

Chemical Nomenclature and Formula Writing

Chemical nomenclature is the system that chemists use to identify and name compounds. Compounds can have two types of names: *systematic names* (names that identify the chemical composition of a chemical compound) and *common names* (traditional names based on historical discovery or reactivity behavior). For example, N_2O has both a systematic name (dinitrogen monoxide) and a common name (laughing gas).

If every substance were assigned a common name, chemists would be expected to memorize over nine million names! This is why chemists generally prefer systematic names for identifying compounds. The International Union of Pure and Applied Chemistry (IUPAC, see <http://www.iupac.com>) was founded in 1921 to provide a system of chemical nomenclature for scientists. IUPAC nomenclature rules can provide valuable structural and reactivity information. On the other hand, most people would be hard pressed to call dihydrogen monoxide by any other name but water, so both types of nomenclature have their place.

Nomenclature leads naturally to formula writing. Compounds exist in distinct combinations of elements, and knowing the proper combinations of elements is essential in chemistry. We expect sodium chloride to be NaCl and not Na_2Cl or NaCl_2 ; knowing which combination or combinations exist in nature is crucial.

The following sections will guide you through the rules of *inorganic nomenclature* and formula writing. Later you may experience the nomenclature for organic chemistry or transition metal chemistry, but most of the compounds observed in first year chemistry will fall in this category.

Part A: Nomenclature of Elemental Ions

The first step in learning nomenclature is to learn the names of the elemental ions you might see in compounds. We make a distinction between the following:

- **fixed charge cations** (metal positive ions from groups IA, IIA and Al, Ga, In, Zn, Cd, Ag (“the stairs”))
- **variable charge cations** (positive ions which do not have a fixed charge; this includes *all* transition metals not in “the stairs”, lanthanides, actinides, Tl, Pb, Sn, and Bi)
- **anions** (negative ions which are generally nonmetals)

Why two types of cations? Many metals have, for all practical purposes, only one ionic charge observed in nature. Lithium is only observed as Li^+ naturally, and even though gas phase studies of lithium ions have produced Li^{2+} and even Li^- ions, they are not observed in most settings. Many metals (such as iron) have many different *oxidation states* or *ionic charges* associated with them. The ions Fe^{2+} , Fe^{3+} and Fe^{6+} can be observed and manipulated quite readily (even at Mt. Hood Community College!); therefore, we need a method to distinguish between the various ions (namely iron(II), iron(III) and iron(VI), respectively).

Fixed Charge Cations use their elemental name.

<i>Example:</i>	Na^+ is the sodium ion	Cs^+ is the cesium ion
	Mg^{2+} is the magnesium ion	Sr^{2+} is the strontium ion
	Al^{3+} is the aluminum ion	In^{3+} is the indium ion

Variable Charge Cations use their elemental name followed by their ionic charge in parentheses. Use *Roman numerals* to distinguish the charge of the ion.

Example: Fe²⁺ is the iron(II) ion Pb²⁺ is the lead(II) ion
Fe³⁺ is the iron(III) ion Pb⁴⁺ is the lead(IV) ion
Mn⁷⁺ is the manganese(VII) ion Co⁹⁺ is the cobalt(IX) ion
U⁴⁺ is the uranium(IV) ion Ti²⁺ is the titanium(II) ion

Anions use their elemental name with the ending changed to *-ide*.

Example: Cl⁻ is the chloride ion I⁻ is the iodide ion
O²⁻ is the oxide ion Te²⁻ is the telluride ion
N³⁻ is the nitride ion As³⁻ is the arsenide ion

Part B: Nomenclature of Polyatomic Ions

Certain combinations of atoms result in stable configurations that are not easily destroyed; these are called *polyatomic ions*. Polyatomic ions can be either positive or negative, but most of them are anions (i.e. they have a negative charge.) Recognizing polyatomic ions in formulas is one of the most difficult concepts to master when learning nomenclature, and it is *very important that you memorize the following list of polyatomic ions*.

A list of polyatomic ions is given below:

nitrate	NO ₃ ⁻	hydroxide	OH ⁻	hypochlorite	ClO ⁻
nitrite	NO ₂ ⁻	cyanide	CN ⁻	chlorite	ClO ₂ ⁻
sulfate	SO ₄ ²⁻	thiocyanide	SCN ⁻	chlorate	ClO ₃ ⁻
sulfite	SO ₃ ²⁻	cyanate	OCN ⁻	perchlorate	ClO ₄ ⁻
phosphate	PO ₄ ³⁻	thiosulfate	S ₂ O ₃ ²⁻	hypobromite	BrO ⁻
phosphite	PO ₃ ³⁻	chromate	CrO ₄ ²⁻	bromite	BrO ₂ ⁻
hydrogen phosphate	HPO ₄ ²⁻	dichromate	Cr ₂ O ₇ ²⁻	bromate	BrO ₃ ⁻
dihydrogen phosphate	H ₂ PO ₄ ⁻	permanganate	MnO ₄ ⁻	perbromate	BrO ₄ ⁻
carbonate	CO ₃ ²⁻	acetate	C ₂ H ₃ O ₂ ⁻	hypoiodite	IO ⁻
hydrogen carbonate	HCO ₃ ⁻	ammonium	NH ₄ ⁺	iodite	IO ₂ ⁻
hydrogen sulfide	HS ⁻	hydrogen	H ⁺	iodate	IO ₃ ⁻
oxalate	C ₂ O ₄ ²⁻	hydride	H ⁻	periodate	IO ₄ ⁻

Part C: Nomenclature of Ionic Compounds from Ions

Knowing the nomenclature rules for ions, we can begin the naming of ionic compounds. Ionic compounds involve a *cation* (either *fixed* or *variable charge*) combining with an *anion*. Naming ionic compounds is straightforward; simply combine the ionic names with the cation first followed by the anion.

Example: **sodium** ion + **chloride** ion give **sodium chloride**
iron(III) ion + **bromide** ion gives **iron(III) bromide**
ammonium polyatomic ion + **oxide** ion gives **ammonium oxide**
aluminum ion + **sulfate** polyatomic ion gives **aluminum sulfate**

Part D: Writing Formulas for Ionic Compounds Using Nomenclature

Another important concept to master is the ability to write a chemical formula using the compound's systematic name. This can be accomplished using the following protocol:

1. Identify the elemental ions and/or polyatomic ions in the compound using the systematic name.
2. Determine the magnitude of the ionic charge on each ion
3. Assume the compound is electrically neutral *unless* the term "ion" appears in the name
4. The sum of the cation charges plus the anion charges must equal zero; combine the ions until this condition is met
5. Write the resulting formula. If more than one polyatomic ion is present, write the polyatomic portion in parentheses with a subscript after it denoting the number of polyatomic ions present.

Example: Write the formula for **sodium chloride**.

1. Sodium chloride has Na^+ and Cl^- ions
2. Sodium has a +1 charge, chloride has a -1 charge
3. Assume sodium chloride is neutral (no "ion" is present in the name)
4. Charge on sodium + charge on chloride = $(+1) + (-1) = 0$; therefore, **one** sodium ion and **one** chloride ion was required for a neutral compound.
5. 1 Na^+ ion and 1 Cl^- ion gives the formula **NaCl**

Example: Write the formula for **aluminum sulfide**.

1. Aluminum sulfide has Al^{3+} and S^{2-} ions
2. Aluminum has a +3 charge, sulfide has a -2 charge
3. Assume aluminum sulfide is neutral (no "ion" is present in the name)
4. Charge on aluminum + charge on sulfide = $(+3) + (-2) = +1$; this would indicate that combining one aluminum ion with one sulfide ion would give an *ion*. We want a *neutral* compound (see step 3); this can be accomplished by multiplying aluminum by **2** and sulfide by **3**, which results in: $2(+3) + 3(-2) = 0$. Therefore, a neutral compound would result by combining **two** aluminum ions with **three** sulfide ions.
5. 2 Al^{3+} ions and 3 S^{2-} ions give the formula **Al₂S₃**.

Example: Write the formula for **magnesium nitrate**.

1. Magnesium nitrate has Mg^{2+} and NO_3^- ions
2. Magnesium has a +2 charge, nitrate has a -1 charge
3. Assume magnesium nitrate is neutral (no "ion" is present in the name)
4. Charge on magnesium + charge on nitrate = $(+2) + (-1) = +1$; this would indicate that combining one magnesium ion with one nitrate ion would give an *ion*. We want a *neutral* compound (see step 3); this can be accomplished by multiplying magnesium by **1** and nitrate by **2**, which results in: $1(+2) + 2(-1) = 0$. Therefore, a neutral compound would result by combining **one** magnesium ion with **two** nitrate ions.
5. 1 Mg^{2+} ions and 2 NO_3^- ions give the formula **Mg(NO₃)₂**. (Note there are *two* nitrate ions, so they are placed in parentheses with a subscript two after it.)

Example: Write the formula for **titanium(IV) oxalate**.

1. Titanium(IV) oxalate has Ti^{4+} and $\text{C}_2\text{O}_4^{2-}$ ions
2. Titanium(IV) has a +4 charge, oxalate has a -2 charge
3. Assume titanium(IV) oxalate is neutral (no "ion" is present in the name)
4. Charge on titanium(IV) + charge on oxalate = $(+4) + (-2) = +2$; this would indicate that combining one titanium(IV) ion with one oxalate ion would give an *ion*. We want a *neutral* compound (see step 3); this can be accomplished by multiplying titanium(IV) by **1** and oxalate by **2**, which results in: $1(+4) + 2(-2) = 0$. Therefore, a neutral compound would result by combining **one** titanium(IV) ion with **two** oxalate ions.
5. 1 Ti^{4+} ions and 2 $\text{C}_2\text{O}_4^{2-}$ ions give the formula **$\text{Ti}(\text{C}_2\text{O}_4)_2$** . (Note there are *two* oxalate ions, so they are placed in parentheses with a subscript two after it.)

Part E: Finding Systematic Names for Ionic Compounds Using Formulas

Determining the systematic name of a compound from its formula is straightforward using these steps:

1. Identify the cation and anion in the formula. *Watch* for polyatomic ions.
2. Assume the compound is electrically neutral *unless* a charge appears in the formula
3. Determine the name of the anion and the charge on the anion
4. If a fixed charge cation is present, determine its name.
5. If a variable charge cation is present, determine its name and use this formula to find the charge on the metal: **$\text{charge}_{\text{metal}} = - (\# \text{ anions})(\text{charge}_{\text{anion}}) / (\# \text{ metal cations})$** (where # = "number of")
6. Combine the cation and anion names as per Part C. The cation goes first, followed by the anion; do not forget the Roman numeral charge in parentheses for variable charge cations.

Example: Determine the name for **NaCl**.

1. The cation is Na and the anion is Cl
2. NaCl is neutral (no charges are present in the formula)
3. The anion, the chloride ion, has a -1 charge
4. Na is a fixed charge cation, and its name is the sodium ion
5. There are no variable charge cations in NaCl
6. The name of this compound is **sodium chloride**.

Example: Determine the name for **$\text{Sr}(\text{NO}_3)_2$** .

1. The cation is Sr and the anion is NO_3 .
2. $\text{Sr}(\text{NO}_3)_2$ is neutral (no charges are present in the formula)
3. The anion, the nitrate polyatomic ion, has a -1 charge
4. Sr is a fixed charge cation, and its name is the strontium ion
5. There are no variable charge cations in $\text{Sr}(\text{NO}_3)_2$
6. The name of this compound is **strontium nitrate**.

Example: Determine the name for **Fe(NO₃)₃**.

1. The cation is Fe and the anion is NO₃.
2. Fe(NO₃)₃ is neutral (no charges are present in the formula)
3. The anion, the nitrate polyatomic ion, has a -1 charge
4. There are no fixed charge cations in Fe(NO₃)₃
5. Iron is a variable charge cation; therefore, we must use the formula to calculate the charge on the iron atom. $\text{charge}_{\text{Fe}} = -(\# \text{ nitrates})(\text{charge}_{\text{nitrate}^-}) / (\# \text{ Fe atoms}) = - (3)(-1) / (1) = +3$; therefore, this is the **iron(III) ion**.
6. The name of this compound is **iron(III) nitrate**.

Example: Determine the name for **Ru₃(PO₄)₂**.

1. The cation is Ru and the anion is PO₄.
2. Ru₃(PO₄)₂ is neutral (no charges are present in the formula)
3. The anion, the phosphate polyatomic ion, has a -3 charge
4. There are no fixed charge cations in Ru₃(PO₄)₂
5. Ruthenium is a variable charge cation; therefore, we must use the formula to calculate the charge on the ruthenium atom. $\text{charge}_{\text{Ru}} = -(\# \text{ phosphates})(\text{charge}_{\text{phosphate}^-}) / (\# \text{ Ru atoms}) = - (2)(-3) / (3) = +2$; therefore, this is the **ruthenium(II) ion**.
6. The name of this compound is **ruthenium(II) phosphate**.

Part F: Nomenclature for Binary Nonmetal Covalent Molecules

Not all compounds are ionic; indeed, many compounds *share* their electrons over the respective atoms. This class of compound is called *covalent*, and they are formed when two nonmetal elements combine.

The simplest covalent compounds are the elements that exist naturally in pairs; we refer to them as *diatomics*. These are crucial to a successful chemistry experience, and memorization is straightforward using the following acronym:

Name	Compound	Acronym
Hydrogen	H ₂	<i>Have</i>
Nitrogen	N ₂	<i>No</i>
Fluorine	F ₂	<i>Fear</i>
Oxygen	O ₂	<i>Of</i>
Iodine	I ₂	<i>Ice</i>
Chlorine	Cl ₂	<i>Clear</i>
Bromine	Br ₂	<i>Brew</i>

In addition to the diatomics, several other nonmetals exist naturally in elemental form as combinations of more than one atom. **Phosphorus** exists naturally as P₄, and **sulfur** exists as S₈.

Most nonmetal covalent compounds have more than one type of element. Since there is no ionic charge present in these molecules, we cannot use the system developed above for ionic compounds, and a new method must be used. We will use the **Greek prefixes** for our compounds; they are:

1	<i>mono</i>	6	<i>hexa</i>
2	<i>di</i>	7	<i>hepta</i>
3	<i>tri</i>	8	<i>octa</i>
4	<i>tetra</i>	9	<i>nona</i>
5	<i>penta</i>	10	<i>deca</i>

The Greek prefixes refer to the number of atoms present in the molecule. For example, "dinitrogen" implies two nitrogen atoms since the prefix *di* stands for two.

When writing systematic names for binary nonmetal covalent compounds, use the *least electronegative atom first*. The topic of electronegativity will be discussed in Chem 222, but for now, the element listed first (either in the formula or the name) will be the least electronegative element.

Just as with cations in ionic compounds, use the normal element name for the least electronegative element. If more than one exist, use the Greek symbols to represent how many. The *most* electronegative element receives an *-ide* ending (as with anions in ionic compounds) as well as a Greek prefix, *even for single elements*. This is an important distinction between the most and least electronegative elements in binary compounds: the least electronegative element uses Greek symbols only if two or more atoms are present, while the more electronegative element gets an *-ide* ending *and* a Greek prefix *regardless* of the number of atoms present.

Examples:

NO	nitrogen monoxide
N₂O	dinitrogen monoxide
NO₂	nitrogen dioxide
P₂O₃	diphosphorus trioxide
P₂O₅	diphosphorus pentoxide

In addition, there are several common names of binary covalent compounds that you should be familiar with including the following:

Common Name	Formula	Systematic Name
water	H ₂ O	dihydrogen monoxide
ammonia	NH ₃	nitrogen trihydride
laughing gas	N ₂ O	dinitrogen monoxide
nitric oxide	NO	nitrogen monoxide
phosphine	PH ₃	phosphorus trihydride
hydrazine	N ₂ H ₄	dinitrogen tetrahydride
hydrogen sulfide	H ₂ S	dihydrogen monosulfide

Part G: Nomenclature for Acids and Bases

Acid and base theory shall be discussed in detail during CH 223, but recognizing common acids and bases is important for all chemists. Acids and bases require water to become active; hence, Part G assumes all of the compounds mentioned have been dissolved in water.

Acids contain **H⁺**, the **hydrogen ion**. Acids are created when hydrogen ions combine with halogens. If no oxygen atoms are present, add the *hydro-* prefix and an *-ic acid* suffix to find the acid name.

<i>Examples:</i>	HCl	hydrochloric acid
	HBr	hydrobromic acid
	HI	hydroiodic acid

If oxygen atoms are present in the halogen acid, use the following table:

Prefix and/or Suffix	Name	Formula
<i>hydro-, -ic</i>	hydrochloric acid	HCl
<i>hypo-, -ous</i>	hypochlorous acid	HClO
<i>-ous</i>	chlorous acid	HClO ₂
<i>-ic</i>	chloric acid	HClO ₃
<i>per-, -ic</i>	perchloric acid	HClO ₄

Similar rules apply to bromide or iodide, but not fluoride.

Other common names for acids include:

HNO₃	nitric acid	H₃PO₄	phosphoric acid
HNO₂	nitrous acid	H₃PO₃	phosphorous acid
H₂SO₄	sulfuric acid	HC₂H₃O₂	acetic acid
H₂SO₃	sulfurous acid	HCN	hydrocyanic acid
H₂CO₃	carbonic acid	HF	hydrofluoric acid

More assistance with naming acids can be found in the handout, “**Guide to Common Polyatomic Ions and the Corresponding Acids**” available in the CH 221 Companion or on the CH 221 website.

One final note about acids: *technically*, an acid is only an acid if dissolved in water (i.e. if *aqueous*, with an *aq* state. If not in water, the acidic properties are lost (at least for CH 221!), and the compound should probably be written as either a binary nonmetal covalent molecule (Section F) or, if the acid contains a polyatomic ion, as a fixed charge metal with a nonmetal. Consider the following examples:

HCl(aq)	hydrochloric acid	<i>This is truly an acid since HCl is dissolved in water</i>
HCl(g)	hydrogen monochloride	<i>This is not an acid - no water! - so name this compound as a covalent compound</i>
HNO₂(aq)	nitrous acid	<i>This is a true acid, dissolved in water</i>
HNO₂(g)	hydrogen nitrite	<i>This is not an acid - no water! - so name this compound as a fixed charge metal + nonmetal due to the polyatomic ion (nitrite) present</i>

If a designation of state (i.e. aqueous, gas, solid, etc.) is not provided, then the naming system used is up to the observer (i.e. take your pick! 😊)

Bases contain **OH⁻**, the **hydroxide ion**. Bases consist of a metal cation with the hydroxide anion; hence, their nomenclature will be similar to that of Parts C, D and E, above.

<i>Examples:</i>	NaOH	sodium hydroxide
	Fe(OH)₃	iron(III) hydroxide
	NH₄OH	ammonium hydroxide

Part H: Final Words

Understanding chemical nomenclature rules and being able to write formulas for compounds can be thought of as learning to read and write a language. At first, the symbols and rules do not make much sense, but as time progresses, you master the language and a moment of euphoric inspiration occurs when "it all falls into place." Regrettably, inspiration only occurs after time has been spent practicing the material. The more you practice, the faster you will master the material.

Remember that there are five general classes of compounds:

<u>Compound Class</u>	<u>Example</u>
Fixed charge cation + anion	Al_2O_3 - aluminum oxide
Variable charge cation + anion	Fe_2O_3 - iron(III) oxide
Nonmetal binary covalent compound	P_2O_3 - diphosphorus trioxide
Acid	HIO_3 - iodic acid
Base	$\text{Al}(\text{OH})_3$ - aluminum hydroxide

Each has specific rules to learn and master. Determining the charge of variable charge cations can be difficult at first, but application of the formulas in Part D and Part E should alleviate the distress.

...oh, wait, one more thing: **Waters of Hydration** or **Hydrated Compounds** show up occasionally with a "dot water" after the name of another chemical. If you see one, add the appropriate Greek prefix plus "hydrate." Examples of hydrated compounds:

$\text{MgSO}_4 \cdot 6 \text{H}_2\text{O}$ would be magnesium sulfate hexahydrate

$\text{Cu}(\text{NO}_3)_2 \cdot 2 \text{H}_2\text{O}$ would be copper(II) nitrate dihydrate

$\text{Mn}(\text{BrO}_3)_3 \cdot 4 \text{H}_2\text{O}$ would be manganese(III) bromate tetrahydrate

Nomenclature & Formula Writing Worksheet

Name: _____

Complete the worksheets below.

This lab does not have to be typed to receive credit (as long as your handwriting is legible!)

Section One: Ion Names

Complete the chart using the appropriate elemental ion or polyatomic ion name or symbol. The first row has been filled in as an example. A list of polyatomic ions (page I-18) might prove helpful.

Ion	Name	Ion	Name
Na ⁺	sodium ion	F ⁻¹	fluoride ion
Li ⁺	_____	_____	hydride ion
_____	gold(III) ion	_____	hydroxide ion
Mo ³⁺	_____	_____	cyanide ion
W ²⁺	_____	SCN ⁻¹	_____
_____	gold(I) ion	BrO ⁻¹	_____
Mn ²⁺	_____	_____	bromite ion
_____	platinum(IV) ion	_____	acetate ion
_____	zirconium(II) ion	CrO ₄ ²⁻	_____
Mt ³⁺	_____	_____	dichromate ion
Mg ²⁺	_____	_____	phosphide ion
_____	vanadium(II) ion	_____	phosphate ion
Cr ³⁺	_____	_____	phosphite ion
Cr ²⁺	_____	S ₂ O ₃ ²⁻	_____
_____	tantalum(V) ion	IO ₄ ⁻¹	_____
Ni ²⁺	_____	_____	iodate ion
_____	silver ion	_____	hypoiodite ion
_____	ammonium ion	MnO ₄ ⁻¹	_____

Section Two: Ions from Formulas

Write the ions that you would expect from the following compounds

Example: NaCl would give: Na^+ , Cl^-

Example: $\text{Fe}_3(\text{PO}_4)_2$ would give: Fe^{2+} , PO_4^{3-}

LiBr would give:

MgCl_2 would give:

Na_2O would give:

VCl_2 would give:

$\text{Fe}(\text{NO}_3)_3$ would give:

$\text{U}(\text{ClO}_3)_4$ would give:

Section Three: Nomenclature from Ion Names

Complete the chart using the appropriate compound name using the ions given. The first row has been filled in as an example.

Cation	Anion	Compound Name
potassium	iodide	potassium iodide
magnesium	oxide	_____
rhodium(III)	chloride	_____
lead(IV)	chlorate	_____
gold(I)	cyanide	_____
cobalt(II)	nitrate	_____
barium	hydroxide	_____
ammonium	phosphate	_____

Section Four: Writing Formulas Using Nomenclature

Complete the chart by providing the correct ion symbols (with the charge) and the correct formula for each compound. The first row has been filled in as an example.

Compound	Cation	Anion	Formula
calcium nitrate	Ca^{2+}	NO_3^{-1}	$\text{Ca}(\text{NO}_3)_2$
gallium bromide	_____	_____	_____
silver nitrate	_____	_____	_____
bismuth(III) chloride	_____	_____	_____
sodium acetate	_____	_____	_____
titanium(II) hypochlorite	_____	_____	_____
lithium permanganate	_____	_____	_____
iron(III) oxalate	_____	_____	_____
cesium chloride	_____	_____	_____

Section Five: Chemical Nomenclature Using Formulas

Complete the chart by providing the correct ion symbols (with the charge) and the correct name for each formula. The first row has been filled in as an example.

Formula	Cation	Anion	Name
$\text{Ca}(\text{IO}_3)_2$	Ca^{2+}	IO_3^-	calcium iodate
ZnS	_____	_____	_____
$\text{Sr}_3(\text{PO}_3)_2$	_____	_____	_____
$\text{Ga}_2(\text{SO}_4)_3$	_____	_____	_____
$\text{V}(\text{SCN})_5$	_____	_____	_____
NaMnO_4	_____	_____	_____
$(\text{NH}_4)_2\text{S}$	_____	_____	_____
NH_4NO_2	_____	_____	_____
CrCl_6	_____	_____	_____

Section Six: Nonmetal Binary Covalent Compounds

Complete the chart by providing either the correct formula or name. The first row has been filled in as an example.

Name	Formula	Name	Formula
nitrogen dioxide	NO ₂	phosphorus trichloride	PCl ₃
_____	SCl ₄	sulfur hexachloride	_____
hydrogen monochloride	_____	_____	H ₂ S(g)
_____	PI ₃	disulfur dichloride	_____
dinitrogen tetraoxide	_____	_____	N ₂ O ₃
antimony trichloride	_____	_____	SbCl ₅
_____	SiO	carbon monoxide	_____
_____	SiO ₃	carbon dioxide	_____
phosphorus trihydride	_____	_____	NO

Section Seven: Acids and Bases

Complete the chart by providing either the correct formula or name. The first entry has been filled in as an example.

Name	Formula	Name	Formula
hydrobromic acid	HBr	phosphoric acid	_____
_____	HBrO	phosphorous acid	_____
bromous acid	_____	_____	HCN
_____	HBrO ₃	acetic acid	_____
perbromic acid	_____	_____	NaOH
sulfuric acid	_____	_____	TiOH
_____	H ₂ SO ₃	potassium hydroxide	_____
_____	HNO ₃	iron(III) hydroxide	_____
nitrous acid	_____	_____	Mg(OH) ₂

Section Eight: Combined Problems: Complete the chart by providing either the correct formula or name.

<u>Name</u>	<u>Formula</u>
_____	HCl(aq)
_____	HCl(g)
potassium chloride	_____
_____	N ₂ O ₄
nitrogen disulfide	_____
_____	LiClO ₃
aluminum dichromate	_____
_____	FeSO ₄
carbonic acid	_____
_____	SO ₃
_____	(NH ₄) ₂ CO ₃
potassium dihydrogen phosphate	_____
potassium hydrogen phosphate	_____
_____	P ₄ O ₁₀
_____	TbBr ₆
_____	ThBr ₃
_____	TlBr
_____	TiBr ₄
_____	TeBr ₂
tetrasulfur decaoxide	_____
sodium hydrogen carbonate	_____
_____	In(C ₂ H ₃ O ₂) ₃
_____	Mg(ClO ₄) ₂ · 6 H ₂ O

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