

Chemical Nomenclature

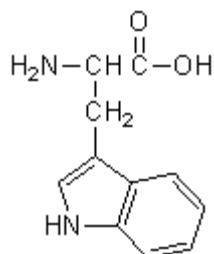
I. The reason that we care about naming compounds

- a. Naming, a.k.a. “nomenclature” of compounds
- b. Significance: must know what we are talking about very clearly when we refer to substances by their names or formulas.
- c. For instance, what is “carbon oxide”? There are two different formulas for a carbon-oxygen compound, both of which obey the octet rule, CO and CO₂. Which one is “carbon oxide”?
Answer: neither. One is carbon monoxide, the other is carbon dioxide.
- d. When we say sodium chloride, what is the formula? This is easy, because we know that Na forms 1+ ions, and Cl will usually form 1- ions. Thus, NaCl.
- e. When we say “iron chloride,” what is this compound’s formula? Well, iron is a transition metal, and it turns out that iron can form Fe²⁺ ions, but can also form Fe³⁺ ions.
 - i. Fe – [Ar]4s²3d⁶ -- can lose two electrons from the 4s to become simply [Ar]3d⁶, which is the electron configuration for Fe²⁺.
 - ii. Alternatively, Fe can become stable by losing both 4s electrons and one from the 3d, leaving it with a rather stable half-filled d sublevel. Fe³⁺ has an electron configuration of [Ar]3d⁵.
 - iii. Transition metals often become stable by obtaining a **pseudo noble gas configuration**. That is, they do not attain a full highest-occupied **energy level** in their outer shells (i.e., do not attain complete octets). Instead, they attain full or half-full **sublevels**.
- f. So, “iron chloride” could be FeCl₂ or FeCl₃. We need to have a way of distinguishing these compounds from one another. As it turns out, FeCl₂ is called iron (II) chloride and the FeCl₃ is iron (III) chloride.

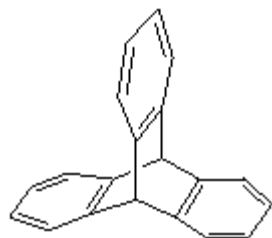
II. The goals of this chapter:

- a. You should be able to name a compound, given its chemical formula. Example: N₂O₄ is dinitrogen tetroxide.
- b. You should be able to write the chemical formula for a compound if you are given the name of that compound. Example: manganese (IV) oxide is MnO₂.
- c. We will only concern ourselves with **simple** ionic and covalent (i.e., molecular) compounds. (It is a rare chemist who can name any compound under the sun – organic compounds, proteins, organometallic compounds, steroids, etc. – without reviewing at least the rules for that particular class of compounds first.)

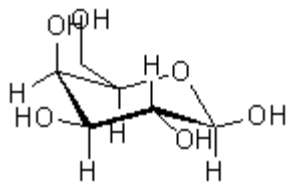
Examples of compounds that you will NOT have to name! ☺



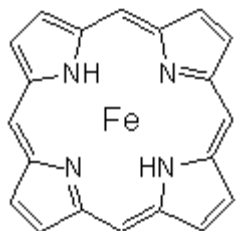
An example of an amino acid.



A propellane, an example of an aromatic compound.



A hexose, which is an example of a carbohydrate.



An example of a porphyrin.

III. Naming binary ionic compounds

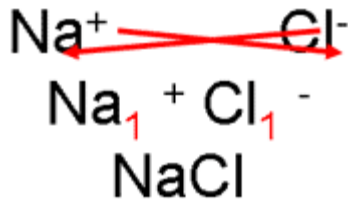
a. Situations in which there is not a “special metal”

- If the metal involved always forms the same charge, then there is no need to specify its charge in the compound's name. Examples: all group I and group II metals, aluminum, Zn, Ag, Cd. Zn always forms Zn^{2+} , Ag always forms Ag^+ , Cd always forms Cd^{2+} .
- The anion takes an -ide ending.
- Examples:
 - NaCl is sodium chloride
 - CaCl_2 is calcium chloride
 - AlCl_3 is aluminum chloride

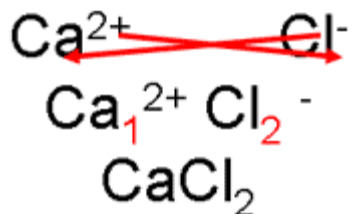
b. Writing formulas for the above compounds

- Write down the ions, cross over the charges.
- “Cancel” these subscripts, if necessary, just as you would a fraction.

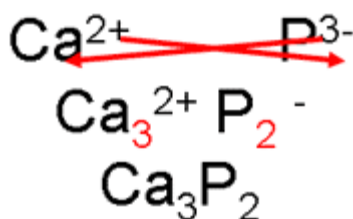
sodium chloride



calcium chloride



calcium phosphide



- c. Situations in which there is a “special metal”, i.e., the metal is a transition metal or a group 14 metal.
- The bad news about these metals is that it is very difficult for a 1st year chemistry student to figure out how a transition metal forms an ion.
 - For instance, iron (Fe) has the electron configuration [Ar]4s²3d⁶. It will not lose 8 electrons to gain the electron configuration of the noble gas Ar. Because it is a metal, it certainly will not gain electrons, and at any rate it would have to gain 10 electrons to have the full octet of Kr.
 - Instead, transition metals and group 14 metals will attain a **pseudo noble gas configuration**, in which the metals become stable ions by losing enough electrons to be left with either a full sublevel, or a half-filled sublevel. Full sublevels and half-filled sublevels are stable, though not nearly as stable as filled highest-occupied energy levels (i.e., filled valence shells).
 - These metals usually have more than one possible ion that they can form.
 - Fe will lose either two electrons to become Fe²⁺ as [Ar]4s¹3d⁵, or it will lose three electrons to become Fe³⁺ as [Ar]4s⁰3d⁵.
 - Ions you should be able to recognize and look up on a table are listed below.

Common Metal Ions		
Ion	Systematic Name	Common Name
Fe ²⁺	iron (II)	ferrous
Fe ³⁺	iron (III)	ferric
Cu ⁺	copper (I)	cuprous
Cu ²⁺	copper (II)	cupric

Pb²⁺	lead (II)	plumbous
Pb⁴⁺	lead (IV)	plumbic
Cr²⁺	chromium (II)	chromous
Cr³⁺	chromium (III)	chromic
Sn²⁺	tin (II)	stannous
Sn⁴⁺	tin (IV)	stannic
Co²⁺	cobalt (II)	cobaltous
Co³⁺	cobalt (III)	cobaltic
Hg₂²⁺	mercury (I)	mercurous
Hg²⁺	mercury (II)	mercuric

**** ALWAYS: Zn²⁺, Ag⁺, Cd²⁺ ****

- vii. The **stock system** uses roman numerals in parentheses to indicate the type of ion in the compound. Example: FeCl₂ is iron (II) chloride, FeCl₃ is iron (III) chloride. You are responsible only for the stock system of naming compounds containing these “special metal” ions.
- viii. The classical system of naming uses the “-ous” and “-ic” prefixes to indicate the lower and higher charges, respectively. Example: iron (II) chloride is **ferrous** chloride and iron (III) chloride is **ferric** chloride. You are not responsible for using or recognizing the classical nomenclature system, though it will be helpful in the lab, understanding chemical names on packaging, etc.
- d. Writing formulas for compounds that contain polyatomic ions.
 - i. The polyatomic ions that you should **recognize**, but which you **do not need to memorize**:

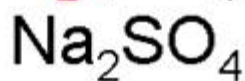
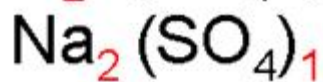
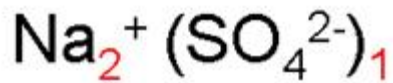
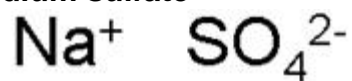
Common Polyatomic Ions							
+1		-1		-2		-3	
NH₄⁺	ammonium	C₂H₃O₂⁻	acetate	CO₃²⁻	carbonate	PO₄³⁻	phosphate
H₃O⁺	hydronium	ClO⁻	hypochlorite	CrO₄²⁻	chromate	PO₃³⁻	phosphite
		ClO₂⁻	chlorite	Cr₂O₇²⁻	dichromate		
		ClO₃⁻	chlorate	SO₄²⁻	sulfate		
		ClO₄⁻	perchlorate	SO₃²⁻	sulfite		
		CN⁻	cyanide	O₂²⁻	peroxide		
		NO₃⁻	nitrate	C₂O₄²⁻	oxalate		
		NO₂⁻	nitrite				
		HCO₃⁻	hydrogen carbonate (bicarbonate)				
		OH⁻	hydroxide				
		MnO₄⁻	permanganate				

- ii. Notice that ions formed from nonmetals take the “-ide” ending. Usually, the converse is true: “ide” ions are formed from nonmetals. Notable exceptions: cyanide (CN⁻) and hydroxide

(OH⁻).

iii. These examples are similar to the examples using monatomic (one-atom ions).

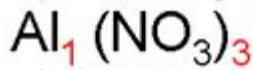
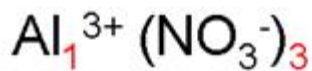
sodium sulfate



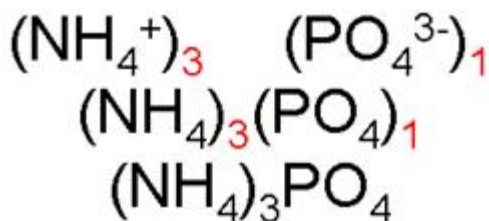
calcium sulfate



aluminum nitrate



ammonium phosphate



IV. Naming covalent compounds (a.k.a. molecular compounds)

- Remember, covalent bonds are formed between nonmetals and nonmetals
- We will only name binary molecular compounds. Binary molecular compounds only contain two different elements.
- Consider these two compounds that can be formed from carbon and oxygen: CO and CO₂.
- Using the word “carbon oxide” would not be sufficient, because there are two different forms of “carbon oxide.”
- Because carbon is not an ion in these compounds, it does not have a charge. That is one way to rationalize to yourself why we do not use the charge on carbon in parentheses (as we do when naming transition metal-containing ionic compounds).
- For covalent compounds, we use prefixes to indicate the number of each atom in a compound.
 - The first word in the name gets no suffix, but the second word in the name gets the “ide” suffix, similar to how we write names for binary ionic compounds.
 - Never use “mono” if there is one atom of the first element. Ex: carbon dioxide = CO₂, *not* “monocarbon dioxide.”
 - DO use the appropriate prefix to indicate the number of atoms of the first element if there IS more than one atom of that element. Example: P₂O₅ is diphosphorus pentoxide.
 - The suffixes are shown below.

Prefixes Used in Naming Binary Molecular Compounds	
Prefix	Number
mono	1
di	2
tri	3
tetra	4
penta	5
hexa	6
hepta	7
octa	8
nona	9
deca	10

V. Naming acids

- For our purposes, for right now, acids are compounds that start with one or more “H’s” and end with some anion. “HX”, H₂A”, “H₃D”, etc., where X⁻, A²⁻, and D³⁻ are anions attached to one or

more hydrogens.

b. We will consider three cases of acids:

- i. Acids which contain an anion that ends with “-ide”
- ii. Acids which contain an anion that ends with “-ite”
- iii. Acids which contain an anion that ends with “-ate”

c. Acids which contain an anion that ends with “-ide”

- i. Examples of ions that end in “ide”: cyanide, chloride, bromide, iodide, sulfide.
- ii. First, chop off the suffix “ide”
- iii. Then, add the prefix “hydro”, write the stem, then add the suffix “-ic acid”
- iv. HCN

1. “hydrogen cyanide”
2. Cyanide
3. Cyan-
4. Hydrocyanic acid

v. HCl

1. “hydrogen chloride”
2. Chloride
3. Chlor-
4. Hydrochloric acid

vi. HBr

1. “hydrogen bromide”
2. Bromide
3. Brom-
4. Hydrobromic acid

vii. HI

1. “hydrogen iodide”
2. Iodide
3. Iod-
4. Hydroiodic acid

viii. H₂S

1. “hydrogen sulfide”
2. Sulfide
3. Sulf-
4. Hydrosulfuric acid
5. Note the weird stem change. Same thing happens with “phosphide”

d. Acids that contain an anion that ends with “-ite”

- i. Examples: sulfite, chlorite, nitrite, hypochlorite.
- ii. First chop off the suffix “ite”.
- iii. Next, write down the stem plus the suffix “-ous acid”
- iv. H₂SO₃

1. “hydrogen sulfite”
2. Sulf-
3. Sulfurous acid
4. Note the weird stem change. Similar thing happens with phosphite ion.

v. HClO₂

1. “hydrogen chlorite”
2. Chlor-
3. Chlorous acid

vi. HNO₂

- 1. “hydrogen nitrite”
 - 2. Nitr-
 - 3. Nitrous acid
- vii. HClO
 - 1. “hydrogen hypochlorite”
 - 2. Hypochlor-
 - 3. Hypochlorous acid
- e. Acids that contain an anion that ends with “ate”
 - i. Examples: chromate, dichromate, nitrate, sulfate
 - ii. First, chop off the suffix “-ate”
 - iii. Then, write down the stem + “ic acid”
 - iv. H_2CrO_4
 - 1. “hydrogen chromate”
 - 2. Chrom-
 - 3. Chromic acid
 - v. $\text{H}_2\text{Cr}_2\text{O}_7$
 - 1. “hydrogen dichromate”
 - 2. Dichrom- (note that “di” is actually a part of the dichromate ion’s name and is therefore not omitted)
 - 3. Dichromic acid
 - vi. HNO_3
 - 1. “hydrogen nitrate”
 - 2. Nitr-
 - 3. Nitric acid
 - vii. H_2SO_4
 - 1. “hydrogen sulfate”
 - 2. Sulf-
 - 3. Sulfuric acid. Note the weird stem change. Same thing happens with phosphoric acid (from phosphate ion)

Rules For Naming Acids			
Anion ending	Example	Acid name	Example
-ide	Cl^- chloride	hydro-(stem)-ic acid	hydrochloric acid
-ite	SO_3^{2-} sulfite	(stem)-ous acid	sulfurous acid
-ate	NO_3^- nitrate	(stem)-ic acid	nitric acid