIIA*	
Li^+	Lithium	Be^{2+}	Beryllium	Al^{3+}	Aluminum
Na^+	Sodium	Mg^{2+}	Magnesium	Bi^{3+}	Bismuth
K^+	Potassium	Ca^{2+}	Calcium	Sb^{3+}	Antimony
Rb^+	Rubidium	Sr^{2+}	Strontium	*Transition Elements*	
Cs^+	Cesium	Ba^{2+}	Barium	Cr^{3+}	Chromium(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		
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)		

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, $CaBr_2$.

Table 9.4 *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	*Oxyanions*		*Oxyanion*	
I^-	Iodide	ClO^-	Hypochlorite	CO_3^{2-}	Carbonate	PO_4^{3-}	Phosphate
Acidic Anions				SO_4^{2-}	Sulfate	PO_3^{3-}	Phosphite
HCO_3^-	Hydrogen carbonate	BrO_3^-	Bromate	SO_3^{2-}	Sulfite		
		BrO_2^-	Bromite	$C_2O_4^{2-}$	Oxalate		
HS^-	Hydrogen sulfide	BrO^-	Hypobromite	CrO_4^{2-}	Chromate		
		IO_4^-	Periodate	$Cr_2O_7^{2-}$	Dichromate		
HSO_4^-	Hydrogen sulfate	IO_3^-	Iodate				
		IO^-	Hypoiodite	*Acidic Anion*			
HSO_3^-	Hydrogen sulfite	NO_3^-	Nitrate	HPO_4^{2-}	Monohydrogen phosphate		
		NO_2^-	Nitrite	*Diatomic*			
$H_2PO_4^-$	Dihydrogen phosphate						
		OH^-	Hydroxide	O_2^{2-}	Peroxide		
		$C_2H_3O_2^-$	Acetate				
Other Anions		MnO_4^-	Permanganate				
SCN^-	Thiocyanate						
CN^-	Cyanide						
H^-	Hydride						

Name | Date | Section

Experiment 9

Advance Study Assignment

1. Give the names and formulas of two oxyacids of chlorine.

2. Give the formula and name of an ionic compound containing (a) the sulfate ion, (b) the nitrite ion.

3. Write the formula of an ionic compound (using the symbols X and Y) derived from:

   a. X^{+} cation and Y^{3-} anion.

   b. X^{2+} cation and HY^{-} anion.

   c. X^{3+} cation and HY^{2-} anion.

Name ____________ Date ____________ Section ____________

Experiment 9

Work Page

General Instructions: For each substance whose name is given, write the formula; if the formula is given, write the name. Unless stated otherwise, a periodic table should be your only reference.

Elements

Write the formulas of the elements in their natural, stable states.

Iron	Na
Calcium	Cl_2
Nitrogen	Cu
Bromine	Mg
Potassium	Ni

Molecular Binary Compounds

Carbon dioxide	CBr_4
Dinitrogen tetroxide	CO
Iodine chloride	P_2O_3
Sulfur trioxide	SiS_2
Diphosphorus pentoxide	S_2F_6

Acids

Hydrobromic acid	HClO
Sulfuric acid	HI
Bromic acid	HNO_3
Phosphoric acid	H_2SO_3
Nitrous acid	HIO_4
Perchloric acid	$HBrO_2$

Weiner_5209_ch09, 12/29/8, 17:31, page: 120

Monatomic and Polyatomic Ions

Calcium ion	Fe^{2+}
Sulfate ion	Br^{-}
Monohydrogen phosphate ion	CIO^{-}
Nitrite ion	CO_3^{2-}
Iron(III) ion	Cr^{3+}
Iodite ion	SO_3^{2-}
Sulfide ion	HCO_3^{-}

Name ____________________ *Date* ____________ *Section* ____________

Experiment 9

Work Page

Ionic Compounds

Sodium nitrate	K_2SO_4
Calcium fluoride	Na_3PO_4
Potassium hydrogen sulfate	$Pb(NO_3)_2$
Sodium carbonate	$FeCl_3$
Potassium bromide	KIO_3
Iron(III) sulfide	$Ca(OH)_2$
Magnesium chloride	$Al_2(SO_4)_3$
Sodium dihydrogen phosphate	$HgCO_3$
Ammonium sulfate	$NaClO_2$
Copper(II) carbonate	KHS
Barium hydroxide	K_2O
Silver bromide	$NaHSO_3$
Mercury(II) sulfate	$(NH_4)_2CO_3$
Potassium nitrite	FeO
Calcium chlorate	$NaHCO_3$
Iron(II) hydroxide	CaI_2
Copper(I) iodate	NH_4Br
Aluminum sulfite	$BaCl_2$
Magnesium oxide	$FePO_4$
Lead(II) iodide	Ag_2SO_4
Sodium hypochlorite	$Co(OH)_2$

Ionic Compounds (Continued)

Lithium hydrogen sulfite	NH_4NO_2
Ammonium carbonate	Cu_2O
Mercury(I) chloride	K_3PO_4
Aluminum oxide	$(NH_4)_2HPO_4$
Potassium periodate	$AgBrO_3$

Name | Date | Section

Experiment 9

Work Sheet

Compounds Containing Less Common Ions

(Refer to tables of cations and anions and the periodic table when writing these formulas.)

Strontium sulfate (strontium, atomic number 38)	
Cesium iodide (cesium, atomic number 55)	
Indium chloride (indium, atomic number 49)	
Tellurium trioxide (tellurium, atomic number 52)	
Calcium hydride	
Sodium cyanide	
Iron(III) thiocyanate	
Nickel(II) chromate	

Name | Date | Section

Experiment 9

Report Sheet

General Instructions: For each substance whose name is given, write the formula; if the formula is given, write the name. Unless stated otherwise, a periodic table should be your only reference.

Elements

Write the formulas of the elements in their natural, stable states.

Iron	Na
Calcium	Cl_2
Nitrogen	Cu
Bromine	Mg
Potassium	Ni

Molecular Binary Compounds

Carbon dioxide	CBr_4
Dinitrogen tetroxide	CO
Iodine chloride	P_2O_3
Sulfur trioxide	SiS_2
Diphosphorus pentoxide	S_2F_6

Acids

Hydrobromic acid	HClO
Sulfuric acid	HI
Bromic acid	HNO_3
Phosphoric acid	H_2SO_3
Nitrous acid	HIO_4
Perchloric acid	$HBrO_2$

Monatomic and Polyatomic Ions

Calcium ion	Fe^{2+}
Sulfate ion	Br^{-}
Monohydrogen phosphate ion	ClO^{-}
Nitrite ion	CO_3^{2-}
Iron(III) ion	Cr^{3+}
Iodite ion	SO_3^{2-}
Sulfide ion	HCO_3^{-}

Name | Date | Section

Experiment 9

Report Sheet

Ionic Compounds

Sodium nitrate	K_2SO_4
Calcium fluoride	Na_3PO_4
Potassium hydrogen sulfate	$Pb(NO_3)_2$
Sodium carbonate	$FeCl_3$
Potassium bromide	KIO_3
Iron(III) sulfide	$Ca(OH)_2$
Magnesium chloride	$Al_2(SO_4)_3$
Sodium dihydrogen phosphate	$HgCO_3$
Ammonium sulfate	$NaClO_2$
Copper(II) carbonate	KHS
Barium hydroxide	K_2O
Silver bromide	$NaHSO_3$
Mercury(II) sulfate	$(NH_4)_2CO_3$
Potassium nitrite	FeO
Calcium chlorate	$NaHCO_3$
Iron(II) hydroxide	CaI_2
Copper(I) iodate	NH_4Br
Aluminum sulfite	$BaCl_2$
Magnesium oxide	$FePO_4$
Lead(II) iodide	Ag_2SO_4
Sodium hypochlorite	$Co(OH)_2$

Ionic Compounds (Continued)

Lithium hydrogen sulfite	NH_4NO_2
Ammonium carbonate	Cu_2O
Mercury(I) chloride	K_3PO_4
Aluminum oxide	$(NH_4)_2HPO_4$
Potassium periodate	$AgBrO_3$

Name | Date | Section

Experiment 9

Report Sheet

Compounds Containing Less Common Ions

(Refer to tables of cations and anions and the periodic table when writing these formulas.)

Strontium sulfate (strontium, atomic number 38)	
Cesium iodide (cesium, atomic number 55)	
Indium chloride (indium, atomic number 49)	
Tellurium trioxide (tellurium, atomic number 52)	
Calcium hydride	
Sodium cyanide	
Iron(III) thiocyanate	
Nickel(II) chromate	