

## AN INTRO TO THE DESCRIPTIVE CHEMISTRY

To this point in our review we've worked on methods for establishing a **QUANTITATIVE** foundation in chemistry. We can now manipulate all manner of unit factors to solve problems involving amounts of chemical materials, and while there was some modest requirement that you understand the nature of the substances involved in the problems, you might just as well have applied the techniques you learned to solve problems associated with preparing a good white sauce—make the units cancel and you've learned the right answer—whatever those **numbers** meant. There was nothing particularly relevant to chemistry in the problem solving. Well now there is a radical change in material. For one thing, there aren't any problems to work. Reading this material is like taking a botany class. Suddenly our primary focus is on **DESCRIBING** chemistry. So when you get down to it, in this chapter there is: a whole lot of **MEMORIZING** going on. But I will also introduce you to the basic theoretical foundation that **significantly reduces memorization** by learning the following basic concepts.

### Concept 1: Periodic Trends Like Metallic Character

- You learn basic trends in the elements that prompted the creation of a periodic table of those elements. The first is which elements are metals and which are not?

### Concept 2: Acids Base Chemistry

- You will learn what things do and don't behave as acid and bases.

### Concept 3: Which Ions dissolve in Water?

- You will learn elementary solubility rules for ions in water.

### Concept 4: Oxidation/Reduction

- You will learn to assign oxidation numbers to elements. The oxidation number is your first indication of where **ELECTRON DENSITY** is located in chemical compounds.

### Concept 5: Active Metals

- You will learn about displacement reactions, which means you will learn to predict what happens when you dump metals into water.

### Concept 6: Chemical Nomenclature

- You will learn the common way to name binary and ternary ionic compounds. A hint: **don't** memorize compound names, memorize naming rules.

## The Periodic Table

Long about 1870, Mendeleev was inspired to put together a tabulation of the known elements in such a way as to describe in a periodic way their physical and chemical properties. The modern version of Mendeleev's Periodic Table is an attractive addition to most science lecture halls—it is worth studying during drier moments of lectures if for no other reason than to prepare for chemistry tests.

### Families of Elements

Note that the table of elements consists of **columns (groups or families)** of elements in which similar properties are observed.

Group IA	alkali metal	Li, Na, K, Rb, Cs
Group IIA	alkaline earth metals	Be, Mg, Ca, Sr, Ba
Group VIIA	halogens	F, Cl, Br, I
Group 0	noble or rare gases	He, Ne, Ar, Kr, Xe

### Periods: Chemical Trends along Rows of Elements

You can also look at the table as a collection of horizontal **rows called periods**:

Row 2                      Li, Be, B, C, N, O, F, Ne

As we will learn, important trends in chemical and physical properties are observed as we move across a row. For example, .....

## Trend 1 :Metals and Non-metals

We will look at numerous trends within the periodic table during the semester. Our first trend will be with respect to metallic character in elements. Note in Table 1 the qualities of metals.

**Table 1: Some Physical Properties of Metals and Nonmetals**

<b>Metals</b>	<b>Nonmetals</b>
1. High electrical conductivity	1. Poor electrical conductivity
2. High thermal conductivity	2. Good heat insulators
3. Metallic gray or silver luster	3. No metallic luster
4. Almost all are solids	4. Solids, liquids, or gases
5. Malleable	5. Brittle in solid state
6. Ductile (can be drawn into wires)	6. Non-ductile

### Your first periodic table trend:

**Metallic:** Character increases from right to left and top to bottom. Thus, Cs is the most active metal. Note that most elements are metals.

**Non-metallic:** Character increases from left to right and bottom to top. The most active non-metal is F.

**Metalloids:** The switch hitters of the world, along the interface between metals and nonmetals. These include elements like B, Si, As, Ge. You may notice that the world of semiconductors revolves around the metalloids.

## Aqueous Chemistry

If there is one molecule worth giving too much attention, it is water. Water makes up 75 % of the earth's surface, 70% of the human body and 100% of a Diet Coke. It is not surprising then that so much effort is devoted to understanding the chemistry of water—almost every reaction we will examine between now and the end of the course will occur in water.

## Electrolytes: Ions Moving Through Water

The chemistry of water is primarily the chemistry of dissolved ions. Water is a polar solvent with the ability to promote the dissociation of ions. One property of these

ions conduct electricity in water.

### Strong Electrolytes

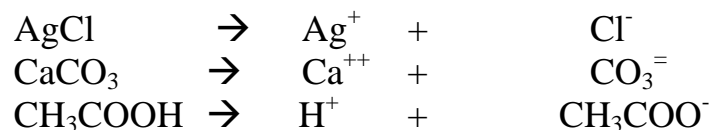
Compounds that **dissociate completely** in water to form ions are called **strong electrolytes**. Examples of these include alkali metal salts and strong acids and bases:



We assume that essentially **100%** of the salts or acids of strong electrolytes **dissociate to form ions**.

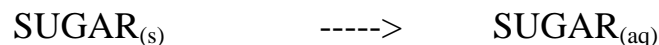
### Weak electrolytes

Some compounds dissociate only slightly, preferring to remain **primarily in an undissociated form**. These molecules are referred to as weak electrolytes and include a wide range of **sparingly soluble salts** as well as weak acids and bases.



## Non-electrolytes

Of course, there are molecules that are too tough to have their bonds ripped apart in water. These are molecules like sugar which possess covalent rather than ionic bonds. Water may be capable of solubilizing these molecules (dissolving them), but **no ions are formed** and electricity is not conducted in solution.



Note that ions are not formed, but the sugar is dissolved in the water (aq means aqueous.)

## Acid Base Chemistry

What is the big deal with acids and bases? Again, it is because we are working with water, which, as we will see, has its own chemistry that produces ions like  $\text{H}^+$  and  $\text{OH}^-$ , and which as everyone knows, are what Arrhenius called acids and bases. Consequently, we will spend a lot of time looking at what happens with these two ions, the proton or **hydronium ion,  $\text{H}^+$** , and the **hydroxide ion,  $\text{OH}^-$** . In CH 302 we will develop many equations that attempt to determine their concentration in solution, their pH. But again, understand the context, we are working in water, and we are examining electrolytes. In other environments,  $\text{H}^+$  and  $\text{OH}^-$  would have no significance. For example, how often do you hear people talk about the pH of gasoline?

### Bronstead/Lowry Definition Acids and Bases:

**acid:** capable of donating a proton

**base:** capable of accepting a proton

The general equation for the dissociation of an acid is simply:



For example:



or



**TABLE D.1** Common Anions and Their Parent Acids

Anion	Parent acid	Anion	Parent acid
fluoride ion, $F^-$	hydrofluoric acid,* HF (hydrogen fluoride)	nitrite ion, $NO_2^-$	nitrous acid, $HNO_2$
chloride ion, $Cl^-$	hydrochloric acid,* HCl (hydrogen chloride)	nitrate ion, $NO_3^-$	nitric acid, $HNO_3$
bromide ion, $Br^-$	hydrobromic acid,* HBr (hydrogen bromide)	phosphate ion, $PO_4^{3-}$	phosphoric acid, $H_3PO_4$
iodide ion, $I^-$	hydroiodic acid,* HI (hydrogen iodide)	hydrogen phosphate ion, $HPO_4^{2-}$	
oxide ion, $O^{2-}$	water, $H_2O$	dihydrogen phosphate ion, $H_2PO_4^-$	
hydroxide ion, $OH^-$		sulfite ion, $SO_3^{2-}$	sulfurous acid, $H_2SO_3$
sulfide ion, $S^{2-}$	hydrosulfuric acid,* $H_2S$ (hydrogen sulfide)	hydrogen sulfite ion, $HSO_3^-$	
hydrogen sulfide ion, $HS^-$		sulfate ion, $SO_4^{2-}$	sulfuric acid, $H_2SO_4$
cyanide ion, $CN^-$	hydrocyanic acid,* HCN (hydrogen cyanide)	hydrogen sulfate ion, $HSO_4^-$	
acetate ion, $CH_3CO_2^-$	acetic acid, $CH_3COOH$	hypochlorite ion, $ClO^-$	hypochlorous acid, $HClO$
carbonate ion, $CO_3^{2-}$	carbonic acid, $H_2CO_3$	chlorite ion, $ClO_2^-$	chlorous acid, $HClO_2$
hydrogen carbonate (bicarbonate) ion, $HCO_3^-$		chlorate ion, $ClO_3^-$	chloric acid, $HClO_3$
		perchlorate ion, $ClO_4^-$	perchloric acid, $HClO_4$

\*The name of the aqueous solution of the compound. The name of the compound itself is in parentheses.

## Distinguishing Strong and Weak Acids and Bases

Now just as there are strong electrolytes which dissociate completely and weak electrolytes which only partially dissociate, there are the equivalent strong and weak acids and bases.

- Strong acids are strong electrolytes
- Weak acids are weak electrolytes

### The definition of a strong or weak acid:

When we start to do a lot of problems involving acid/base chemistry, it is important to be able to distinguish which compounds are strong acids and bases and which are weak acids and bases. Remember that something is defined as an acid or base in water

by an ability to donate a proton or hydroxide, respectively. Whether a compound is a strong or weak acid or base has to do with the extent to which it dissociates.

**Strong acids and bases:** Compounds that are assumed to dissociate completely upon addition to water.

	$\text{HA} \rightarrow$	$\text{H}^+$	+	$\text{A}^-$
initial	100 %	0%		0%
equilibrium	0%	100 %		100%

Common Strong Acids		Anions of These Strong Acids	
Formula	Name	Formula	Name
HCl	hydrochloric acid	$\text{Cl}^-$	chloride ion
HBr	hydrobromic acid	$\text{Br}^-$	bromide ion
HI	hydroiodic acid	$\text{I}^-$	iodide ion
$\text{HNO}_3$	nitric acid	$\text{NO}_3^-$	nitrate ion
$\text{HClO}_4$	perchloric acid	$\text{ClO}_4^-$	perchlorate ion
$\text{HClO}_3$	chloric acid	$\text{ClO}_3^-$	chlorate ion
$\text{H}_2\text{SO}_4$	sulfuric acid	$\text{HSO}_4^-$	hydrogen sulfate ion

How do we know which compounds are strong acids or bases? **We memorize the table.**

**Weak acids.** Of course if an acid or base doesn't dissociate completely, it behaves as a weak electrolyte. Most of the acid stays in the molecular form. Only a small amount dissociates to form a proton and the anion.

	$\text{CH}_3\text{COOH} \rightleftharpoons$	$\text{H}^+$	+	$\text{CH}_3\text{COO}^-$
initial	100 %	0%		0%
equilibrium	99%	1 %		1%

How do we know which acids are weak? Simple. It is any acid that isn't one of the seven strong acids. Several examples are shown in the table above.

**Bases:** We can describe an analogous collection of strong and weak bases that produce  $\text{OH}^-$  in aqueous solution. The strong bases include the Group IA and IIA metal **hydroxides**. For example:



### Table of Strong Bases

LiOH	lithium hydroxide	CsOH	cesium hydroxide
NaOH	sodium hydroxide	Ca(OH) <sub>2</sub>	calcium hydroxide
KOH	potassium hydroxide	Sr(OH) <sub>2</sub>	strontium hydroxide
RbOH	rubidium hydroxide	Ba(OH) <sub>2</sub>	barium hydroxide

**Weak bases.** If a base is a weak electrolyte, it does not dissociate significantly in solution. The most famous weak base is ammonia,  $\text{NH}_3$



## Solubility

Throw a compound into water. Does it dissolve or does it sink to the bottom of the flask? Whether this happens or not determines the **SOLUBILITY** of a compound. A soluble compound **dissolves**. An **insoluble** compound usually sinks to the bottom of the flask as a **PRECIPITATE**, although it may form a **SUSPENSION** in solution. For the most part, we are interested in what ionic compounds do in water. How do we know what these ionic materials do in solution? We learn rules about solubility—see the next page.



## Solubility of Common Ionic Compounds in Water

There are two ways to learn the solubility rules.

- Memorize the individual solubility rules for each kind of cation and anion. The two tables below list the ions that are generally soluble, and the ions that are generally insoluble.
- If the thought of memorizing a lengthy table makes you sad, see if you can spot some generalizations that make it easy to slide by with some simple rules that account for most of the ion solubilities.

**TABLE I.1** Solubility Rules for Inorganic Compounds

Soluble compounds	Insoluble compounds
compounds of Group 1 elements	
ammonium ( $\text{NH}_4^+$ ) compounds	carbonates ( $\text{CO}_3^{2-}$ ), chromates ( $\text{CrO}_4^{2-}$ ), oxalates ( $\text{C}_2\text{O}_4^{2-}$ ), and phosphates ( $\text{PO}_4^{3-}$ ), <b>except</b> those of the Group 1 elements and $\text{NH}_4^+$
chlorides ( $\text{Cl}^-$ ), bromides ( $\text{Br}^-$ ), and iodides ( $\text{I}^-$ ), <b>except</b> those of $\text{Ag}^+$ , $\text{Hg}_2^{2+}$ , and $\text{Pb}^{2+}$ *	sulfides ( $\text{S}^{2-}$ ), <b>except</b> those of the Group 1 and 2 elements and $\text{NH}_4^+$
nitrates ( $\text{NO}_3^-$ ), acetates ( $\text{CH}_3\text{CO}_2^-$ ), chlorates ( $\text{ClO}_3^-$ ), and perchlorates ( $\text{ClO}_4^-$ )	hydroxides ( $\text{OH}^-$ ) and oxides ( $\text{O}^{2-}$ ), <b>except</b> those of the Group 1 and 2 elements <sup>‡</sup>
sulfates ( $\text{SO}_4^{2-}$ ), <b>except</b> those of $\text{Ca}^{2+}$ , $\text{Sr}^{2+}$ , $\text{Ba}^{2+}$ , $\text{Pb}^{2+}$ , $\text{Hg}_2^{2+}$ , and $\text{Ag}^+$ <sup>†</sup>	

\*  $\text{PbCl}_2$  is slightly soluble.

<sup>†</sup>  $\text{Ag}_2\text{SO}_4$  is slightly soluble.

<sup>‡</sup>  $\text{Ca}(\text{OH})_2$  and  $\text{Sr}(\text{OH})_2$  are sparingly (slightly) soluble;  $\text{Mg}(\text{OH})_2$  is only very slightly soluble.

## There has to be an easier way!!

Some **cation** solubility rules:

- Rule 1: All of the alkali metals ( $K^+$ ,  $Na^+$ . etc.) ions plus  $NH_4^+$  are always soluble.
- Rule 2. Heavy cations (down the periodic table) like  $Ba^{++}$  and  $Ag^+$  and  $Pb^{++}$  are insoluble.

Some **anion** solubility rules:

- Rule 3: Anions that are always insoluble except for Rule 1 are mostly the multiply charged species ( $S^{=}$ ,  $CO_3^{=}$ ,  $PO_4^{-3}$ ,  $O^{=}$ , etc.)
- Rule 4. The anions that are primarily soluble are the conjugate bases of strong acids like  $Cl^-$  from  $HCl$ ,  $Br^-$  from  $HBr$ ,  $NO_3^-$  from  $HNO_3$ ,  $ClO_4^-$  from  $HClO_4$  etc.

These rules aren't perfect but they are good enough to get an A in CH301 without doing a lot of memorizing.

## The Best Solubility Rule of All:

But the biggest trend of all, the rule of thumb if you only want to remember one thing is this:

**the singly charged ions are usually soluble**

and

**the multiply charged ions are usually insoluble.**

The reason for this is something we will learn in CH 302.

Solubility Example:

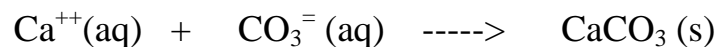
We can use this information to decide the types of reactions that occur when we place compounds in solution.

Consider the reaction of  $CaNO_3$  and  $KCO_3$ .

Note from above that nitrates and alkali metals are soluble. So when these compounds are added to solution, four ions are formed:



Note from above that carbonates are generally insoluble, so it is expected that the following precipitation reaction occurs:



### Oxidation Numbers

Reactions in which substances undergo changes in oxidation number are called oxidation-reduction reactions or redox reactions. Redox reactions involve the transfer of electrons. Some definitions:

- **oxidation:** an algebraic increase in oxidation number in which electrons are lost from a compound.
- **reduction:** an algebraic decrease in oxidation number in which electrons are gained from a compound.
- **oxidizing agent:** substance that gains electrons and oxidizes other substances by being reduced.
- **reducing agent:** substance that loses electrons and reduces other substances by being oxidized.

### Oxidation Rules:

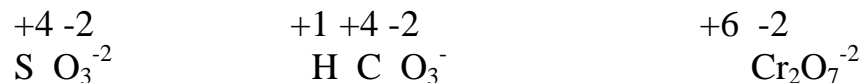
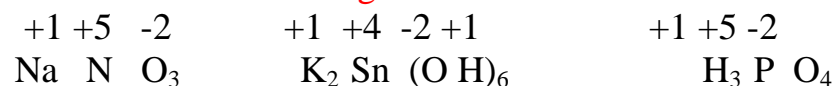
Assignment of the oxidation number for a compound follows certain rules:

1. The oxidation of a free element is zero. **Na is 0**
2. The oxidation number of an element in a monatomic ion is the charge on the ion. **Na<sup>+</sup> is +1**
3. The charge of a monatomic ion is usually determined by its group (column on periodic table). See table below.
4. In the formula for any compound, the sum of the oxidation numbers of all elements in the compound is zero. **NaCl is 0**
5. In a polyatomic ion, the sum of the oxidation numbers of the constituent elements is equal to the charge in the ion.  
**NO<sub>3</sub><sup>-</sup> is -1**

### Some typical oxidation numbers for each:

element	oxidation number
H	+1 (but sometimes -1)
O	-2 (but sometimes -1 or -1/2)
group IA	+1 always
group IIA	+2
group IIIA	+3
group VIIA in reaction with metals	-1
group VIA in reaction with metals	-2
group VA in reaction with metals	-3

### Examples of Oxidation Number Assignment:



### Displacement Reactions.

The world is filled with two kinds of metals, those that dissolve in your pants and those that don't.

Coinage metals, like gold, silver, copper, don't dissolve in your pants, or when you bury them in the ground. This is why they are so valuable.

But valuable also means boring. If you want real fun, you hang out with active metals like potassium and magnesium and lithium, active metals that explode at the slightest provocation (like placing them in water.)

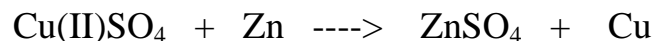
So what does happen when you throw a metal in water?



- The **metal** is **oxidized** (becomes a cation)
- The **water** is **reduced** making **hydrogen** (this is why Na explodes in water)
- The **water** turns basic (you can see this with an indicator turning color)

Wow, what a famous reaction!!

There is a special kind of reaction in which one element displaces another element from a compound. For example:



In general, active metals displace less active metals and hydrogen. There are three categories to consider.

- Reactions that displace hydrogen from non-oxidizing acids.
- Reactions involving metals that displace hydrogen from steam.
- Reaction of metals that displace hydrogen from cold water.

These results are tabulated below:

<b>Type I Elements</b>	<b>Common Reduced Form</b>	<b>Common Oxidized Forms</b>
Li	Li	Li <sup>+</sup>
K Displace hydrogen	K	K <sup>+</sup>
Ca from cold water	Ca	Ca <sup>2+</sup>
Na	Na	Na <sup>+</sup>

Note these elements are the ones that explode in water.

### **Type II Elements**

Mg	Mg	Mg <sup>2+</sup>
Al	Al	Al <sup>3+</sup>
Mn Displace hydrogen	Mn	Mn <sup>2+</sup>
Zn from steam	Zn	Zn <sup>2+</sup>
Cr	Cr	Cr <sup>3+</sup> , Cr <sup>6+</sup>
Fe	Fe	Fe <sup>2+</sup> , Fe <sup>3+</sup>

### Type III Elements

Cd		Cd	$\text{Cd}^{2+}$
Co	Displace hydrogen	Co	$\text{Co}^{2+}$
Ni	from nonoxidizing acids	Ni	$\text{Ni}^{2+}$
Sn		Sn	$\text{Sn}^{2+}$ , $\text{Sn}^{4+}$
Pb		Pb	$\text{Pb}^{2+}$ , $\text{Pb}^{4+}$

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### Type IV Elements

Cu		Cu	$\text{Cu}^+$ , $\text{Cu}^{2+}$
Hg	don't react	Hg	$\text{Hg}_2^{2+}$ , $\text{Hg}^{2+}$
Ag		Ag	$\text{Ag}^+$
Pt		Pt	$\text{Pt}^{2+}$ , $\text{Pt}^{4+}$

In reading the **metal activity table**, the metals at the top of the chart displace salts of the less active metals below them. An easy way to think of this, a reaction occurs if the more stable metal (the one lower on the chart) is a reaction product.

Again, what are some of the general rules you might use rather than have to memorize everything.

The farther to the right, the more reactive the metal.

- **Type I:** The alkali metals (column 1) explode in cold water
- **Type II:** The alkali earths (column 2) dissolve in hot water
- **Type III:** The transition metals dissolve in acid
- **Type IV:** The coinage metals don't do anything.

Not 100% accurate, but good enough for an A (alright, a low A) in CH301.

## Foreign Language Class: Chemistry 506

All of you must have taken a foreign language at some time. This means you've had practice at memorizing large collections of random letters that make up the names we assign to different types of chemical compounds. We are specifically interested in procedures for naming BINARY compounds, which are formed by combining two elements, and TERNARY compounds, formed by putting together three elements. To assist us in our endeavor, it will be of great assistance to memorize the now expanded list of common cations and anions shown in the Table below. Try putting these on flash cards as well.

### Naming Binary Compounds

Let's look at rules for BINARY IONIC COMPOUNDS. In putting together two elements as ions, the following rules apply:

**Rule 1: The element that is more metallic (more electronegative) is listed first. Example:**

Sodium is always listed first because it is always +1

**Rule 2: If the element has more than one common oxidation state, that oxidation number is included in parenthesis after the element**

Copper is either

Copper (I) if it is +1 oxidation state

or

Copper (II) if it is +2 oxidation state

**Rule 3. The element with less metallic character is listed second. This second element is named by adding the suffix "ide" to the STEM of the element.**

Naming the nonmetals:

element	symbol	stem	name
boron	B	bor	boride
carbon	C	carb	carbide
silicon	Si	silic	silicide
nitrogen	N	nitr	nitride
phosphorous	P	phosph	phosphide
arsenic	As	arsen	arsenide
antimony	Sb	antimon	antimonide
oxygen	O	ox	oxide

sulfur	S	sul	sulfide
selenium	Se	selen	selenide
hydrogen	H	hydr	hydride
fluorine	F	fluor	fluoride
chlorine	Cl	chlor	chloride
bromine	Br	brom	bromide
iodine	I	iod	iodide

### Examples of Binary Compound Names

Compound	Name
LiBr	lithium bromide
MgCl <sub>2</sub>	magnesium chloride
Li <sub>2</sub> S	lithium sulfide
Al <sub>2</sub> O <sub>3</sub>	aluminum oxide
Na <sub>3</sub> P	sodium phosphide
Mg <sub>3</sub> N <sub>2</sub>	magnesium nitride

### Multiple Oxidation Numbers

If only life were this simple. Recall that not all elements have a single common oxidation number in ionic form. The idea of using a Roman numeral after the name was proposed to name these compounds. However, there is an older way of naming these ambiguous compounds using -ous and -ic to terminate the metal. The -ic stood for the higher oxidation state, and -ous was attached to the metal with the lower oxidation number.

copper (I) is also called cuprous

copper (II) is also called cupric

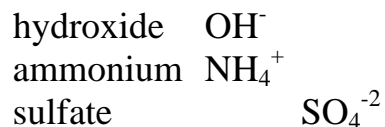


## Examples

compound	older name	modern name
FeBr <sub>2</sub>	ferrous bromide	iron (II) bromide
FeBr <sub>3</sub>	ferric bromide	iron (III) bromide
SnO	stannic oxide	tin (II) oxide
SnO <sub>2</sub>	stannous oxide	tin (I) oxide
TiCl <sub>2</sub>		titanium (II) chloride
TiCl <sub>3</sub>		titanium (III) chloride
TiCl <sub>4</sub>		titanium (IV) chloride

## Pseudo Binary Compounds Names (Polyatomic Ions)

But what about all the compounds containing polyatomic anions and cations? Ions like



Well these ions behave as simple monatomic anions and cations, so we include them in our naming of binary ionic species using the same rules as above, except that we have to memorize all of the polyatomic ion exceptions. There are only 50 or 100 listed below for your memorization pleasure:

### Common Cations

Li <sup>+</sup>	1+
Na <sup>+</sup>	1+
K <sup>+</sup>	1+
NH <sub>4</sub> <sup>+</sup>	1+
Ag <sup>+</sup>	1+
Co <sup>2+</sup>	2+
Mg <sup>2+</sup>	2+
Ca <sup>2+</sup>	2+
Ba <sup>2+</sup>	2+
Cd <sup>2+</sup>	2+
Zn <sup>2+</sup>	2+
Cu <sup>2+</sup>	2+
Hg <sub>2</sub> <sup>2+</sup>	2+
Hg <sup>2+</sup>	2+
Mn <sup>2+</sup>	2+
Ni <sup>2+</sup>	2+
Pb <sup>2+</sup>	2+
Sn <sup>2+</sup>	2+
Fe <sup>2+</sup>	2+
Fe <sup>3+</sup>	3+
Al <sup>3+</sup>	3+
Cr <sup>3+</sup>	3+

### Common Anions

lithium ion	F <sup>-</sup>	1-	fluoride ion
sodium ion	Cl <sup>-</sup>	1-	chloride ion
potassium ion	Br <sup>-</sup>	1-	bromide ion
ammonium ion	I <sup>-</sup>	1-	iodide ion
silver ion	OH <sup>-</sup>	1-	hydroxide ion
cobalt (II) ion	CN <sup>-</sup>	1-	cyanide ion
magnesium ion	ClO <sup>-</sup>	1-	hypochlorite
calcium ion	ClO <sub>2</sub> <sup>-</sup>	1-	chlorite ion
barium ion	ClO <sub>3</sub> <sup>-</sup>	1-	chlorate ion
cadmium ion	ClO <sub>4</sub> <sup>-</sup>	1-	perchlorate
zinc ion	CH <sub>3</sub> COO <sup>-</sup>	1-	acetate ion
copper (II) ion	MnO <sub>4</sub> <sup>-</sup>	1-	permanganate
mercury (I) ion	NO <sub>2</sub> <sup>-</sup>	1-	nitrite ion
mercury (II) ion	NO <sub>3</sub> <sup>-</sup>	1-	nitrate ion
manganese (II) ion	SCN <sup>-</sup>	1-	thiocyanate
nickel (II) ion	O <sup>2-</sup>	2-	oxide ion
lead(II) ion	S <sup>2-</sup>	2-	sulfide ion
tin (II) ion	HSO <sub>3</sub> <sup>-</sup>	1-	bisulfite ion
iron (II) ion	SO <sub>3</sub> <sup>2-</sup>	2-	sulfite ion
iron (III) ion	HSO <sub>4</sub> <sup>-</sup>	1-	bisulfate ion
aluminum ion	SO <sub>4</sub> <sup>2-</sup>	2-	sulfate ion
chromium (III) ion	PO <sub>4</sub> <sup>3-</sup>	3-	phosphate ion
	HCO <sub>3</sub> <sup>-</sup>	1-	bicarbonate
	CO <sub>3</sub> <sup>2-</sup>	2-	carbonate ion
	CrO <sub>4</sub> <sup>2-</sup>	2-	chromate ion
	Cr <sub>2</sub> O <sub>7</sub> <sup>2-</sup>	2-	dichromate ion

### Examples of compounds made with pseudobinary ion:

compound	name
$\text{Na}_2\text{SO}_4$	sodium sulfate
$\text{Cu}(\text{NO}_3)_2$	copper (II) nitrate
$\text{LiCN}$	lithium cyanide
$\text{NH}_4\text{Cl}$	ammonium chloride

So what if it isn't an ion:

### Naming Binary Molecules (Covalent Bonds):

Often two non-metals are put together. In these cases neither has a great desire to rip away or give up electrons, so ionic species aren't formed. Moreover, these types of compounds may exist in a variety of oxidation states. Rather than deal with oxidation state, the number of each atom is listed with a prefix noting the number of atoms of each element.

# atoms	2	3	4	5	6	7	8
prefix	di-	tri-	tetra-	penta-	hexa-	hepta-	octa-

### Examples of molecules:

compound	name
$\text{CO}_2$	carbon dioxide
$\text{SO}_3$	sulfur trioxide
$\text{OF}_2$	oxygen difluoride
$\text{P}_4\text{O}_6$	tetraphosphorus hexaoxide
$\text{Cl}_2\text{O}_7$	dichlorine heptaoxide

## Ternary compounds (after two comes three)

Ternary compounds contain combinations of atoms from three different elements. Things can get really complicated with the additional element, so we will stick to just a few simple classes.

First, in working with ionic compounds there were a variety of polyatomic ions that you were asked to memorize. These followed the basic rules of binary ions and are easy to name IF you have them all memorized.

## Ternary Acids and Their Salts

There are a collection of well know ternary acids which consist of hydrogen, oxygen and a nonmetal.

The most famous example is



## Rules for Naming Ternary Acids

There are some basic rules for naming such compounds which depends upon the oxidation state of the central atom, XXX :

	Ternary acid	anion salt	
decreasing oxidation number of X	perXXXic acid	perXXXate	decreasing oxygens on X
	XXXic	XXXate	
	XXXous acid	XXXite	
	hypoXXXous acid	hypoXXXite	

## Examples of Ternary Acids

acid	name	sodium salt	name
$\text{HNO}_2$	nitrous acid	$\text{NaNO}_2$	sodium nitrite
$\text{HNO}_3$	nitric acid	$\text{NaNO}_3$	sodium nitrate
$\text{H}_2\text{SO}_3$	sulfurous acid	$\text{Na}_2\text{SO}_3$	sodium sulfite
$\text{H}_2\text{SO}_4$	sulfuric acid	$\text{Na}_2\text{SO}_4$	sodium sulfate
$\text{HClO}$	hypochloric acid	$\text{NaClO}$	sodium hypochlorite
$\text{HClO}_2$	chlorous acid	$\text{NaClO}_2$	sodium chlorite
$\text{HClO}_3$	chloric acid	$\text{NaClO}_3$	sodium chlorate
$\text{HClO}_4$	perchloric acid	$\text{NaClO}_4$	sodium perchlorate

**I have never done anything so boring in my life  
as type this particular set of notes.**

The end.