Exercise 1–1: Symbols and Formulas

Chemical symbols are used for convenience to represent the names of the elements. Some are merely initials of the names. All of these are older elements. No new future elements that are discovered will be allowed to have symbols with only one letter.

H	________	O	________	C	________
F	________	I	________	B	________
U	________	P	________	N	________
Y	________	S	________	V	________

Others consist of the first letter in the name of the element and a second letter which is prominent in pronouncing the name of the element. (This is the modern day rule that is followed.)

Si	________	Ca	________	Bi	________
Ra	________	Al	________	Mn	________
Cl	________	Cd	________	Br	________
Li	________	Ni	________	Zn	________
Mg	________	As	________	Ba	________
Pt	________	Pu	________	Np	________
Co	________	Ti	________	Cr	________
Sg	________	Sr	________	Ga	________

Some symbols are derived from non-English words. i.e., Latin, Greek, or German names.

Fe	(ferrum)	________	Cu	(cuprum)	________
Na	(natrium)	________	K	(kalium)	________
Ag	(argentum)	________	Hg	(hydrargyrum)	________
Sn	(stannum)	________	Sb	(stibium)	________
Pb	(plumbum)	________	Au	(aurium)	________
W	(wolfram)	________			

Symbols are used as a sort of shorthand in writing the names of elements. The use of symbols to represent atoms, or definite quantities by mass of the elements is also important in writing chemical formulas and in describing reactions. Thus, the symbol C represents the element carbon, but it also represents one atom of carbon. ***Note:*** The first letter of the symbol is always printed uppercase; the second letter is always printed lowercase!

Fill in the blanks given above with the names of the elements the symbols represent. Be able to give either the symbol or its name from memory.

CHAPTER 2

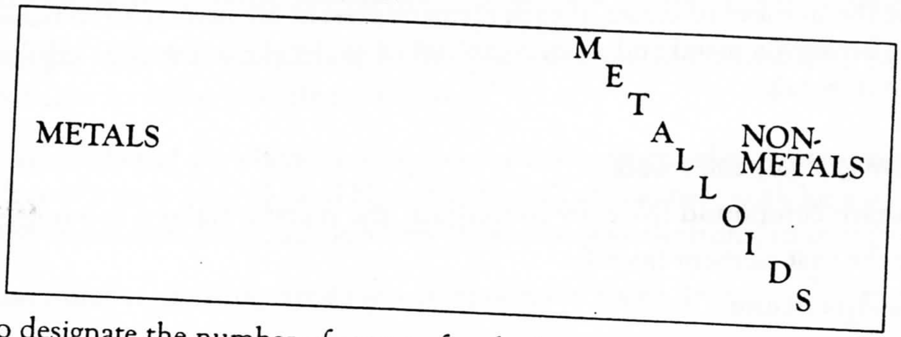
Simple Inorganic Formulas and Nomenclature

Compounds consisting of two different type elements are considered to be *binary compounds*. Binary compounds usually end in the suffix "ide." There are two types of binary compounds–binary molecules and binary salts. A binary molecule consists of two nonmetals bonded via covalent bonding. A binary salt consists of a metal and a nonmetal exhibiting ionic bonding.

General Rules:

A. Binary Molecules (Nonmetal + Nonmetal Compounds) i.e., CO_2 or N_2O_3

Molecules are formed when two nonmetals or metalloids combine and prefixes must be used to designate the number of atoms of each element present in one molecule. Nonmetals are found just to the right of the zigzag line on the periodic table. Metalloids are near the zigzag line and have some properties of metals and some properties of nonmetals.



Prefixes are used to designate the number of atoms of each element present in the formula of a binary compound. *Mono* is **never** used in front of the first element (standard convention). If there is only one atom, the mono is assumed.

1 = mono	4 = tetra	7 = hepta	10 = deca
2 = di	5 = penta	8 = octa	11 = undeca
3 = tri	6 = hexa	9 = nona	12 = dodeca

Name the following binary molecules—CO_2 and N_2O_3

First word in the compound:

1. Give the prefix designating the number of atoms of the first element present. Remember, *mono* is never used (by standard convention) for the first element.

   CO_2: No prefix for C

   N_2O_3: di

2. Name the first element.

   CO_2: carbon

   N_2O_3: dinitrogen

Second word in the compound:

3. Give the prefix designating the number of atoms of the second element present.
   CO_2: carbon di
   N_2O_3: dinitrogen tri
4. Name the root of the second element.
   (*Note:* The root is the base name which designates the element.)
   CO_2: carbon diox
   N_2O_3: dinitrogen triox
5. Add *–ide* to the root of the second element.
   CO_2: carbon dioxide (official name)
   N_2O_3: dinitrogen trioxide (official name)

B. Binary Salts (Metal + Nonmetal compounds) i.e., $CaCl_2$

Prefixes giving the number of atoms of each element present are *never* used to name a salt. Salts exhibit ionic bonding between a metal and a nonmetal, while molecular substances exhibit covalent bonding between two nonmetals.

Name the following binary salt—$CaCl_2$

First word in each compound (*Note:* By convention, the metal is written before the nonmetal.):

1. Name the first element (metal).
   $CaCl_2$: calcium

Second word in the compound:

2. Name the root of the second element (nonmetal).
   $CaCl_2$: calcium chlor
3. Add the suffix *–ide* to the second element.
   $CaCl_2$: calcium chloride

Exercise 2–1: In column 1, classify each of the following compounds as binary molecules (M) or binary ionic salts (I). Then in column 2, use the rules to name each binary compound.

1. CaF_2	____	____________	10. SrI_2	____	____________
2. P_4O_{10}	____	____________	11. CO	____	____________
3. K_2S	____	____________	12. Cs_2Po	____	____________
4. NaH	____	____________	13. $ZnAt_2$	____	____________
5. Al_2Se_3	____	____________	14. P_2S_3	____	____________
6. N_2O	____	____________	15. AgCl	____	____________
7. O_2F	____	____________	16. Na_3N	____	____________
8. SBr_6	____	____________	17. Mg_3P_2	____	____________
9. Li_2Te	____	____________	18. XeF_6	____	____________

CHAPTER 3

Oxidation Numbers: Anions and Cations

Metals with Variable Charges (Oxidation Numbers)

A number of metallic elements can form compounds in which the metallic ions (cations) may have different charges. These charges are known as oxidation numbers and are sometimes referred to as valences. The transition metals in the middle of the periodic table have variable oxidation numbers as do many of the representative elements in columns III, IV, V, and VI. Cations with variable oxidation numbers use a Roman numeral system (the Stock System) enclosed in parentheses to designate the charge on the cation (metallic ion). For example, the oxidation number of iron in the following two compounds cannot be the same: $FeCl_2$ and $FeCl_3$. Calling both of these compounds iron chloride would only lead to confusion. The Stock System is used to differentiate between ions that have two or more possible charges. $FeCl_2$ is known as iron(II) chloride and $FeCl_3$ is officially called iron(III) chloride. The Roman numeral represents the charge on the cation (metal) and does *not* represent the number of atoms of the element present. To name these types of ionic compounds, the oxidation numbers of all the elements present must be known.

Here are some simple rules that should help in the determination of the oxidation numbers of metallic ions (cations) from the formulas of their compounds.

1. The oxidation number of any **element** in its free state (uncombined with other elements) is zero. e.g., Fe in a bar of iron is zero. O_2 and N_2 in the Earth's atmosphere both have oxidation numbers of zero. When an element has equal numbers of protons and electrons, its overall charge is zero.

2. The oxidation number of **alkali metals** in a compound is always 1+, e.g., Li^+, Na^+, K^+, etc.

3. The oxidation number of **alkaline earth metals** in a compound is always 2+, e.g., Mg^{2+}, Ca^{2+}, Sr^{2+}, etc.

4. **Fluorine** is always assigned a value of 1^- in a compound, e.g., F^-.

5. The oxidation number of **oxygen** is almost always 2^- in a compound. Exceptions to this rule would be peroxides, O_2^{2-} where the oxidation number of each oxygen is 1^-, and superperoxides, O_2^- where the oxidation number of each oxygen is $\frac{1}{2}^-$. Neither peroxides nor superperoxides are common. Peroxides are only known to form compounds with the elements in the first two columns of the periodic table, e.g., H_2O_2, Na_2O_2, CaO_2, etc. Potassium, rubidium, and cesium are the only elements that form superperoxides, e.g., KO_2.

6. In covalent compounds (with nonmetals), **hydrogen** is assigned an oxidation number of 1+, e.g., HCl, H_2O, NH_3, CH_4. The exception to this rule is when hydrogen combines with a metal to form a **hydride**. Under these conditions, which are rare, hydrogen is assigned an oxidation number of 1^-, e.g., NaH.

7. In **metallic halides** the halogen (F, Cl, Br, I, At) always has an oxidation number equal to 1^-.

8. Sulfide, selenide, telluride, and polonide are always 2^- in binary salts.

9. Nitrides, phosphides, and arsenides are always 3^- in binary salts.

10. All other oxidation numbers are assigned so that the sum of the oxidation numbers of each element equals the net charge on the molecule or polyatomic ion. In neutral compounds, the sum of the positive and negative charges must equal zero.

Example:

Determine the oxidation number of the underlined element: $K\underline{Mn}O_4$. Since K is an alkali metal, its charge must be 1+. Oxygen is 2– but there are four of them, therefore, 4 times 2– equals 8–. If 1+ and 8– are added together, we get 7–. In order for the compound to be neutral, the Mn must be 7+.

Other Examples:

$\underline{N}H_4^+$: The sum of the charges on this polyatomic ion must equal 1+. Since hydrogen has a 1+ charge and there are four hydrogen atoms, the nitrogen must be 3– because (3–) + (4+) = 1+!

$K_2\underline{Cr}_2O_7$: Potassium is 2 times 1+ = 2+ and oxygen is 7 times 2– = 14–. (14–) + (2+) = 12–. Since there are two chromium atoms and the compound is neutral overall, the charge on the two chromium atoms must be equal to 12+ and each chromium atom must have a charge of 6+ (since 12+/2 = 6+).

$\underline{O}_2$: This is an element in its free state, so the oxidation number must be zero.

(*Note:* Ions written alone, such as peroxide, must be written with a charge on them, e.g., O_2^{2-}. In a compound, the charges on individual atoms or ions are not shown.)

Exercise 3–1: Determine the oxidation number of each underlined element.

1. $K_2\underline{S}$
2. $Na\underline{Cl}O_4$
3. $\underline{Br}Cl$
4. $Li_2\underline{C}O_3$
5. $\underline{O}F_2$
6. $\underline{S}_8$
7. $\underline{Mg}$
8. $K_2\underline{W}_4O_{13}$
9. $Mg(\underline{B}F_4)_2$
10. $\underline{Au}_2O_3$
11. $\underline{C}_{60}$
12. $\underline{Zr}O_2$
13. $\underline{Nb}OF_6^{3-}$
14. $Al_2(\underline{Cr}O_4)_3$
15. $Cs_2\underline{Te}F_8$

Remember, free elements, no matter how complex the molecule, have an oxidation number (valence or charge) equal to zero. The following are diatomic or polyatomic elements in nature which must be committed to memory. These elements exist as neutral molecules in nature!

Polyatomic Elements

Hydrogen, H_2	Bromine, Br_2
Nitrogen, N_2	Iodine, I_2
Oxygen, O_2	Ozone, O_3
Fluorine, F_2	Phosphorus, P_4
Chlorine, Cl_2	Sulfur, S_8

Most common forms of buckminsterfullerenes (buckyballs): C_{60} & C_{70}

Representative Elements (s- or p-Differentiating) Cations and Anions

Charges can be determined by position (family) on the Periodic Table. Cations (+ ions) come from metals that lose electrons (oxidation) in order to become isoelectronic with a noble gas. Anions (- ions) come from nonmetals that gain electrons (reduction) to become isoelectronic with a noble gas.

Oxidation Numbers (Valence) of Representative Element Cations and Anions

Cations		Anions	
1+ *Alkali metals*	lithium, Li^+; sodium, Na^+; potassium, K^+; rubidium, Rb^+; cesium, Cs^+; francium, Fr^+; hydrogen, H^+	1– *Halogens*	fluoride, F^-; chloride, Cl^-; bromide, Br^-; iodide, I^-; astatide, At^-
2+ *Alkaline Earth metals*	beryllium, Be^{2+}; magnesium Mg^{2+}; calcium, Ca^{2+}; strontium, Sr^{2+}; barium, Ba^{2+}; radium, Ra^{2+}	2– *Oxygen family*	oxide, O^{2-}; sulfide, S^{2-}; selenide, Se^{2-}; telluride, Te^{2-}; polonide, Po^{2-}
3+	aluminum, Al^{3+}; boron, B^{3+}	3– *Nitrogen family*	nitride, N^{3-}; phosphide, P^{3-}; arsenide, As^{3-}
		4–	carbide (covalent), C^{4-}

More on Metallic Elements with Variable Oxidation Numbers

Transition metals, representative metals with *p* and *d* sublevels, and the inner transition metals typically have more than one oxidation state in compounds. Electrons for these metallic elements are lost (oxidized) in the following order: *p, s, d*. Such elements are *not* isoelectronic with a noble gas when the outermost (valence) electrons are lost and if enough energy is available, will begin to lose *d* level electrons.

Example 1: A neutral vanadium atom has an electron configuration of [Ar] $4s^2$ $3d^3$. The outermost electrons are always lost first, therefore, vanadium will lose its $4s^2$ electrons and form the vanadium(II) ion, V^{2+}. With additional energy, the V^{2+} cation can lose its $3d^3$ electrons in order, forming vanadium(III), V^{3+}, vanadium(IV), V^{4+}, and vanadium(V), V^{5+} cations.

Example 2: The electron configuration for an atom of Fe is [Ar] $4s^2$ $3d^6$. The first cation that forms when the $4s^2$ electrons are lost is the iron(II) ion, Fe^{2+}. Additional energy will cause the iron(II) ion to lose one of its 3d electrons to form the iron(III) ion, Fe^{3+}. The remaining *d* electrons are all spinning in the same direction and the energy required to oxidize them is greater than normally encountered in an ordinary chemical reaction. The repulsive forces between the only two paired electrons in the 3d sublevel make the formation of the iron(III) ion relatively easy.

Example 3: The electronic configuration of a neutral lead atom is [Xe] $6s^2$ $4f^{14}$ $5d^{10}$ $6p^2$. The two common oxidation numbers of lead are lead(II) when the two $6p^2$ electrons are lost and lead(IV) when the two $6s^2$ electrons are also oxidized. Tin behaves in a similar manner when it forms tin(II) and tin(IV) cations. Bismuth with an electron configuration of [Xe] $6s^2$ $4f^{14}$ $5d^{10}$ $6p^3$, forms bismuth(III) and bismuth(V) ions.

Inner transition elements are sometimes called by such names as the lanthanides, actinides, rare earth elements, and the transuranium elements. All of these elements are quite rare, and many of them exist for only short periods of time. Reactions involving such elements are seldom encountered in a beginning chemistry course and there is little need to pursue this topic in any detail. Two inner transition elements worth mentioning are uranium (U^{3+}, U^{4+}, and U^{5+}) and cerium (Ce^{3+} and Ce^{4+}).

Both inner transition and transition elements are known for their variable oxidation numbers. The most common oxidation number for transition elements is 2+. The *d* sublevel in transition elements is responsible for the various oxidation numbers that result. Incomplete *d* sublevels are also responsible for the many colorful transition compounds that are known to exist. Complete *d* sublevels in cations of silver and zinc result in white compounds.

Summary of Cations with Variable Oxidation Numbers—Stock System

1+, 2+	copper(I), Cu^{+}; copper(II), Cu^{2+};
	mercury(I)*, Hg_2^{2+}; mercury(II), Hg^{2+}
	**Note: mercury(I) actually exists as a diatomic ion and is written as Hg_2^{2+} and not Hg^{+}.*
1+, 3+	gold(I), Au^{+}; gold(III), Au^{3+};
	indium(I), In^{+}; indium(III), In^{3+};
	thallium(I), Tl^{+}; thallium(III), Tl^{3+}
2+, 3+	chromium(II), Cr^{2+}; chromium(III), Cr^{3+};
	cobalt(II), Co^{2+}; cobalt(III), Co^{3+};
	iron(II), Fe^{2+}; iron(III), Fe^{3+};
	manganese(II), Mn^{2+}; manganese(III), Mn^{3+}
2+, 4+	lead(II), Pb^{2+}; lead(IV), Pb^{4+};
	platinum(II), Pt^{2+}; platinum(IV), Pt^{4+};
	tin(II), Sn^{2+}; tin(IV), Sn^{4+};
	zirconium(II), Zr^{2+}; zirconium(IV), Zr^{4+}
3+, 4+	cerium(III), Ce^{3+}; cerium(IV), Ce^{4+}
3+, 5+	antimony(III), Sb^{3+}; antimony(V), Sb^{5+};
	arsenic(III), As^{3+}; arsenic(V), As^{5+};
	bismuth(III), Bi^{3+}; bismuth(V), Bi^{5+};
	phosphorus(III), P^{3+}; phosphorus(V), P^{5+}
2+, 3+, 4+	iridium(II), Ir^{2+}; iridium(III), Ir^{3+}; iridium(IV), Ir^{4+};
	titanium(II), Ti^{2+}; titanium(III), Ti^{3+}; titanium(IV), Ti^{4+}
2+, 4+, 5+	tungsten(II), W^{2+}; tungsten(IV), W^{4+}; tungsten(V), W^{5+}
3+, 4+, 5+	uranium(III), U^{3+}; uranium(IV), U^{4+}; uranium(V), U^{5+}
2+, 3+, 4+, 5+	vanadium(II), V^{2+}; vanadium(III), V^{3+};
	vanadium(IV), V^{4+}; vanadium(V), V^{5+}

Note: When reading the name of an ion such as Pb^{2+}, the ion is read in English as the "lead two ion."

Special Metallic Cations

The following transition metal cations do not exhibit variable oxidation numbers and normally are written without Roman numerals:

cadmium, Cd^{2+} silver, Ag^{+} zinc, Zn^{2+}

Nickel, on the other hand, has variable oxidation numbers, and even though it almost always appears as the nickel(II) ion, Ni^{2+}, the Roman numeral must be written.

The ions of the representative elements gallium, germanium, and indium do not have variable oxidation numbers, but are written with Roman numerals:

gallium(III), Ga^{3+} germanium(IV), Ge^{4+} indium(III), In^{3+}

Polyatomic Ions

The term *polyatomic ion* is used to describe a group of atoms that behave as a single ion. The bonding within a polyatomic ion is covalent, but because there is always an excess or shortage of electrons when compared to the number of protons present, an ion results. A common polyatomic positive ion (cation) is the ammonium ion, NH_4^{+}. A common polyatomic negative ion (anion) is the sulfate ion, SO_4^{2-}. Polyatomic ions *must be memorized!* There is no simple way to learn these ions. Remember, polyatomic ions stay together as a group. The ammonium ion is always written as NH_4^{+} and *never* as $N^{3-} + 4H^{+}$ or H_4^{+} or H_4^{4+}. If two or more of the same polyatomic ion are needed within a compound in order to reach electrical neutrality, the polyatomic group is enclosed in parentheses. For example, ammonium sulfate is written as $(NH_4)_2SO_4$. The compound consists of two ammonium ions and one sulfate ion. The letters are read as "N, H, four taken twice, S, O, four."

Common Polyatomic Ions

Anions

1–

acetate, CH_3COO^-
amide, NH_2^-
azide, N_3^-
benzoate, $C_6H_5COO^-$
bromate, BrO_3^-
chlorate, ClO_3^-
chlorite, ClO_2^-
cyanate, OCN^-
cyanide, CN^-
dihydrogen phosphate, $H_2PO_4^-$
formate, $HCOO^-$
hydrogen carbonate, HCO_3^-
(bicarbonate)
hydrogen sulfate, HSO_4^-
(bisulfate)
hydrogen sulfide, HS^-
(bisulfide or hydrosulfide)
hydroxide, OH^-
(called hydroxyl when aqueous)
hypochlorite, ClO^-
iodate, IO_3^-
nitrate, NO_3^-
nitrite, NO_2^-
perchlorate, ClO_4^-
permanganate, MnO_4^-
thiocyanate, SCN^-
(thiocyanato)
triiodide, I_3^-
vanadate, VO_3^-

2–

carbide, C_2^{2-}
(saltlike)
carbonate, CO_3^{2-}
chromate, CrO_4^{2-}
dichromate, $Cr_2O_7^{2-}$
imide, NH^{2-}
manganate, MnO_4^{2-}
metasilicate, SiO_3^{2-}
monohydrogen phosphate, HPO_4^{2-}
oxalate, $C_2O_4^{2-}$
peroxide, O_2^{2-}
peroxydisulfate, $S_2O_8^{2-}$
phthalate, $C_8H_4O_4^{2-}$
polysulfide, S_x^{2-}
selenate, SeO_4^{2-}
sulfate, SO_4^{2-}
sulfite, SO_3^{2-}
tartrate, $C_4H_4O_6^{2-}$
tellurate, TeO_4^{2-}
tetraborate, $B_4O_7^{2-}$
thiosulfate, $S_2O_3^{2-}$
tungstate, WO_4^{2-}
zincate, ZnO_2^{2-}

3–

aluminate, AlO_3^{3-}
arsenate, AsO_4^{3-}
borate, BO_3^{3-}
citrate, $C_6H_5O_7^{3-}$
phosphate, PO_4^{3-}

4–

orthosilicate, SiO_4^{4-}
pyrophosphate, $P_2O_7^{4-}$

5–

tripolyphosphate, $P_3O_{10}^{5-}$

Cations

1+

ammonium, NH_4^+
hydronium, H_3O^+

Exercise 3–2: Name the following substances.

1. $FeSO_3$
2. $Cu(NO_3)_2$
3. Hg_2Cl_2
4. $AgBr$
5. $KClO_3$
6. $MgCO_3$
7. BaO_2
8. KO_2
9. SnO_2
10. $Pb(OH)_2$
11. $Ni_3(PO_4)_2$
12. $CuCH_3COO$
13. N_2O_4
14. Rb_3P
15. S_8
16. Fe_2O_3
17. $(NH_4)_2SO_3$
18. $Ca(MnO_4)_2$
19. PF_5
20. LiH
21. $Ti(HPO_4)_2$

Exercise 3–3: Write formulas for the following substances.

1. vanadium(V) oxide
2. dihydrogen monoxide
3. ammonium oxalate
4. polonium(VI) thiocyanate
5. tetraphosphorus decaoxide
6. zinc hydroxide
7. potassium cyanide
8. cesium tartrate
9. oxygen molecule
10. mercury(II) acetate
11. silver chromate
12. tin(II) carbonate
13. sodium hydrogen carbonate
14. manganese(VII) oxide
15. copper(II) dihydrogen phosphate
16. francium dichromate
17. calcium carbide
18. mercury(I) nitrate
19. cerium(IV) benzoate
20. potassium hydrogen phthalate

CHAPTER 4

Ternary Nomenclature: Acids and Salts

The halogens, with their variable oxidation numbers, allow for a great variety of compounds. The problem arises on how these compounds should be named. For example, chlorine is found with a different oxidation state in each of the following five compounds:

$HClO_4$ (Cl = 7+)
$HClO_3$ (Cl = 5+)
$HClO_2$ (Cl = 3+)
$HClO$ (Cl = 1+)
HCl (Cl = 1-)

A good way to learn ternary nomenclature is to start with a certain "home base" polyatomic ion. This is the polyatomic ion ending with the suffix *–ate* (see page 16). Remembering that salts are named by adding the name of the metallic ion (cation) to the nonmetallic polyatomic ion (anion), the following rules apply:

Number of Oxygen Atoms (Compared to Home Base)	Polyatomic Ion Name		Acid Name (H+ Combined with Polyatomic Ion)	
Plus One Oxygen Atom	ClO_4^-	*per*chlor*ate*	$HClO_4$	*per*chlor*ic* acid
Home Base	ClO_3^-	chlor*ate*	$HClO_3$	chlor*ic* acid
Minus One Oxygen Atom	ClO_2^-	chlor*ite*	$HClO_2$	chlor*ous* acid
Minus Two Oxygen Atoms	ClO^-	*hypo*chlor*ite*	$HClO$	*hypo*chlor*ous* acid
No Oxygen Atoms	Cl^-	chlor*ide*	HCl*	*hydro*chlor*ic* acid

*Water solutions of binary hydrides form acids. The root derived from the hydride is given the prefix *hydro-* and the suffix *-ic* and is followed by the word acid. The binary hydride HCl is called hydrogen chloride (hydrogen monochloride) gas, but is known as hydrochloric acid when it is aqueous.

Common Binary Acids

Formula	Name	Anion
HF(aq)	*hydro*fluor*ic* acid	F^-, fluor*ide* ion
HCl(aq)	*hydro*chlor*ic* acid	Cl^-, chlor*ide* ion
HBr(aq)	*hydro*brom*ic* acid	Br^-, brom*ide* ion
HI(aq)	*hydro*iod*ic* acid	I^-, iod*ide* ion
H_2S(aq)	*hydro*sulfur*ic* acid	S^{2-}, sulf*ide* ion

Many common acids contain only oxygen, hydrogen, and a nonmetallic ion or a polyatomic ion. Such acids are called *oxyacids.* The suffixes *-ous* and *-ic* give the oxidation state of the atom bonded to the oxygen and the hydrogen. The *-ous* suffix always indicates the lower oxidation state and *-ic* the higher.

Common Oxyacids

Formula	Name	Anion
$HClO_4$	perchloric acid	ClO_4^- perchlorate
$HClO_3$	chloric acid	ClO_3^- chlorate
$HClO_2$	chlorous acid	ClO_2^- chlorite
$HClO$	hypochlorous acid	ClO^- hypochlorite
HNO_3	nitric acid	NO_3^- nitrate
HNO_2	nitrous acid	NO_2^- nitrite
H_2SO_4	sulfuric acid	SO_4^{2-} sulfate
H_2SO_3	sulfurous acid	SO_3^{2-} sulfite
CH_3COOH or $HC_2H_3O_2$	acetic acid	CH_3COO^- acetate or $C_2H_3O_2^-$
H_2CO_3	carbonic acid	CO_3^{2-} carbonate
$H_2C_2O_4$	oxalic acid	$C_2O_4^{2-}$ oxalate
H_3PO_4	phosphoric acid	PO_4^{3-} phosphate

Exercise 4–1: Name the following compounds.

1. HIO_3
2. $NaBrO_2$
3. $Ca_3(PO_4)_2$
4. HIO_4
5. $Fe(IO_2)_3$
6. $HAt(aq)$
7. C_6H_5COOH
8. $Hg_2(IO)_2$
9. H_3PO_3
10. NH_4BrO_3

Exercise 4–2: Write formulas for the following compounds.

1. tartaric acid
2. calcium hypochlorite
3. hydrotelluric acid
4. copper(II) nitrite
5. carbonic acid
6. hypoiodous acid
7. cyanic acid
8. phthalic acid
9. tin(IV) chromate

DO YOU KNOW YOUR ACIDS?

-IC FROM -ATE | -OUS FROM -ITE | HYDRO-, -IC, -IDE

Exercise 4–3: Complete the following table.

Name of Acid	Formula of Acid	Name of Anion
hydrochloric acid	HCl	chloride
sulfuric acid	H_2SO_4	sulfate
	HI	
		sulfite
chlorous acid		
		nitrate
	$HC_2H_3O_2$ or CH_3COOH	
hydrobromic acid		
		sulfide
	HNO_2	
chromic acid		
		phosphate