

# Naming Inorganic Compounds

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by

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## Abstract

This document provides background and reference information for identifying and naming chemical substances. The focus is on inorganic species that occur in aqueous equilibria and on substances that serve as reagents in analytical applications. For comprehensive guidance and advanced examples, see the IUPAC Red Book:

Connelly, N. G. et al. *Nomenclature of Inorganic Chemistry*,  
(IUPAC Recommendations 2005);

[http://www.iupac.org/fileadmin/user\\_upload/databases/Red\\_Book\\_2005.pdf](http://www.iupac.org/fileadmin/user_upload/databases/Red_Book_2005.pdf).

The most current version of this document is available online at:

<http://www.achem.org>.

This website and associated files are provided by the author to assist users of the text:

B. M. Tissue, *Basics of Analytical Chemistry and Chemical Equilibria*,  
(John Wiley: New York, 2013).

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## I. Introduction

The first step when trying to predict aqueous equilibria and reactions is to identify the chemical species that are present in solution. Common mistakes are failing to identify soluble ionic species or breaking bonds of polyatomic species that remain as intact molecular entities. When chemical formulas are misinterpreted, it is impossible to consider the resulting chemical equilibria. You are trying to write sentences with the wrong alphabet.

As an example, zinc chloride,  $\text{ZnCl}_2$ , is a soluble salt.<sup>1</sup> The key term here is salt, meaning that this compound exists as an ionic solid. An ionic solid consists of positively charged cations and negatively charged anions in relative amounts that make the solid neutral overall. Since this salt is soluble, we predict that dissolving it in water causes the ions to separate and form  $\text{Zn}^{2+}$  and  $\text{Cl}^-$  ions surrounded by water molecules. The “ $\text{Cl}_2$ ” in the  $\text{ZnCl}_2$  formula represents two  $\text{Cl}^-$  ions and not the dichlorine molecule. Chlorine gas,  $\text{Cl}_2$ , is not present unless an additional redox reagent is added to convert  $\text{Cl}^-$  to  $\text{Cl}_2$ . Only after identifying that  $\text{Zn}^{2+}$  and  $\text{Cl}^-$  ions are what we expect to exist in the water can we then think about what equilibria these species might undergo in solution.

Use the information in this document to familiarize yourself with common chemical reagents. You want to be able to recognize the atomic and polyatomic ions that exist in compounds and as distinct entities when dissolved in water. The principles of bonding and how chemical species behave in solution is described more fully in the document on “Solvents and Solutions.” These help documents do not attempt to discuss all of the types of reactions that these water-soluble species undergo, those topics are left for the text book.

## II. Pure Elements

Pure elements are named using the name on the periodic table. When writing chemical reactions, we often indicate the form of the species after the chemical symbol. If not specified otherwise, the form is the standard state at standard temperature and pressure (STP) of 25°C and 1 atm. The common cases are solid, (s), liquid, (l), gaseous, (g), and for a solute in aqueous solution, (aq). For example, tin refers to the pure solid metal,  $\text{Sn(s)}$ , argon is the monatomic gas,  $\text{Ar(g)}$ , and oxygen is the diatomic gas,  $\text{O}_2\text{(g)}$ . Molten tin is shown in a reaction as  $\text{Sn(l)}$ . Oxygen atoms, O, can exist in plasmas and vacuum systems, and in those cases we specify that species as atomic oxygen.

In some cases, an element symbol will have other additional symbols to indicate a specific state of matter. Some examples are a dot to specify a radical species, e.g.,  $\cdot\text{OH}$  or  $\text{HO}\cdot$  (the hydroxyl radical), or an indication that a species is in an electronic excited state, e.g.,  $\text{Ca}^*$  (an excited neutral calcium atom) or  $^1\text{O}_2$  (singlet molecular oxygen).

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<sup>1</sup> When we use the term soluble salt, we almost always mean a water-soluble salt.

## Oxidation State

The oxidation state of an atom indicates the extent to which it has lost or gained electrons. The oxidation state of the pure, neutral element is zero whether it is in a monatomic, molecular, or condensed-phase form. Another way to say this is that an oxidation state of zero indicates that an atom has equal numbers of electrons and protons. We usually do not show a zero oxidation state on the element symbol, but it can be shown explicitly with a superscript to avoid confusion, e.g.,  $\text{Sn}^0$ . For monatomic species that have gained or lost electrons relative to their neutral state, the oxidation state is the same as the ionic charge.

Oxidation states, also called oxidation numbers, are assigned only to atoms and not to molecular entities. To assign the oxidation states of atoms in polyatomic ions, oxygen atoms are almost always  $-2$  (peroxide is an exception) and hydrogen atoms are almost always  $+1$  (hydrides are exceptions). The oxidation states of other atoms in a molecule or polyatomic ion are then assigned so that the sum of all oxidation states equals the overall charge of the molecular entity. As one example, the hydrogen sulfate ion,  $\text{HSO}_4^-(\text{aq})$ , has a charge of  $-1$ . We use the rules for H and O to find the oxidation state of the sulfur atom:

$$\begin{aligned}\Sigma \text{ oxidation states} &= \text{overall charge} \\ +1 + 4(-2) + \text{S oxidation state} &= -1 \\ \text{S oxidation state} &= -1 - 1 + 8 \\ \text{S oxidation state} &= +6\end{aligned}$$

## isotopes

Specific isotopes of an element are specified by the mass number. The three natural isotopes of carbon are carbon-12, carbon-13, and carbon-14. The symbol is given the mass number as a leading superscript, e.g.,  $^{12}\text{C}$ ,  $^{13}\text{C}$ , and  $^{14}\text{C}$ , respectively, for carbon. The atomic weight of an element is the weighted average for the natural abundance of all isotopes of that element. The natural abundance is assumed if no other information is given. When working with mass spectrometry, the pattern of peaks in a spectrum will depend on the isotopic pattern of the elements being analyzed.

## allotropes

Allotropes are different forms of the same pure element that differ due to differences in bonding. Common examples are dioxygen,  $\text{O}_2$ , and trioxygen (ozone),  $\text{O}_3$ , for oxygen and graphite, diamond, and fullerenes for carbon. Allotropes of elements with no common names are designated by a prefix:  $\alpha$ ,  $\beta$ ,  $\gamma$  etc, to indicate the different forms. The term allotrope is used for a pure element. The term polymorph is used for the same concept for multi-element compounds (discussed in section IV). Allotropes are mostly of concern for materials science. Once an element is converted to an ion to get it into aqueous solution, the original form of the element is not significant.

### Practice Questions

Use a periodic table and other resources to answer the following questions. Cite your source of information to complete your answer.

1. How many groups are there in the periodic table and what are they?
2. How many elements exist as gases at room temperature? List the molecular form of the ones that are not noble gases.
3. List the elements that exist in the liquid phase at room temperature.
4. What is the naturally occurring element with the highest atomic number? List the two elements of lower atomic number that are not naturally occurring.
5. List the allotropes of tin. Which allotropes are stable at room temperature?

### III. Common Cations

First let me explain how I choose the “common” examples that I use throughout this document. Many of the lighter elements are truly common since they are prevalent in aquatic and biological systems in one form or another. Many transition metals are common in analytical measurements because they are toxic, such as  $\text{Pb}^{2+}$ ,  $\text{Cd}^{2+}$ , and  $\text{Hg}^{2+}$  in Table 1, and must be monitored in food, drinking water, etc. Some of the species we discuss are not common at all in nature, but they are useful chemical reagents in analytical applications.  $\text{Ba}^{2+}$  is relatively rare in the environment, but it is used to precipitate sulfate from solution for analysis. Other cases are rather uncommon in aqueous systems, although found in geologic or technological materials, e.g.,  $\text{Hf}^{4+}$  and  $\text{In}^{3+}$ , but they follow the periodic trend seen in Table 1. Finally, the oxidation states that I list as common are for ions in aqueous solution or in solid compounds. The distribution of atomic, molecular, and ionic species in gas-phase environments such as a heated vapor, flame, or plasma is dependent on the temperature, pressure and other ambient conditions.

The symbol for a monatomic metal ion is simply the element symbol with the ionic charge as a superscript, e.g., lithium ion,  $\text{Li}^+$ , calcium ion,  $\text{Ca}^{2+}$ , and aluminum ion,  $\text{Al}^{3+}$ . Table 1 lists the most common oxidation state for the metallic groups of the periodic table. The group 1 – 4 examples are rarely found in other oxidation states (besides zero). This empirical observation is easy to remember by recalling that stable ions in solution and ionic solids tend to have a filled valence shell containing 8 electrons. The metal ions of groups 5 – 10 tend to have multiple stable oxidation states and no trend is listed here. Likewise, there are numerous exceptions to the general trends seen in Table 1. Copper in group 11 is more often found as  $\text{Cu}^{2+}$  rather than  $\text{Cu}^+$ , mercury in group 12 exists as a stable  $\text{Hg}_2^{2+}$  dimer with an oxidation state of +1, and lead in group 14 is more stable as  $\text{Pb}^{2+}$  than as  $\text{Pb}^{4+}$ .

**Table 1. Most Common Oxidation States of Metal Cations**

group	1	2	3	4	5-10	11	12	13	14
oxidation state	+1	+2	+3	+4	(var)	+1	+2	+3	
examples	H <sup>+</sup> Li <sup>+</sup> Na <sup>+</sup> K <sup>+</sup>	Mg <sup>2+</sup> Ca <sup>2+</sup> Sr <sup>2+</sup> Ba <sup>2+</sup>	Sc <sup>3+</sup> Y <sup>3+</sup> La <sup>3+</sup> Ln <sup>2</sup>	Ti <sup>4+</sup> Zr <sup>4+</sup> Hf <sup>4+</sup>		Ag <sup>+</sup> Au <sup>+</sup>	Zn <sup>2+</sup> Cd <sup>2+</sup> Hg <sup>2+</sup>	Al <sup>3+</sup> Ga <sup>3+</sup> In <sup>3+</sup>	Sn <sup>4+</sup>

There are only two common inorganic polyatomic cations – the mercury dimer and the ammonium ion (see Table 3). Having said that, a large number of organic compounds exist that contain an amine group that can be protonated. Figure 1 shows several examples for illustration. Although these examples deviate from our theme of inorganic nomenclature, I include them here due to the similarity with the ammonium ion and also because many of them are used as pH buffers in analytical protocols.

**Table 2. Examples of Polyatomic Cations**

formula	name
Hg <sub>2</sub> <sup>2+</sup>	mercury(I) ion
NH <sub>4</sub> <sup>+</sup>	ammonium ion
(CH <sub>3</sub> ) <sub>3</sub> N <sup>+</sup>	trimethylammonium ion
C <sub>3</sub> H <sub>5</sub> N <sub>2</sub> <sup>+</sup> or C <sub>3</sub> H <sub>4</sub> N <sub>2</sub> H <sup>+</sup>	imidazolium ion
C <sub>2</sub> H <sub>6</sub> O <sub>2</sub> N <sup>+</sup> or <sup>+</sup> H <sub>2</sub> NCH <sub>2</sub> CO <sub>2</sub> H or HGly <sup>+</sup>	glycinium ion

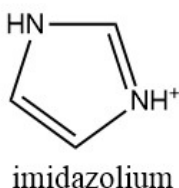
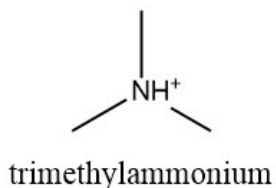


Figure 1. Structural formulas of several protonated amines.

The hydronium ion, H<sub>3</sub>O<sup>+</sup>, is used to represent a proton in water. This species is not isolable as a cation in ionic solid, and I do not include it in the table above. The protonated amines do form solid salts and you will often see the formulas written as the neutral amine followed by a hydrated anion. Examples are discussed in the last section on naming compounds.

<sup>2</sup> Including the ions of the lanthanoid series, with the exception of Ce<sup>4+</sup>.

## IV. Common Anions

Name simple nonmetal anions by adding **-ide** to the end of the element name. Table 4 shows the same trend as for cations, the stable anions have a complete shell of electrons. Most of these particular anions exist only in solids and polyatomic species. Only sulfide and the halides have any appreciable solubility in water as free ions.

**Table 3. Monatomic Nonmetal Ions**

symbol	name	symbol	name	symbol	name	symbol	name
$\text{C}^{4-}$	carbide	$\text{N}^{3-}$	nitride	$\text{O}^{2-}$	oxide	$\text{F}^-$	fluoride
$\text{Si}^{4-}$	silicide	$\text{P}^{3-}$	phosphide	$\text{S}^{2-}$	sulfide	$\text{Cl}^-$	chloride
		$\text{As}^{3-}$	arsenide	$\text{Se}^{2-}$	selenide	$\text{Br}^-$	bromide
				$\text{Te}^{2-}$	telluride	$\text{I}^-$	iodide

Table 5 lists polyatomic anions, which usually end with **-ite** or **-ate**. Note that hydroxide,  $\text{OH}^-$ , is an exception.

**Table 4. Common Anions Containing Nonmetals**

form.	name	form.	name	form.	name	form.	name
$\text{CO}_3^{2-}$	carbonate	$\text{NO}_2^-$	nitrite	$\text{OH}^-$	hydroxide	$\text{ClO}^-$	hypochlorite
$\text{CN}^-$	cyanide	$\text{NO}_3^-$	nitrate	$\text{O}_2^{2-}$	peroxide	$\text{ClO}_2^-$	chlorite
$\text{OC}\equiv\text{N}^-$	cyanate	$\text{PO}_4^{3-}$	phosphate	$\text{S}_2^{2-}$	disulfide	$\text{ClO}_3^{2-}$	chlorate
$\text{SC}\equiv\text{N}^-$	thiocyanate	$\text{AsO}_3^{3-}$	arsenite <sup>3</sup>	$\text{SO}_3^{2-}$	sulfite	$\text{ClO}_4^-$	perchlorate
$\text{SiO}_4^{4-}$	silicate	$\text{AsO}_4^{3-}$	arsenate	$\text{SO}_4^{2-}$	sulfate	$\text{BrO}_3^{2-}$	bromate
						$\text{IO}_3^{2-}$	iodate

These polyatomic anions exist as stable molecular entities with strong covalent bonds between the oxygen atoms and the central atom. They can change in redox reactions, but they do not dissociate when they are simply dissolved in water. We will discuss protonated forms below, but a common error is to take a molecule like  $\text{HClO}_2$  in water and think that it forms  $\text{ClO}^-$  and  $\text{OH}^-$  in solution. This dissociation does not occur, and the tip-off should be the lack of charge balance in the reaction:



A neutral species cannot dissociate into two anions.

<sup>3</sup> Often called orthoarsenite to distinguish it from the metaarsenite anion,  $\text{AsO}_2^-$ , which occurs in polymeric solids. In drinking water analysis, arsenite and arsenate species are often referred to as As(III) and As(V) respectively, disregarding the extent of protonation.

## polymers

Usually when we talk about polymers we are referring to large organic molecules that contain repeat units. Numerous inorganic species will also form polymeric chains or rings. Phosphoric acid,  $\text{H}_3\text{PO}_4$ , is also called orthophosphoric acid to distinguish it from polymeric forms. Multiple phosphoric acid molecules can combine with loss of  $\text{H}_2\text{O}$  to form pyrophosphoric acid,  $\text{H}_4\text{P}_2\text{O}_7$ , and higher-order polyphosphoric acids. The polymeric forms will exist in concentrated solution. An analyst must be aware of these possible species to avoid quantitative errors in some types of measurements.

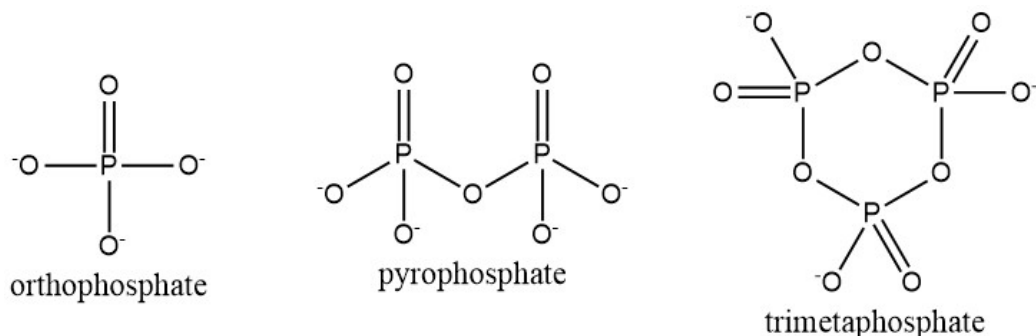


Figure 2. Structural formulas of the unprotonated forms of phosphate, a phosphate dimer, and a cyclic phosphate trimer.

Although not polymers, sulfur oxyanions likewise form more complicated formulas.

Some examples, which find use as reagents in redox reactions, include:

- thiosulfate ion:  $\text{S}_2\text{O}_3^{2-}$  or  $\text{SO}_3\text{S}^{2-}$
- metabisulfite ion:  $\text{S}_2\text{O}_5^{2-}$
- peroxydisulfate ion:  $\text{S}_2\text{O}_8^{2-}$

## polymorphs

A somewhat related concept is a polymorph, which is a different structural form of a solid substance. Pure calcium carbonate,  $\text{CaCO}_3$ , has three polymorphs: calcite, aragonite, and vaterite. Dissolving any of these polymorphs with acid will produce the same species in solution, but these different polymorphs can vary in solid-state properties such as solubility, mineral habit, and refractive index. These physical differences have little impact on making quantitative measurements in solution, but they can be very important in geochemistry, aquatic systems, pharmaceutical products, and materials science.

## V. Common Anions Containing Metals

The following table lists the common polyatomic anions that contain metals. Most metals have fewer oxidation states than the non-metals seen in the previous section, so most metal containing polyatomic anions end with **-ate**. These polyatomic species exist as distinct entities that do not dissociate when dissolved in water. The heptamolybdate ion might seem unusual, but ammonium heptamolybdate is a soluble salt that is used as a reagent to test for phosphate in water.

**Table 5. Common Anions Containing Metals**

symbol	name
$\text{MnO}_4^-$	permanganate ion
$\text{CrO}_4^{2-}$	chromate ion
$\text{Cr}_2\text{O}_7^{2-}$	dichromate ion
$\text{Mo}_7\text{O}_{24}^{6-}$	heptamolybdate ion

There are numerous other oxyanions, but most of them form insoluble compounds. As above, these anions usually ends with the **-ate** suffix, including some other anionic forms, not only for oxyanions. Unless the metal oxidation state is given, you usually just need to look up the formula of these species. Different oxidation states or polymeric forms will have prefixes to indicate a difference. The two simplest vanadates for example are orthovanadate,  $\text{VO}_4^{3-}$ , and pyrovanadate,  $\text{V}_2\text{O}_7^{4-}$ .

## VI. Common Inorganic Ions in Protonated Form

Just as a base can accept a proton to become a cation, many multivalent anions can accept a proton to form a new species with a charge one greater than the parent anion. It is equivalent to also form these species by neutralizing a proton from the fully protonated acid. These intermediate forms are called amphiprotic species because they can either accept or lose a proton. Table 7 lists examples of inorganic amphiprotic ions. Chapter 6 of the text discusses numerous examples of polyprotic organic weak acids, which have the same general form. I write the hydrogen separated from the anion name for clarity, but IUPAC recommends writing them together, e.g., hydrogensulfite ion.



**Table 6. Common Anions Containing Metals**

symbol	name	common name
$\text{HCO}_3^-$	hydrogen carbonate ion	bicarbonate ion
$\text{HSO}_3^-$	hydrogen sulfite ion	bisulfite ion
$\text{HSO}_4^-$	hydrogen sulfate ion	bisulfate ion
$\text{H}_2\text{PO}_4^-$	dihydrogen phosphate ion	
$\text{HPO}_4^{2-}$	hydrogen phosphate ion	

## VII. Common nonmetal Compounds

The name of a nonmetallic compound is constructed from the name of the first element in the formula followed by the name for the other elements in the compound, modified as discussed in Sections IV through VI. A prefix indicates the number of each atom in the compound, although mono- is often omitted in unambiguous cases. Table 8 lists numerous examples of covalently bound molecules, which exist as distinct entities in the gas phase. As the last two examples illustrate, many reagents go by common names rather than the systematic name. In the case of HCl, it is typical to use hydrogen chloride when talking about the gas and hydrochloric acid for the condensed phase.

Simple gases such as  $\text{N}_2$  and  $\text{O}_2$  dissolve in water, but do not react with water. Most of the other gases react with water either partially or completely to form other species in aqueous solution. As a typical example, dissolved  $\text{CO}_2$  forms carbonic acid,  $\text{H}_2\text{CO}_3$ , which then loses a proton to form  $\text{H}^+$  and  $\text{HCO}_3^-$  in solution. This reaction leads to pure water in equilibrium with the atmosphere having a pH of approximately 5.6 rather than 7.0.

**Table 7. Nonmetal Compounds**

formula	systematic name	common name
CO	carbon monoxide	
$\text{CO}_2$	carbon dioxide	
$\text{SO}_2$	sulfur dioxide	
$\text{SO}_3$	sulfur trioxide	
$\text{N}_2$	dinitrogen	nitrogen
$\text{N}_2\text{O}_5$	dinitrogen pentoxide	
HCl	hydrogen chloride	hydrochloric acid
$\text{HNO}_3$		nitric acid

## VIII. Putting Cations and Anions Together

The formulas of ionic compounds are written with the cation symbol(s) followed by the anion symbol. The stoichiometry is specified with subscripts so that the ionic solid is neutral overall. The names of ionic compounds follow the same order with the name of the cation followed by name of the anion. Although I usually specified “ion” in the name of cations and anions listed above, we drop the “ion” name in compounds, e.g., NaCl is sodium chloride, not sodium ion chloride ion.

When writing the names of ionic compounds, we use a Roman numeral in parentheses to specify the oxidation number of metals that can exist in different oxidation states. The oxidation state of the metal is the positive value of this oxidation number. We usually omit this designation for the metals that have only one common ionic form. Table 8 lists some examples and older nomenclature that is still used by some chemical suppliers. Unlike the nonmetal cases, we usually do not add a numerical prefix to the anion. It is not incorrect to do so, but it is redundant when we have specified the metal oxidation state.

If there are multiple cations, we write them sequentially. The cations are listed alphabetically, but it is common to list hydrogen after the other cations, e.g., NaHSO<sub>3</sub> is sodium hydrogen sulfite, sodium hydrogensulfite, or sodium bisulfite.

**Table 8. Examples of Written Names**

formula	write as	read as	obsolete terminology
FeCl <sub>2</sub>	iron(II) chloride	iron two chloride	ferrous chloride
FeCl <sub>3</sub>	iron(III) chloride	iron three chloride	ferric chloride
Cu <sub>2</sub> O	copper(I) oxide	copper one oxide	cuprous oxide
Cu(NO <sub>3</sub> ) <sub>2</sub>	copper(II) nitrate	copper two nitrate	cupric nitrate

Inorganic solids often have waters of hydration in the crystal structure. These water molecules affect the formula weight of an ionic solid and must be indicated in the formula. Place a center dot and the number of water molecules, e.g.,  $\cdot x\text{H}_2\text{O}$ , after the formula. The written names have the Greek prefix and hydrate appended at the end.

As noted in Section III, the salts of protonated bases are often named with the base name followed by a hydrated anion. Several examples are shown below.

**Table 9. Examples of Polyatomic Cations**

systematic name	common name	formula
imidazolium chloride	imidazole hydrochloride	$\text{C}_3\text{H}_4\text{N}_2 \cdot \text{HCl}$
glycinium chloride	glycine hydrochloride	$\text{H}_2\text{NCH}_2\text{CO}_2\text{H} \cdot \text{HCl}$ or $\text{C}_2\text{H}_5\text{O}_2\text{N} \cdot \text{HCl}$
---	tris HCl	$\text{NH}_2\text{C}(\text{CH}_2\text{OH})_3 \cdot \text{HCl}$
Where tris is an abbreviation for tris(hydroxymethyl)aminomethane.		

### Practice Questions

- Write out the name of the following compounds.
  - $\text{CaF}_2$
  - $(\text{NH}_4)_2\text{SO}_4$
  - $\text{NaH}_2\text{PO}_4$
  - $\text{NaAsO}_2$
  - $\text{Pb}(\text{CH}_3\text{COO})_2 \cdot 3\text{H}_2\text{O}$
  - $\text{Cu}(\text{NO}_3)_2 \cdot 3\text{H}_2\text{O}$
  - $\text{CuSO}_4$
  - $\text{Na}_2\text{HAsO}_4 \cdot 7\text{H}_2\text{O}$
- Write the chemical formula for each of the following chemicals. You will need a reference source to find the formulas of the organic ions.
  - mercury(I) chloride (common name: calomel)
  - lithium perchlorate
  - ammonium sodium phosphate dibasic tetrahydrate
  - magnesium citrate
  - boron nitride (predict the charge on boron from a periodic table)
  - pseudoephedrine hydrochloride

## IX. Answers to Practice Exercises

### Pure Elements

Use a periodic table and other resources to answer the following questions. Cite your source of information to complete your answer.

1. How many groups are there in the periodic table and what are they?

There are eighteen groups in the periodic table, numbered sequentially from group 1 to group 18. The columns in the lanthanide and actinide series do not have group numbers. The 18 groups arise from filling the s, p, and d orbitals, which contain 2, 6, and 10 electrons, respectively. [source: Atomic Properties of the Elements, NIST SP 966 (2010); <http://www.nist.gov/pml/data/periodic.cfm>.]

2. How many elements exist as gases at room temperature? List the molecular form of the ones that are not noble gases.

Eleven elements exist as gases at room temperature. The molecular ones are H<sub>2</sub>, N<sub>2</sub>, O<sub>2</sub>, F<sub>2</sub>, Cl<sub>2</sub>. The other six gases are the noble gases in group 18. [source: Atomic Properties of the Elements, NIST SP 966 (2010); <http://www.nist.gov/pml/data/periodic.cfm>.]

3. List the elements that exist in the liquid phase at room temperature.

Bromine, Br<sub>2</sub>, and mercury, Hg. [source: Atomic Properties of the Elements, NIST SP 966 (2010); <http://www.nist.gov/pml/data/periodic.cfm>.]

4. What is the naturally occurring element with the highest atomic number? List the two elements of lower atomic number that are not naturally occurring.

Uranium, U, has the highest atomic number of the elements that exist in nature. Technetium, Tc, and Promethium, Pm, have lower atomic numbers but do not occur in nature. [source: Atomic Properties of the Elements, NIST SP 966 (2010); <http://www.nist.gov/pml/data/periodic.cfm>.]

5. List the allotropes of tin. Which allotrope is stable at room temperature?

The four allotropes are  $\alpha$ -tin (gray tin),  $\beta$ -tin (white tin),  $\gamma$ -tin (rhombic tin), and  $\sigma$ -tin (sigma tin). Only  $\alpha$ -tin and  $\beta$ -tin are stable at room temperature.

## Putting Cations and Anions Together

1. Write out the name of the following compounds.

- a.  $\text{CaF}_2$   
calcium fluoride  
[Writing calcium difluoride is also correct. Since the calcium ion is always +2, we usually leave off the prefix.]
- b.  $(\text{NH}_4)_2\text{SO}_4$   
ammonium sulfate
- c.  $\text{NaH}_2\text{PO}_4$   
sodium dihydrogen phosphate
- d.  $\text{NaAsO}_2$   
sodium (meta)arsenite
- e.  $\text{Pb}(\text{CH}_3\text{COO})_2 \cdot 3\text{H}_2\text{O}$   
lead acetate trihydrate, lead(II) acetate trihydrate, or lead diacetate trihydrate  
[Since the lead ion is almost always +2, any of these names are acceptable.]
- f.  $\text{Cu}(\text{NO}_3)_2 \cdot 3\text{H}_2\text{O}$   
copper(II) nitrate trihydrate  
[Since copper can have different oxidation states, we use the Roman numerals to specify the oxidation state. Less common, but correct, is the name “copper dinitrate trihydrate”.]
- g.  $\text{CuSO}_4$   
copper(II) sulfate anhydrous  
[Many metal salts contain waters of hydration in the crystal structure. Salts that are unhydrated are often specified as “anhydrous” or “basic”.]
- h.  $\text{Na}_2\text{HAsO}_4 \cdot 7\text{H}_2\text{O}$   
disodium hydrogen arsenate heptahydrate  
[Chemical suppliers will often use names like “sodium arsenate dibasic heptahydrate” where the dibasic indicates a –2 charge when the anion is dissolved.]

2. Write the chemical formula for each of the following chemicals. You will need a reference source to find the formulas of the organic ions.
- a. mercury(I) chloride (common name: calomel)  
 $\text{Hg}_2\text{Cl}_2$
  - b. lithium perchlorate  
 $\text{LiClO}_4$
  - c. ammonium sodium phosphate dibasic tetrahydrate  
 $\text{NH}_4\text{NaHPO}_4 \cdot 4\text{H}_2\text{O}$
  - d. magnesium citrate  
 $\text{Mg}(\text{C}_6\text{H}_5\text{O}_7)$  or  $\text{C}_6\text{H}_5\text{MgO}_7$   
[The name magnesium citrate is ambiguous, but is common for the 1:1 complex.  $\text{Mg}_3(\text{C}_6\text{H}_5\text{O}_7)_2$ , which is not as common in applications, is called “trimagnesium citrate” or “citric acid, magnesium salt (2:3)”.]
  - e. boron nitride  
 $\text{BN}$   
[Boron is above aluminum in the periodic table (group 13), and we expect the ionic form to be  $\text{B}^{3+}$ .]
  - f. pseudoephedrine hydrochloride  
 $\text{C}_{10}\text{H}_{16}\text{ClNO}$  or  $\text{C}_{10}\text{H}_{15}\text{NO} \cdot \text{HCl}$