# Nomenclature of Simple Inorganic Compounds

(extensively modified by Dr. Arno Papazyan)

## Nomenclature of <u>Ionic</u> Compounds

Ionic compounds are composed of ions. An ion is an atom or molecule with an electrical charge.

- Monatomic ions are formed from single atoms that have gained or lost electrons.
- Polyatomic ions are like molecules (groups of atoms covalently bonded together), except that they have a net overall electrical charge.
- Negative ions are called anions
- Positive ions are called cations

Nonmetals form anions, while metals form cations.

- Ionic charge is shown as a superscript (when the ion is shown alone)
  - Magnitude of the charge is put before the sign: 3+ instead of +3;  2- instead of -2
  - If the magnitude of the charge is 1, only the sign is shown.

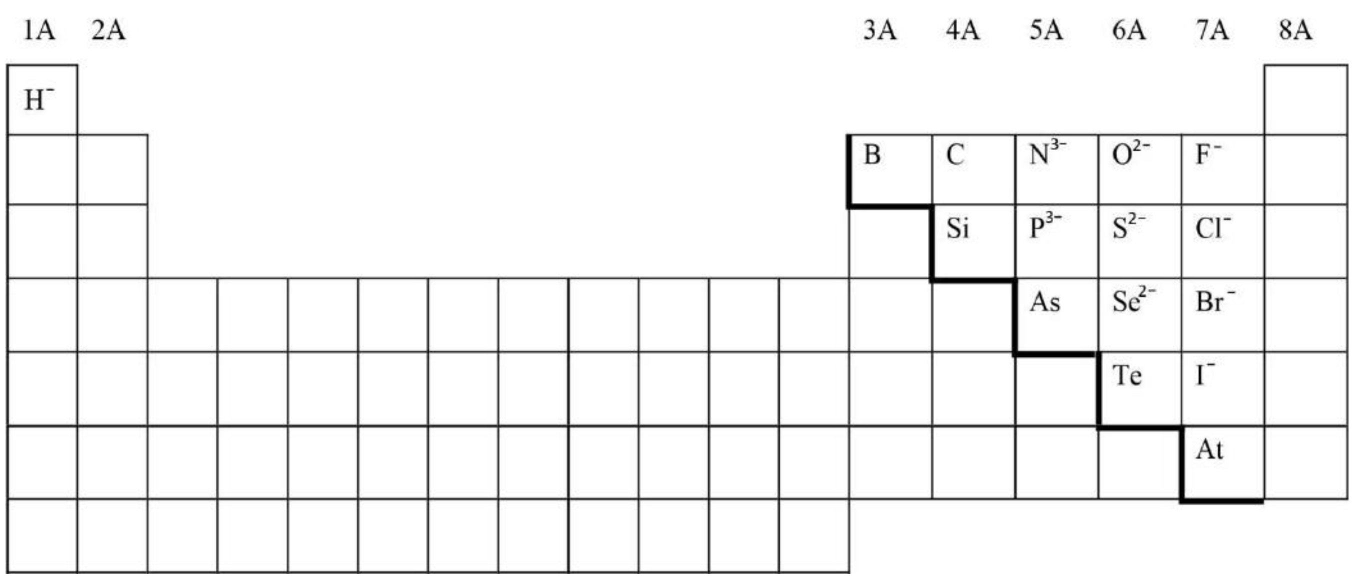
Ions with opposite charges (positive metal cations and negative non-metal anions) experience a strong electrostatic attraction and form an ionic bond, which leads to the formation of the ionic compound.

*Side note: There is some pretense in what we call an "ionic compound" when the supposed charge magnitude on a monatomic (single-atom) ion is greater than 2 (e.g. with* $P^{3-}$ *or* $Ga^{3+}$*). In those cases, the bonding is largely covalent (and the actual charge on the atom isn't nearly as large as its 'official' value), but we still name the compound as if it were truly ionic.*

**If the first element in the formula is a metal, we name the compound according to the nomenclature of ionic compounds.**

**<u>Non-metal Anions</u>**

Non-metals form anions with only one possible negative charge (hydrogen can also form a cation with +1 charge). Periodic Table below shows the charges of non-metal anions:



The magnitude of the anion charge is equal “Main Group Number” minus 8 (except for hydrogen).

The names of these anions are formed by taking the “root” of the element name, and adding an “ide”.

Example 1: Sulfur in Group 6A forms “sulf**<u>ide</u>**” anion with a charge of [6-8] = -2, shown as $S^{2-}$.

Example 2: Chlorine in Group 7A forms “chlor**<u>ide</u>**” anion with a charge of [7-8] = **-**1, shown as $Cl^-$

**<u>Metal Cations</u>**

Most main group metals (except for the heavier metals in Groups 3A-7A) will form cations with only one possible charge. Most transition metals (except for Ag, Zn, and Cd) form cations with more than one possible charge. The following Periodic Table shows the charges for metal cations commonly found in ionic compounds. There is no need to memorize the multiple charges a metal can take on. You can deduce the metal cation charge from the context (formula of the compound and the neutralizing anion).

| 1A | 2A | Transition Elements (B) | | | | | | | | | | 3A | 4A | 5A |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| $H^+$ | | | | | | | | | | | | | | |
| $Li^+$ | $Be^{2+}$ | | | | | | | | | | | | | |
| $Na^+$ | $Mg^{2+}$ | | | | | | | | | | | $Al^{3+}$ | | |
| $K^+$ | $Ca^{2+}$ | | $Ti^{2+}$<br>$Ti^{4+}$ | | $Cr^{2+}$<br>$Cr^{3+}$<br>$Cr^{6+}$ | $Mn^{2+}$<br>$Mn^{3+}$<br>$Mn^{4+}$ | $Fe^{2+}$<br>$Fe^{3+}$ | $Co^{2+}$<br>$Co^{3+}$ | $Ni^{2+}$<br>$Ni^{3+}$ | $Cu^+$<br>$Cu^{2+}$ | $Zn^{2+}$ | $Ga^{3+}$ | Ge | |
| $Rb^+$ | $Sr^{2+}$ | | | | | | | | | $Ag^+$ | $Cd^{2+}$ | $In^+$<br>$In^{3+}$ | $Sn^{2+}$<br>$Sn^{4+}$ | Sb |
| $Cs^+$ | $Ba^{2+}$ | | | | | | | | | $Au^+$<br>$Au^{3+}$ | $Hg_2^{2+}$<br>$Hg^{2+}$ | | $Pb^{2+}$<br>$Pb^{4+}$ | $Bi^{3+}$<br>$Bi^{5+}$ |
| | | | | | | | | | | | | | | |

The magnitude of the positive charge* on the <u>main group</u> metal cations is generally equal to their Group Number.

- The names of metal cations with only one possible charge are the same as the names of the metals themselves, with no further modification.
- For metal cations with more than one possible charge, the ion charge must be indicated in the ion name. In the IUPAC system, the ion charge (or “oxidation state”) is indicated in the name as Roman numerals in parentheses.

Example 3: Magnesium in Group 2A forms cations with a **+2 charge**.

Example 4: $Al^{3+}$ is called the **aluminum** cation.
$Ag^+$ is called the **silver** cation.

Example 5: $Pb^{2+}$ is called **lead(II)**.
$Ti^{4+}$ is called **titanium(IV)**.

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* A “charge” greater than +2 is not an actual charge; it is merely an “oxidation state”, which we will learn about when we study oxidation-reduction reactions. With such high supposed charges, we are only pretending that we have an ion. For naming purposes, we <u>will</u> treat them as if they are actual ions.

**Polyatomic Ions**

Polyatomic ions are just like molecules (groups of atoms covalently bonded together) but carry an overall charge. The table below includes a list of common polyatomic ions (mostly anions).

| | | | |
|---|---|---|---|
| $CO_3^{2-}$ | Carbonate | $CN^-$ | Cyanide |
| $HCO_3^-$ | Bicarbonate (Hydrogen Carbonate) | $OCN^-$ | Cyanate |
| $SO_3^{2-}$ | Sulfite | $SCN^-$ | Thiocyanate |
| $HSO_3^-$ | Bisulfite (Hydrogen Sulfite) | $CrO_4^{2-}$ | Chromate |
| $SO_4^{2-}$ | Sulfate | $Cr_2O_7^{2-}$ | Dichromate |
| $HSO_4^-$ | Bisulfate (Hydrogen Sulfate) | $MnO_4^-$ | Permanganate |
| $NO_2^-$ | Nitrite | $C_2H_3O_2^-$ | Acetate |
| $NO_3^-$ | Nitrate | $C_2O_4^{2-}$ | Oxalate |
| $ClO^-$ | Hypochlorite | $OH^-$ | Hydroxide |
| $ClO_2^-$ | Chlorite | $O_2^{2-}$ | Peroxide |
| $ClO_3^-$ | Chlorate | | |
| $ClO_4^-$ | Perchlorate | | |
| $PO_4^{3-}$ | Phosphate | $NH_4^+$ | Ammonium |
| $S_2O_3^{2-}$ | Thiosulfate | $Hg_2^{2+}$ | Mercury (I) |

You are expected to recognize and remember the names and formulas (and charges) of these ions. But don't rush to memorize them right away. Practice with nomenclature let at least some of them become your knowledge naturally.

- Almost all the polyatomic ions are negatively charged anions (except $NH_4^+$ and $Hg_2^{2+}$).
- The names of polyatomic anions containing oxygen, called "oxyanions", end in either -ate or -ite. For a given element's oxyanions, the -ate ending corresponds to more oxygen than -ite. For example, sulf**ate**, $SO_{\underline{4}}^{2-}$, has one more oxygen than sulf**ite**, $SO_{\underline{3}}^{2-}$.
- Oxyanions are anions formed by a non-oxygen atom with one or more oxygen atoms.
    - The magnitude of the negative charge of an oxyanion is reduced by one with each hydrogen cation (which brings a +1 charge) that is added to it. If the original oxyanion has a charge of -2, the one with a hydrogen will have a charge of -1, halving it (which is why its informal name gets the prefix "bi" meaning "half"). For example $CO_3^{2-}$ is "carbonate", and $HCO_3^-$ is "bicarbonate". Note that, once the $H^+$ is added to an oxyanion, it becomes covalently bonded. The new anion is still a covalent entity like a molecule, except for carrying an overall charge.
    - Oxyanions of halogens (F, Cl, Br, I), can have 1, 2, 3, or 4 oxygens. The middle two, with 2 and 3 oxygens, get the "-ite" and "-ate" endings respectively. The one with even less oxygen than the "-ite" ions (i.e. only one oxygen) gets the additional prefix of "hypo" as well as the "-ite" ending. The one with even more oxygen than the "-ate" ion gets the "per" prefix as well as the "-ate" ending.
    - When an oxygen atom is replaced with a sulfur (S) atom, the new ion get the prefix "thio". It has the same charge as the original anion. For example $SO_4^{2-}$ is "sulfate", whereas $S_2O_3^{2-}$ is "thiosulfate".

## Formulas and Names of Ionic Compounds

The following are the basic rules for writing the formulas and names of ionic compounds:

**Writing the formula**

1. Determine the formulas and charges on the cation and anion involved in the compound.
2. Combine the ions in a ratio that results in the formation of a neutral ionic compound. In other words, the total charge of all the positive cations must equal the total charge of all the negative anions in the compound. The numbers of each ion present in the compound are shown as subscripts.

**Naming the compound, given the cation and the anion**

Ionic compound name is simply the cation name followed by the anion name:

{cation name} {anion name}

**Naming the compound, given the formula**

You must recognize the polyatomic ion in the compound formula, which does not show the ion separately, with its charge. For example, you must recognize that the "$SO_4$" in $MgSO_4$ is the sulfate ion, $SO_4^{2-}$.

Once you recognize the cation and the anion, again the ionic compound name is simply the cation name followed by the anion name:

{cation name} {anion name}

Example 7: Write the formula and name for the compound formed between calcium and fluorine.

Ca (metal) forms a +2 cation $Ca^{2+}$ the **calcium** cation.
F (non-metal) forms a -1 anion $F^-$ the **fluoride** anion.

To obtain a neutral compound, **1** $Ca^{2+}$ is needed for every **2** $F^-$

The formula of the compound is **$CaF_2$**

The name of the compound is **Calcium Fluoride**

Example 8: Write the formula for iron(III) chloride.

First identify the cation and the anion in this compound.

Cation = iron(III) = $Fe^{3+}$
Anion = chloride = $Cl^-$ (non-metal anion)
To obtain a neutral compound, **1** $Fe^{3+}$ is needed for every **3** $Cl^-$

The formula of the compound is **$FeCl_3$**

Example 9: Write the formula for magnesium phosphate.

First identify the cation and anion in this compound.

Cation = magnesium = $Mg^{2+}$
Anion = phosphate = $PO_4^{3-}$

To obtain a neutral compound, **3** $Mg^{2+}$ are needed for every **2** $PO_4^{3-}$
$3 \times (+2) = +6$ neutralizes $2 \times (-3) = -6$

The formula of the compound is **$Mg_3(PO_4)_2$**

Note in the above Example 9 that parentheses are placed around the polyatomic ion portion of compound, to indicate that it must be treated as a distinct, whole unit.

Example 10: Name the ionic compound $Al(NO_3)_3$.

First identify the cation and anion in this compound.

Cation = $Al^{3+}$ = the aluminum cation
Anion = $NO_3^-$ = the nitrate anion

The name of this compound is **Aluminum Nitrate**

Example 11: Name the compound $TiO_2$.

First identify the cation and anion in this compound.

Cation = $Ti^{+4}$ = the titanium(IV) cation

Anion = $O^{2-}$ = the oxide anion

The name of the compound is **Titanium(IV) Oxide**

# Nomenclature of Simple Covalent Compounds

Covalent compounds are compounds formed between non-metals only. Simple binary covalent compounds contain just two different types of non-metal elements. When non-metals combine they can form several different covalent compounds, and there is no easy, immediate way to predict the formula of the covalent compounds that can be formed.

Example 12: Carbon and oxygen combine to form two common covalent compounds $CO_2$ and CO.

Nomenclature of simple covalent compounds is based on the ionic compound nomenclature, with the addition of greek prefixes denoting how many of each element is present in the formula. The prefixes are needed because there is no easy, general, way to deduce the number of atoms from the identities of the elements.

**Formulas and Names of Simple Covalent Compounds**

1. Always write/name the element with more "metallic" character first (even though both are non-metals). Metallic character increases going from right to left, and top to bottom on the Periodic Table.
2. Then write/name the second (less metallic) element, changing the ending of its name to -ide.

   So far, it's as if we are naming an ionic compound. But:
3. Since nonmetals often combine in different proportions to form a number of different compounds, prefixes must be included in the names to indicate the numbers of each kind of atom present. Prefixes for 1-10 atoms are given in the following table.

| Number | **Prefix** | **Number** | **Prefix** |
|---|---|---|---|
| 1 | Mono | 6 | Hexa |
| 2 | Di | 7 | Hepta |
| 3 | Tri | 8 | Octa |
| 4 | Tetra | 9 | Nona |
| 5 | Penta | 10 | Deca |

Example 13: A compound contains 3 atoms of sulfur and 4 atoms of phosphorus. Write its name and formula.

Since **phosphorus is the more metallic element** (left of sulfur), it must be written/named first.

Sulfur, being the less metallic element, is named second with an -ide ending = **sulfide**.

The prefix for **three** is **tri**, and the prefix for **four** is **tetra**.

The name of this compound is **tetraphosphorus trisulfide**

The formula of this compound is $\mathbf{P_4S_3}$

Example 14: Write the name of the compound $N_2O$.

The **two N atoms** require the prefix **di** in the name.

The **one O atom** requires the prefix **mono** in the name. The ending of oxygen must be changed to –ide = **oxide**.

The name of this compound is **dinitrogen monoxide**

There are two important exceptions to the naming rules outlined so far:

- Never use the prefix "mono" for the first element, even if just one atom is present.
- Never use any prefixes at all for simple covalent compounds containing **hydrogen**[†]. Those compounds are named according to other conventions (such as for acids), and when they are not, the prefix is still not necessary because:
  - Hydrogen is a very simple element, always making a single bond. You can predict the formula if you pretend that it is $H^+$ and the other nonmetal formed its usual anion (not true, but it works)

    e.g. water, $H_2O$, is definitely a covalent comppound, but its formula can be predicted by imagining as if two $H^+$ cations combined with an oxide ion, $O^{2-}$. By the way, while nobody calls it that, its systematic name would have been hydrogen oxide (and not dihydrogen monoxide).

Example 15: Write the name of the compound $BCl_3$.

Although just one B atom is present, the prefix **mono is not used** since it is the first element in the formula.

The **three Cl atoms** require the prefix **tri** in the name. The ending of chlorine must be changed to –ide = **chloride**.

The name of this compound is **boron trichloride**

Example 16: HF is hydrogen fluoride, not hydrogen monofluoride or monohydrogen monofluoride.

Traditional names are used for the following binary covalent compounds:

- **$H_2O$ (water)**
- **$NH_3$ (ammonia)**
- **$CH_4$ (methane)**

Please know the formulas and names of those three compounds.

Covalent compounds containing more than two elements have different systems of nomenclature, which we will not learn in this course, except for acids, which are indeed covalent, molecular compounds and usually contain more than two elements.

[†] Prefixing the hydrogen may be necessary for naming an anion (itself not a full compound) formed by combining an oxyanion with a charge 3- or greater with two or more hydrogen cations; e.g $H_2PO_4^-$ is dihydrogen phosphate, distinguished from $HPO_4^{2-}$ (hydrogen phosphate). Still, note that the original oxyanion part is still not prefixed (i.e. not "dihydrogen monophosphate").

# Nomenclature of Acids

Acids can be described as compounds that release hydrogen cations ($H^+$) when dissolved in water.

| | | | | |
|---|---|---|---|---|
| $HCl\ (aq)$ | → | $H^+\ (aq)$ | + | $Cl^-\ (aq)$ |
| hydrochloric acid | | hydrogen ions | | chloride ions |
| $HNO_3\ (aq)$ | → | $H^+\ (aq)$ | + | $NO_3^-\ (aq)$ |
| nitric acid | | hydrogen ions | | nitrate ions |
| $H_2SO_3\ (aq)$ | → | $2\ H^+\ (aq)$ | + | $SO_3^{2-}\ (aq)$ |
| sulfurous acid | | hydrogen ions | | sulfite ions |

Acid molecules are formed by combining enough $H^+$ cations with an anion to make a neutral molecule. Note that acids are covalent, molecular compounds. They only form ions when dissolved in water, splitting into an anion and $H^+$ cations, as shown above. While it is convenient to regard the hydrogen in the acid as acting like a Group 1A metal cation, it does not exist as an ion when it is attached to the "anion" part of the acid (before being released to the aqueous solution).

Although molecular compounds, the formulas of acids are obtained in a similar fashion to ionic compounds:

- Number of acidic hydrogens (i.e. number of $H^+$ cations needed to add to the anion to make a neutral compound) is equal to the magnitude of the anion's charge. $Cl^-$ needs one hydrogen, $SO_4^{2-}$ needs two hydrogens, $PO_4^{3-}$ needs three hydrogens, and so on.

In the inorganic nomenclature used here, "acidic" hydrogens (those that can be released as $H^+$ in aqueous solution) are always written first in the formulas of all acids.

How do we recognize that we have an acid, given a formula?

- If the first element in the formula is hydrogen (H), you can regard it as an acidic compound (except, of course, for $H_2O$). That is the convention used in inorganic chemistry and introductory chemistry.

In organic chemistry (involving the vast majority of the compounds of carbon) the convention for writing formulas is different, but we will generally not use that convention here, and your usage it will be interpreted as obtained from an internet search (as it would not be using the proper convention).

Also, do not use overall molecular formulas that fail to put acidic (and only acidic) H atom(s) at the beginning of the formula. It will also be interpreted as improperly obtained from the internet.

## **The anion produced by the acid determines how the acid is named**.

- The acid name is derived from the root-name of the anion it produces.

Example 17: HCl produces the chloride ion, $Cl^-$ when dissolved in water.
Root-name of chloride is "chlor" (the part remaining when the -ide suffix is removed).
The acid name is derived from "chlor".

- Also, the **ending** of the anion name determines **how** the acid name is derived.

## Acidic compounds that produce anions whose name end with **-ide**

If the anion of the acidic compound has a name ending with **-ide**, the compound is named as an acid **only if it is dissolved in water**, as follows:

**hydro**{anion root-name}**ic acid**

Example 18: Name the acid $HBr(aq)$.

This acid produces the bromide anion, $Br^-$.

The anion name ends with -ide, and the compound is in aqueous solution, so it is named as an acid as follows, based on the anion root-name "brom":

**hydro** + brom + '**-ic**' + **acid** ⇒ **hydrobromic acid**

If the compound is not in aqueous solution, it is named as an ordinary compound of hydrogen and the anion (which is almost always monatomic, except for cyanide, $CN^-$). For example, HCN(*g*) would be named "hydrogen cyanide", but HCN(*aq*) would be named "hydrocyanic acid".

- Remember that we do not use the greek prefixes with the binary compounds of hydrogen (cyanide, $CN^-$, is treated as if it were a monatomic anion). So HCN(g) is "hydogen cyanide" and not "hydrogen monocyanide"!
- Also remember that this convention is specific to the acidic compounds whose anion-names end with **-ide** **only**.

Example 19: Write the formula and name for the acid containing the sulfide anion.

The sulfide, $S^{2-}$, anion needs to combine with 2 $H^+$ cations to produce a neutral acid molecule, $H_2S$.
Since we want the acid name, $H_2S$ needs to be in the form of an aqueous solution:
The formula for the acid is: **$H_2S$ (*aq*)**
The root-name of "sulfide" is actually the full element name "sulfur" (no good justification; simply convention). The acid name is:

**hydro** + sulfur + '**-ic**' + **acid** ⇒ **hydrosulfuric acid**

## Acidic compounds that produce anions whose name end with <u>-ite</u> or <u>-ate</u>

Anions whose names end with **-ite** or **-ate** are polyatomic anions (the only polyatomic anion whose name does not end with -ite or -ate is cyanide), and when they combine with the appropriate number of $H^+$ cations, they produce an acid molecule.

Example 20: Nitric acid, $HNO_3$(*aq*), is obtained by combining one $H^+$ cation with each nitrate anion, $NO_3^-$. And each $HNO_3$ molecule can produce one $H^+$ cation and one $NO_3^-$ anion in aqueous solution.

When the anion name ends with with -ite or -ate, it doesn't matter if the acid is dissolved in aqueous solution or not. We always name the corresponding acid as an <u>acid</u>, even if it's the pure liquid, or a bunch of crystals of the pure acid.

<u>Warning</u>: While the above statement is true, the idea that all acids must be written as aqueous solution, the formula followed by (*aq*), has taken root in introductory and general chemistry instruction. Be prepared to answer accordingly in subsequent courses where you might be expected to always put (*aq*) after acid formulas. Just be aware that such a requirement is not true.

If the polyatomic ion has the ending **-ate**, in the acid the ending is changed to **-ic** + **acid**.
In other words, the corresponding acid name is:

{anion root-name}**ate** ⇒ {anion root-name}**ic acid**

Example 21: Name the acid $HClO_3$ (aq).

$HClO_3$ has one acidic hydrogen, and produces the $ClO_3^-$, chlor<u>ate</u> ion.

To name this acid, the anion ending -ate is switched to –ic, and the term "acid" is added.

$HClO_3$ (*aq*) is thus called <u>chlor**ic acid**</u>

If the polyatomic ion has the ending **-ite**, in the acid the ending is changed to **-ous** + **acid**.

{anion root-name}**ite** ⇒ {anion root-name}**ous acid**

Example 22: Name the acid $HNO_2$ (aq).

$HNO_2$ has one acidic hydrogen, and produces the $NO_2^-$, nitr<u>ite</u> ion.

To name this acid, the anion ending -ite is switched to –ous, and the term "acid" is added.

$HNO_2$(*aq*) is thus called nitr<u>**ous acid**</u>

Example 23: Write the formula for oxalic acid.

Oxal**ic acid** must contain (by reverse logic) the oxal**ate** anion, $C_2O_4^{2-}$, combined with <u>two</u> acidic hydrogens (to neutralize the -2 charge of $C_2O_4^{2-}$).

The formula of oxalic acid is $H_2C_2O_4$

Note that we omitted the (*aq*) here, as solid crystals of oxalic acid are simply called oxalic acid. But as warned earlier, you may be required to put (*aq*) in subsequent courses.

## Nomenclature of Hydrates

A hydrate is typically an ionic compound with a certain number of water molecules (per formula) loosely bound to it (typically not by actual chemical bonds).

The general formula of a hydrate is $MX \bullet nH_2O$ (*s*), where M is the cation in the ionic compound, X is the anion in the ionic compound and $nH_2O$ are the *n* water molecules loosely bound to each formula unit of the ionic compound.

Hydrates are named by writing the name of the ionic compound first, followed by the word "hydrate". To indicate the number of water molecules present per formula unit, prefixes must be used.

Example 24: Name the hydrate $MgSO_4 \cdot 7H_2O$.

$MgSO_4$ is the ionic compound **magnesium sulfate**.

Since there are **seven water molecules** present, the correct prefix to use is **hepta**.

The name of this hydrate is **magnesium sulfate heptahydrate**

Example 25: Write the formula for copper(II) chloride dihydrate.

Copper(II) chloride has the formula **$CuCl_2$**

The prefix **di** indicates that there are **two** water molecules present.

The formula of this hydrate is **$CuCl_2 \cdot 2H_2O$**

The water molecules in a hydrate can be removed with relative ease by heating the hydrate. The ionic compound that remains after heating is called an anhydrous salt.

| $MX \bullet nH_2O$ (*s*) → | $MX$ (*s*) | + | $n\ H_2O$ |
|---|---|---|---|
| Hydrate | Anhydrous salt | | Free Water |

Often the anhydrous salt has a completely different color and texture from the hydrate (in addition to having different physical properties).

Example 26: Copper(II) sulfate pentahydrate is blue and crystalline, whereas anhydrous copper(II) sulfate is white and powdery.