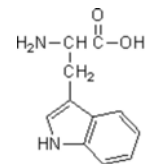


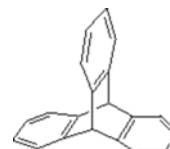
### Chemical Nomenclature

- 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 CO<sub>2</sub>. 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 Fe<sup>2+</sup> ions, but can also form Fe<sup>3+</sup> ions.
    - i. Fe – [Ar]4s<sup>2</sup>3d<sup>6</sup> -- can lose two electrons from the 4s to become simply [Ar]3d<sup>6</sup>, which is the electron configuration for Fe<sup>2+</sup>.
    - 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. Fe<sup>3+</sup> has an electron configuration of [Ar]3d<sup>5</sup>.
    - 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 FeCl<sub>2</sub> or FeCl<sub>3</sub>. We need to have a way of distinguishing these compounds from one another. As it turns out, FeCl<sub>2</sub> is called iron (II) chloride and the FeCl<sub>3</sub> is iron (III) chloride.
- II. The goals of this chapter:
  - a. You should be able to name a compound, given its chemical formula. Example: N<sub>2</sub>O<sub>4</sub> 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 MnO<sub>2</sub>.
  - 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.)

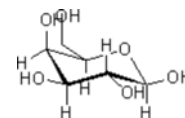
### Examples of compounds that you will NOT have to name! J



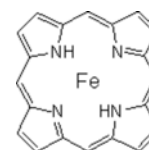
An example of an amino acid.



A propellane, an example of an aromatic compound.



A hexose, which is an example of a carbohydrate.

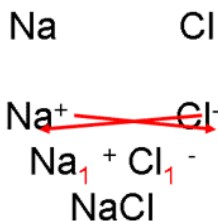


An example of a porphyrin.

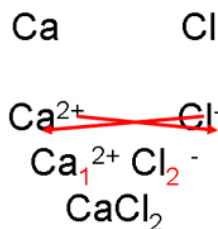
## III. Naming binary ionic compounds

- a. Situations in which there is not a "special metal"
  - i. 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  $Cd^{2+}$ .
  - ii. The anion takes an -ide ending.
  - iii. Examples:
    1. NaCl is sodium chloride
    2.  $CaCl_2$  is calcium chloride
    3.  $AlCl_3$  is aluminum chloride
- b. Writing formulas for the above compounds
  - i. Write down the ions, cross over the charges.
  - ii. "Cancel" these subscripts, if necessary, just as you would a fraction.

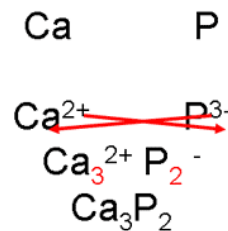
sodium chloride



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.
  - i. The bad news about these metals is that it is very difficult for a 1<sup>st</sup> year chemistry student to figure out how a transition metal forms an ion.
  - ii. For instance, iron (Fe) has the electron configuration  $[Ar]4s^23d^6$ . 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.
  - iii. 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).
  - iv. These metals usually have more than one possible ion that they can form.
  - v. Fe will lose either two electrons to become  $Fe^{2+}$  as  $[Ar]4s^13d^5$ , or it will lose three electrons to become  $Fe^{3+}$  as  $[Ar]4s^03d^5$ .
  - vi. Ions you should be able to recognize and look up on a table are listed below.

Common Metal Ions		
Ion	Systematic Name	Common Name
$\text{Fe}^{2+}$	iron (II)	ferrous
$\text{Fe}^{3+}$	iron (III)	ferric
$\text{Cu}^+$	copper (I)	cuprous
$\text{Cu}^{2+}$	copper (II)	cupric
$\text{Pb}^{2+}$	lead (II)	plumbous
$\text{Pb}^{4+}$	lead (IV)	plumbic
$\text{Cr}^{2+}$	chromium (II)	chromous
$\text{Cr}^{3+}$	chromium (III)	chromic
$\text{Sn}^{2+}$	tin (II)	stannous
$\text{Sn}^{4+}$	tin (IV)	stannic
$\text{Co}^{2+}$	cobalt (II)	cobaltous
$\text{Co}^{3+}$	cobalt (III)	cobaltic
$\text{Hg}_2^{2+}$	mercury (I)	mercurous
$\text{Hg}^{2+}$	mercury (II)	mercuric

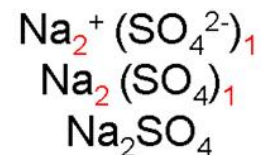
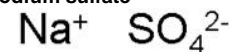
\*\* ALWAYS:  $\text{Zn}^{2+}$ ,  $\text{Ag}^+$ ,  $\text{Cd}^{2+}$  \*\*

- vii. The **stock system** uses roman numerals in parentheses to indicate the type of ion in the compound. Example:  $\text{FeCl}_2$  is iron (II) chloride,  $\text{FeCl}_3$  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**:

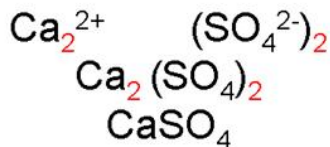
Common Polyatomic Ions							
+1		-1		-2		-3	
$\text{NH}_4^+$	ammonium	$\text{C}_2\text{H}_3\text{O}_2^-$	acetate	$\text{CO}_3^{2-}$	carbonate	$\text{PO}_4^{3-}$	phosphate
$\text{H}_3\text{O}^+$	hydronium	$\text{ClO}^-$	hypochlorite	$\text{CrO}_4^{2-}$	chromate	$\text{PO}_3^{3-}$	phosphite
		$\text{ClO}_2^-$	chlorite	$\text{Cr}_2\text{O}_7^{2-}$	dichromate		
		$\text{ClO}_3^-$	chlorate	$\text{SO}_4^{2-}$	sulfate		
		$\text{ClO}_4^-$	perchlorate	$\text{SO}_3^{2-}$	sulfite		
		$\text{CN}^-$	cyanide	$\text{O}_2^{2-}$	peroxide		
		$\text{NO}_3^-$	nitrate	$\text{C}_2\text{O}_4^{2-}$	oxalate		
		$\text{NO}_2^-$	nitrite				
		$\text{HCO}_3^-$	hydrogen carbonate (bicarbonate)				
		$\text{OH}^-$	hydroxide				
		$\text{MnO}_4^-$	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 ( $\text{CN}^-$ ) and hydroxide ( $\text{OH}^-$ ).
- iii. These examples are similar to the examples using monatomic (one-atom ions).

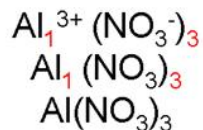
sodium sulfate



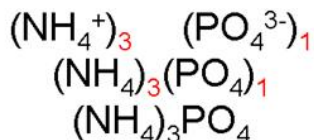
calcium sulfate



## 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 CO<sub>2</sub>.
- 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 = CO<sub>2</sub>, *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: P<sub>2</sub>O<sub>5</sub> 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", H<sub>2</sub>A", "H<sub>3</sub>D", etc., where X<sup>-</sup>, A<sup>2-</sup>, and D<sup>3-</sup> are anions attached to one or more hydrogens.
- We will consider three cases of acids:
  - Acids which contain an anion that ends with "-ide"
  - Acids which contain an anion that ends with "-ite"
  - Acids which contain an anion that ends with "-ate"

- c. Acids which contain an anion that ends with "-ide"
- Examples of ions that end in "ide": cyanide, chloride, bromide, iodide, sulfide.
  - First, chop off the suffix "ide"
  - Then, add the prefix "hydro", write the stem, then add the suffix "-ic acid"
- iv. HCN
- "hydrogen cyanide"
  - Cyanide
  - Cyan-
  - Hydrocyanic acid
- v. HCl
- "hydrogen chloride"
  - Chloride
  - Chlor-
  - Hydrochloric acid
- vi. HBr
- "hydrogen bromide"
  - Bromide
  - Brom-
  - Hydrobromic acid
- vii. HI
- "hydrogen iodide"
  - Iodide
  - Iod-
  - Hydroiodic acid
- viii. H<sub>2</sub>S
- "hydrogen sulfide"
  - Sulfide
  - Sulf-
  - Hydrosulfuric acid
  - Note the weird stem change. Same thing happens with "phosphide"
- d. Acids that contain an anion that ends with "-ite"
- Examples: sulfite, chlorite, nitrite, hypochlorite.
  - First chop off the suffix "ite."
  - Next, write down the stem plus the suffix "-ous acid"
- iv. H<sub>2</sub>SO<sub>3</sub>
- "hydrogen sulfite"
  - Sulf-
  - Sulfurous acid
  - Note the weird stem change. Similar thing happens with phosphite ion.
- v. HClO<sub>2</sub>
- "hydrogen chlorite"
  - Chlor-
  - Chlorous acid
- vi. HNO<sub>2</sub>
- "hydrogen nitrite"
  - Nitr-
  - Nitrous acid
- vii. HClO
- "hydrogen hypochlorite"

- Hypochlor-
  - Hypochlorous acid
- e. Acids that contain an anion that ends with "-ate"
- Examples: chromate, dichromate, nitrate, sulfate
  - First, chop off the suffix "-ate"
  - Then, write down the stem + "ic acid"
- iv. H<sub>2</sub>CrO<sub>4</sub>
- "hydrogen chromate"
  - Chrom-
  - Chromic acid
- v. H<sub>2</sub>Cr<sub>2</sub>O<sub>7</sub>
- "hydrogen dichromate"
  - Dichrom- (note that "d" is actually a part of the dichromate ion's name and is therefore not omitted)
  - Dichromic acid
- vi. HNO<sub>3</sub>
- "hydrogen nitrate"
  - Nitr-
  - Nitric acid
- vii. H<sub>2</sub>SO<sub>4</sub>
- "hydrogen sulfate"
  - Sulf-
  - 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 <sup>-</sup> chloride	hydro-(stem)-ic acid	hydrochloric acid
-ite	SO <sub>3</sub> <sup>2-</sup> sulfite	(stem)-ous acid	sulfurous acid
-ate	NO <sub>3</sub> <sup>-</sup> nitrate	(stem)-ic acid	nitric acid