Chapter 3

Chemical Compounds

Chemical Bonds

Strong attraction holding two atoms together

Ionic Bonds

- Due to attraction between oppositely charged **ions**.
- Form an extensive lattice of ions (**ionic crystals**), **not molecules**.

Covalent Bonds

- Formed by sharing electrons.
- The nuclei of the bonded atoms are attracted to the electrons in between.
- **Molecules** are formed by covalent bonds

  But covalent bonds can lead to other structures too, like covalent crystals

Groups or "Families" of elements and their ions

Group or Family	Charge of ion
Alkali Metals (1A)	1+
Alkaline Earth Metals (2A)	2+
Aluminum *(not a group)* (3A)	3+
Group 5 nonmetals (5A)	3–
Chalcogens (6A)	2–
Halogens (7A)	1–
Noble Gases (8A)	0

Main Group #

Main Group #

Main Group # of noble gases

$\boxed{5-8=-3}$

$\boxed{6-8=-2}$

$\boxed{7-8=-1}$

(they don't ionize)

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By the way:

These charges are taken on **only when they are stabilized by nearby charges of the opposite sign.**

-- such as in an ionic compound

-- or when that ionic compound is dissolved in a solvent with "polar" molecules, like water

Some parts are electron-rich (–), some parts are electron poor (+)

If you ignore this point, you might think those elements always exist as ions!

A common misconception!

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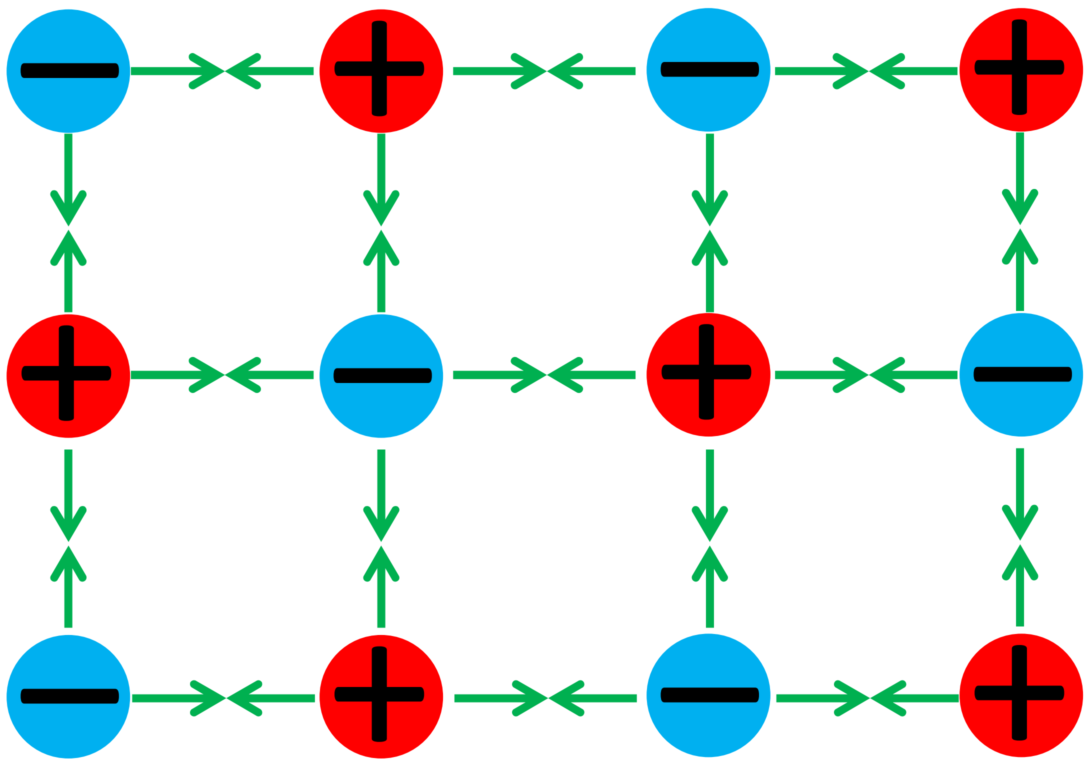
Ionic bonding

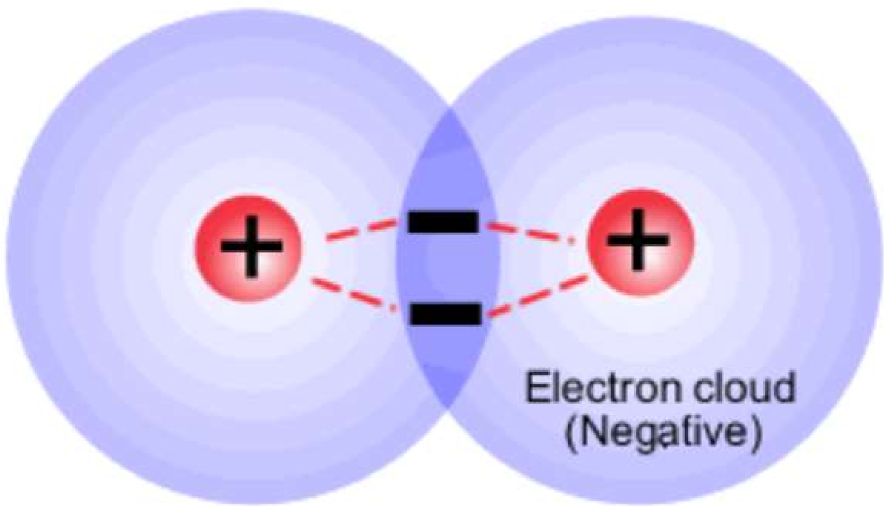
Due to the electrostatic attraction between positive and negative ions

The ions form an extended lattice instead of a molecule



Covalent bonding

Two electrons (normally one from each atom) spend most of their time between the two nuclei, attracted and stabilized by both, instead of just one.



The two nuclei in turn are attracted to the accumulated electrons between them.

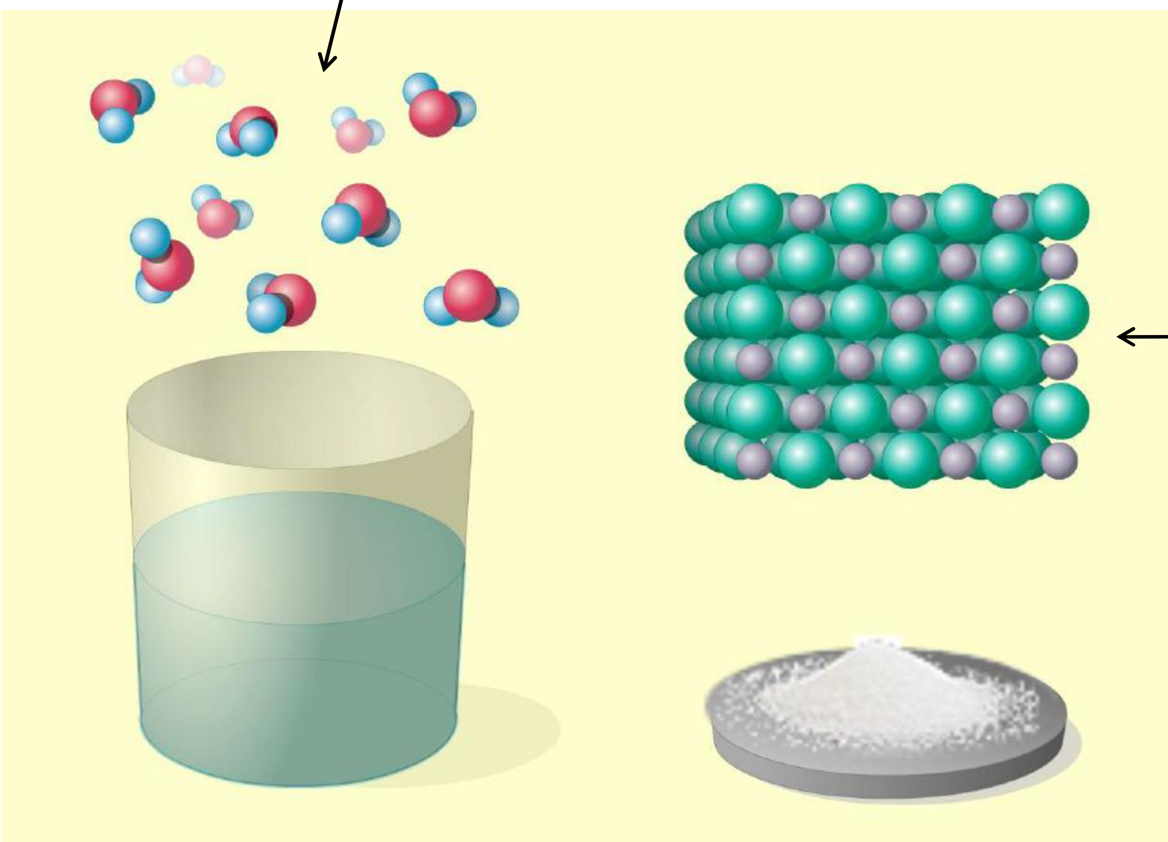
This attraction holds the two atoms together.

Picture from http://myibchemistry.blogspot.com/2013/11/topic-42-covalent-bonding.html

Molecular vs Ionic Compounds

Atoms covalently bonded in a molecule always travel together

It’s clear which atoms belong together



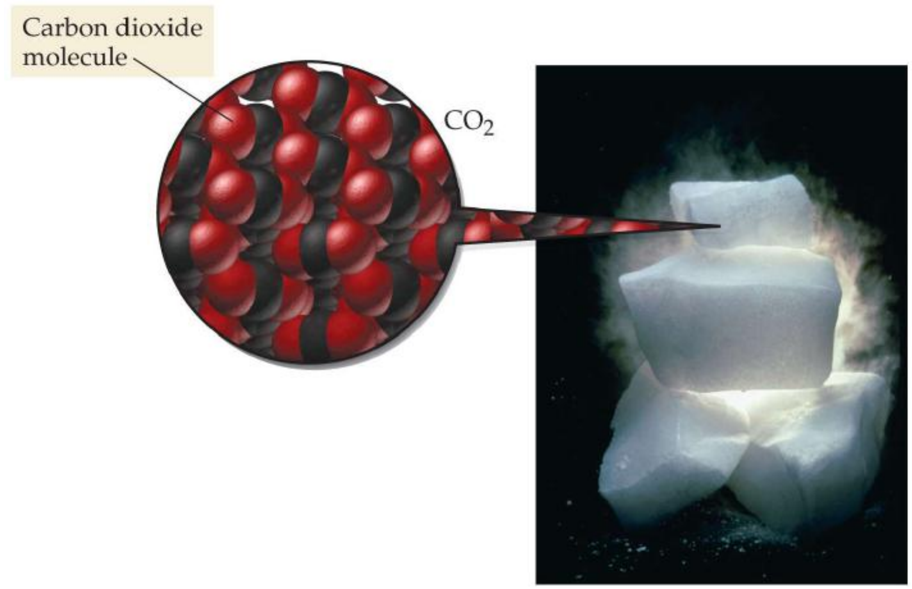
Ions in an ionic compound form a lattice but don’t travel together.

When dissolved or melted, they go their separate ways

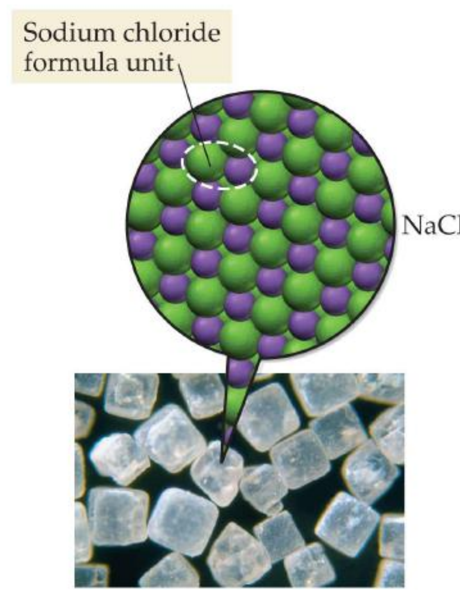
No specific cations and anions “belong” together

- **Covalent compounds** are formed from two or more nonmetals (usually), and **only** have covalent bonds.
  - Often “Molecular” is used to mean “Covalent”
  - Molecular compounds are covalent
  - But not all covalent compounds are molecular
  - Some covalent compounds exist as covalent crystals
- **Ionic compounds** have at least one ionic bond per formula
  - one or more cations with one or more anions

The basic units that compose dry ice, a molecular compound, are CO_2 **molecules.**



The basic units that compose table salt, an ionic compound, are NaCl **formula units.**



Ionic Compounds

- We will focus on **binary** ionic compounds
- "Binary" indicates "two elements"
- Cation is formed by a metal
- Anion is formed by a nonmetal
- We will treat metal-nonmetal binary compounds as **ionic** (even though some are covalent)
- The basic unit of ionic compounds is the **formula unit.**
- Formula unit is the "pretend molecule"
- Formula unit has the smallest overall-neutral group of cations and anions of the ionic compound

- Ionic compounds always contain positive and negative ions.
- In the chemical formula of an ionic compound, the sum of the charges of the positive ions (cations) must always equal the sum of the charges of the negative ions (anions).
    - Compounds are neutral

Writing Formulas for Ionic Compounds

1. Write the symbol for the metal and its charge followed by the symbol of the nonmetal and its charge.

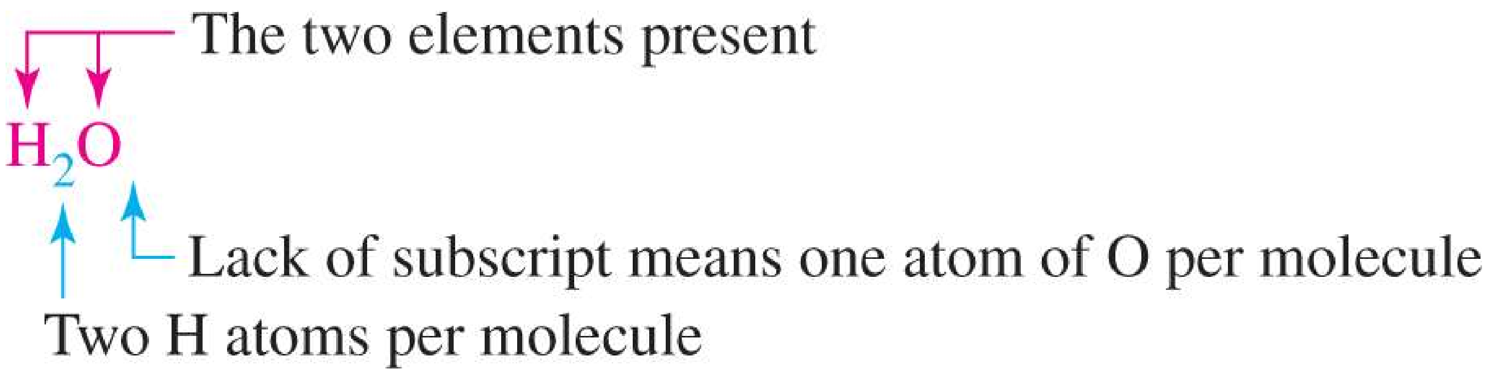
2. Make the magnitude of the charge on each ion (without the sign) become the subscript for the other ion.

3. If possible, reduce the subscripts to give a ratio with the smallest whole numbers.
   – Don't write a subscript of 1

4. Make sure charges add up to zero.

Ca^{2+} O^{2-}	Al^{3+} S^{2-}
Ca^{2+}_{2} O^{2-}_{2}	Al^{3+}_{2} S^{2-}_{3}
Ca O	Al_2S_3
+2 + (-2) = 0	2(+3) + 3(-2) = 0

Formulas for Molecular Compounds



Unlike for ionic compounds, the chemical formula for a molecular compound indicates literally how many atoms of an element is in one molecule of the compound.

The subscripts for each element symbol indicates the actual number of atoms of that element in one molecule (rather than the ratio of the numbers of the ions in the crystal).

Empirical formula: CH_2O

Shows the simplest ratio of the number of atoms of each element in the molecular compound

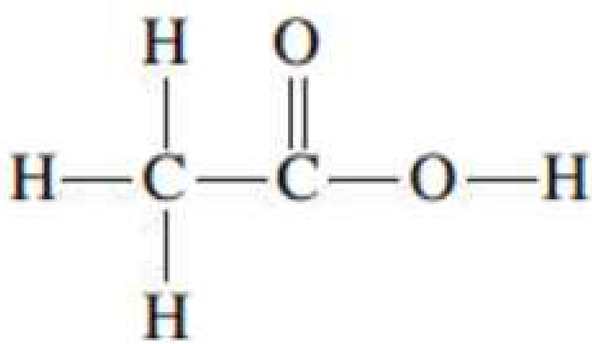
Molecular formula: $C_2H_4O_2$

Shows the actual number of atoms for each element in the molecule

Does not show who is connected to whom

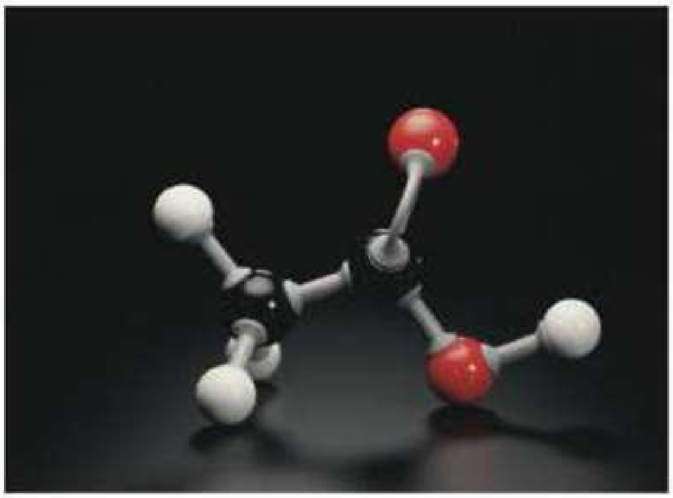
Does not show anything about the shape of the molecule

Structural formula:



Shows who is connected to whom

Does not show anything about the shape of the molecule



Molecular model
("ball and stick")

Molecular model
("space filling")

Molecular models show the three-dimensional shape of the molecule

The mass of one molecule (**molecular mass**) is the sum of the masses of all the atoms in the molecule.

The mass of one formula unit (**formula mass**) is the sum of the masses of all the atoms in the molecule.

Multiply the atomic mass of each atom by its count in the formula

$$\text{Molecular Mass of } H_2O = (2 \times \overset{\text{H}}{1.008}\ u) + (1 \times \overset{\text{O}}{16.00}\ u) = 18.02\ u$$

$$\text{Formula Mass of } Ba(NO_3)_2 = (1 \times \overset{\text{Ba}}{137.33}\ u) + (2 \times \overset{\text{N}}{14.01}\ u) + (6 \times \overset{\text{O}}{16.00}\ u)$$

$$= 261.35\ u$$

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Molar mass of a compound has the same numerical value as molecular mass or formula mass, but with units of grams instead of atomic mass units.

We can simply sum up the molar masses of each atom in the formula.

$$\text{Molar Mass of } H_2O = (2 \times \overset{\text{H}}{1.008}\text{ g}) + (1 \times \overset{\text{O}}{16.00}\text{ g}) = 18.02\ \frac{\text{g}}{\text{mol}}$$

Molar Mass of $Ba(NO_3)_2$

$$= (1 \times \overset{\text{Ba}}{137.33}\text{ g}) + (2 \times \overset{\text{N}}{14.01}\text{ g}) + (6 \times \overset{\text{O}}{16.00}\text{ g})$$

$$= 261.35\text{ g/mol}$$

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Revisiting molar masses and moles of elements …

We considered earlier the concept of “mole” for elements.

1 mole of an element contained 6.022×10^{23} atoms of that element.

That was because an individual atom is the “characteristic particle” that defines an element.

But if the element normally exists as, say, diatomic molecules, we may need to consider moles of those diatomic molecules.

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For example:

1 mol of <u>oxygen, as an element</u>, is composed of 6.022 x 10^{23} oxygen <u>atoms</u>, and has a mass of 16.0 grams.

But 1 mol of <u>oxygen gas</u> (made of O_2 molecules) contains 6.022 x 10^{23} O_2 <u>molecules</u>, and has a mass of 32.0 grams.

So, when O_2 is the material of interest (as a gas, or a reactant, or a product), the "characteristic particle" is O_2 molecule, with a molecular mass of 32 u and a molar mass of 32 g/mol.

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But, when we consider how many moles of oxygen exists in a certain amount of a compound, we will be thinking of <u>atoms</u> of oxygen, <u>not</u> molecules of oxygen.

After all, a compound contains atoms of oxygen, not O_2 molecules.

So, we would use moles of O, with a molar mass of 16.0 g/mol

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Same considerations apply to the following elements, which normally exist as molecular substances, and we need to be aware of the context and use the appropriate molar masses.

Hydrogen	Nitrogen	Oxygen	Fluorine	Chlorine	Bromine	Iodine	Phosphorus	Sulfur
H_2	N_2	O_2	F_2	Cl_2	Br_2	I_2	P_4	S_8

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What is the “characteristic entity” of Carbon?

A Carbon **atom**

What is the “characteristic entity” of Oxygen gas?

An O_2 **molecule**

What is the “characteristic entity” of Calcium Chloride?

A $CaCl_2$ **formula unit**

Remember, ionic compounds don’t have molecules.
Formula unit is the “pretend molecule” that contains the smallest number of ions in the right proportions.

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How many C atoms in 1 mole of Carbon?

How many O_2 molecules in 1 mole of Oxygen gas?

How many O atoms in 1 mole of Oxygen gas?

How many formula units in 1 mole of $CaCl_2$?

How many Cl atoms in 1 mole of $CaCl_2$?

What is the mass of 1 mole of Carbon?

What is the mass of 1 mole of Oxygen gas?

What is the mass of 1 mole of $CaCl_2$?

Practice

For which of the following compounds does 1.00 g represent 2.27×10^{-2} mol?

a) H_2O
b) CO_2
c) NH_3
d) C_2H_6

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Practice

How many moles of oxygen is contained in 2.0 mol CO_2?

This is simple enough to think in terms of proportions:

1 mol CO_2 contains 2 mol O

2.0 mol CO_2 contains $\frac{2.0}{1}(2) = 4.0$ mol O

But we really should start thinking in terms of dimensional analysis, for when things get complicated.

Using dimensional analysis:

$$2.0\ \cancel{\text{mol CO}_2} \times \frac{2\text{ mol O}}{1\ \cancel{\text{mol CO}_2}} = 4.0\text{ mol O}$$

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Practice

How many moles of hydrogen is contained in 1.5 mol NH_3?

Again, in terms of proportions:

1 mol NH_3 contains 3 mol H

1.5 mol NH_3 contains $\frac{1.5}{1}(3) = 4.5$ mol H

Using dimensional analysis:

$$1.5\ \cancel{\text{mol NH}_3} \times \frac{3\ \text{mol H}}{1\ \cancel{\text{mol NH}_3}} = 4.5\ \text{mol H}$$

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Practice

The mass of 0.820 mol of a substance composed of diatomic molecules is 131 g. Identify the molecule.

a) F_2
b) Cl_2
c) Br_2
d) I_2

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Practice

Consider separate 100.0 gram samples of each of the following:

H_2O, N_2O, $C_3H_6O_2$, CO_2

Rank them from greatest to least number of oxygen atoms in the sample.

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Percent Composition of Compounds

Mass percent of an element:

$$\text{mass \%} = \frac{\text{mass of element in compound}}{\text{mass of compound}} \times 100\%$$

Mass % is an intensive property

It doesn't care about sample mass

We can choose a convenient sample mass:

Molar mass if we know the formula

For iron in iron(III) oxide, (Fe_2O_3):

$$\text{mass \%} = \frac{\overbrace{2(55.85\text{ g})}^{\text{Mass of Fe in one mol of } Fe_2O_3}}{\underbrace{2(55.85\text{ g}) + 3(16.00\text{ g})}_{\text{Mass of one mol of } Fe_2O_3}} \text{ x } 100 = 69.94\%$$

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Practice

Consider separate 100.0 gram samples of each:

H_2O, N_2O, $C_3H_6O_2$, CO_2

Rank them from highest to lowest percent oxygen by mass.

The ranking is the same as for number of oxygen atoms per 100 grams

more oxygen atoms ➔ more oxygen mass

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Molecular and Empirical Formulas

Molecular formula

Actual formula of the compound

for example, C_6H_6

Empirical formula

The number of atoms (their subscripts) are the smallest set of integers possible, in the same ratios as in the molecular formula

For $C_6H_{12}O_6$, the empirical formula is CH_2O

For C_6H_6 , the empirical formula is CH

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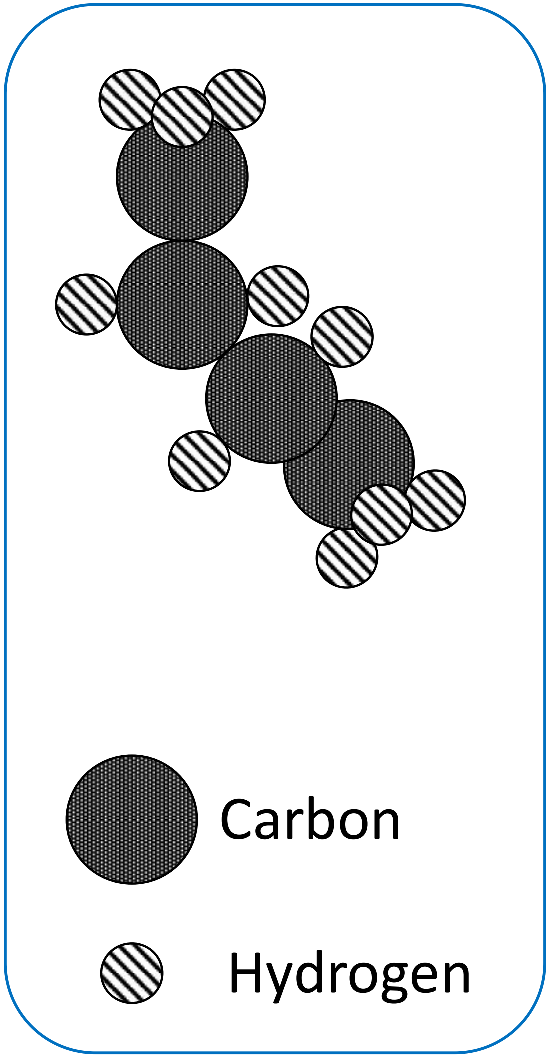
What is the empirical formula of the compound whose molecule looks as follows?



Molecular formula: C_4H_{10}

$\div 2$

Empirical formula: $\mathbf{C_2H_5}$

Subscripts 2 and 5 can’t be simplified further

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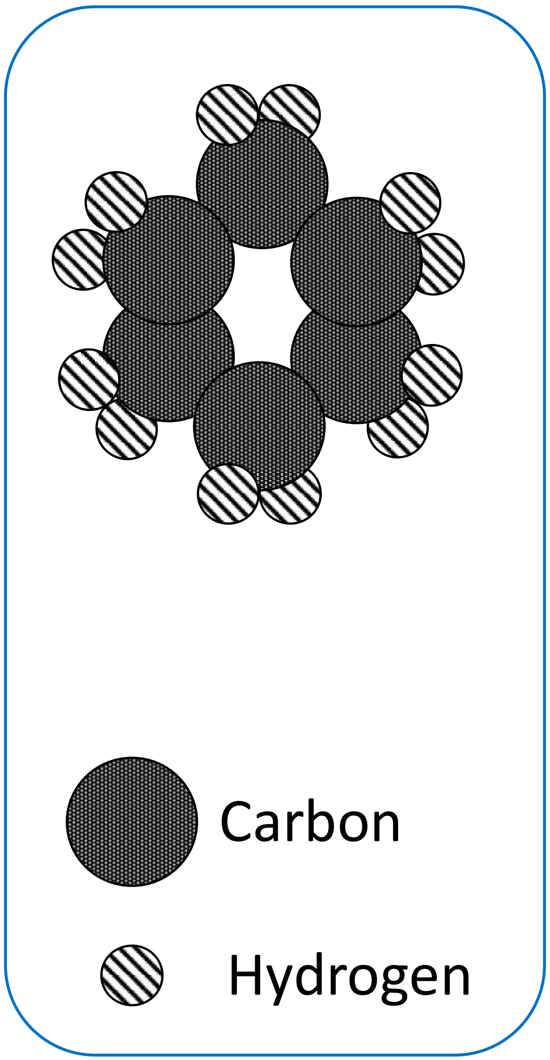
What is the empirical formula of the compound whose molecule looks as follows?



Molecular formula: C_6H_{12}

$\downarrow \div 6$

Empirical formula: CH_2

Subscripts (implied) 1 and 2 can’t be simplified further

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Finding molecular formula, given the molar mass of the compound and its empirical formula

$$\mathbf{n} = \frac{\text{Molar mass of molecular formula}}{\text{“Molar mass” of empirical formula}}$$

Number of atoms in the empirical formula is smaller by the factor **n**

Number of atoms in the molecular formula is greater by the factor **n**

$$\text{Empirical formula} \xrightarrow{\times\ \mathbf{n}} \text{Molecular formula}$$

$$\text{Molecular formula} \xrightarrow{\div\ \mathbf{n}} \text{Empirical formula}$$

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Suppose we know the following:

- The empirical formula is CH
- The molar mass of the compound is 78 g/mol

We can find the molecular formula using the ratio of molar mass to that of the empirical formula (which is 13 g/mol)

Molar mass of molecular formula = 78 g/mol
"Molar mass" of empirical formula CH = 13 g/mol

n = 78 / 13 = **6**

$$CH \xrightarrow{\times 6} C_6H_6$$

$$C_6H_6 \xrightarrow{\div 6} CH$$

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Of course we are not always given the empirical formula

- We can find the empirical formula from elemental percent mass compositions
  - Mass % of each element making up the compound
- Even if we don't know the molar mass

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Finding the **empirical formula** from mass % of elements

-- If we <u>don't know</u> the <u>molar mass</u>

1. Take (exactly) 100g sample (just a convenient number; see next step)
2. Then the mass of each element is equal to its percentage
3. Convert each element's mass to moles of the element
   -- will be "ugly" non-integer values
4. Divide each mole amount by the smallest mole amount
5. If a result is close to an integer (e.g. 1.97=2, 1.02=1, etc.), round off to nearest integer
6. If a result ends with a fraction that resembles **½**, **⅓, ¼** or their multiples (e.g 1.75, 2.49, 1.33, 1.67), multiply all results by the corresponding denominator
   - e.g. multiply each by 3 if one of the results is 1.67 since 0.67 is a multiple (double) of **⅓**
7. Repeat for all fractions (if you multiply each by 4 because of a .25 or .75, do not multiply by 2 because of a .5)
8. The integers obtained are the subscripts in the empirical formula

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Practice

The composition of adipic acid is 49.3% C, 6.9% H, and 43.8% O (by mass). Find its empirical formula.

We consider a 100-g sample of the compound.

Mass percentages equal masses in that case.

$$\textbf{C:}\quad 49.3\text{ g} \times \frac{1\text{ mol}}{12.01\text{ g}} = 4.10\text{ mol} \xrightarrow{\div\, 2.74} \boxed{1.5} \xrightarrow{\times\, 2} \mathbf{3}$$

$$\textbf{H:}\quad 6.9\text{ g} \times \frac{1\text{ mol}}{1.01\text{ g}} = 6.83\text{ mol} \xrightarrow{\div\, 2.74} \boxed{2.5} \xrightarrow{\times\, 2} \mathbf{5}$$

$$\textbf{O:}\quad 43.8\text{ g} \times \frac{1\text{ mol}}{16.0\text{ g}} = 2.74\text{ mol} \xrightarrow{\div\, 2.74} \boxed{1} \xrightarrow{\times\, 2} \mathbf{2}$$

$$\mathbf{C_3H_5O_2}$$

Smallest mol

Need doubling to become integers

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Practice

Using the empirical formula ($C_3H_5O_2$) and the molar mass of adipic acid (146 g/mol), find the molecular formula.

"molar mass" of empirical formula $C_3H_5O_2$ = 73 g/mol

n = 146 / 73 = **2**

$$\mathbf{C_3H_5O_2} \xrightarrow{\times\, 2} \mathbf{C_6H_{10}O_4}$$

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Alternative (better?) method to find molecular formula

We still need molar mass, but don't need the empirical formula

We can apply the mass composition directly on the molar mass of the compound to find the mass (and then moles) of each element in one mole of the compound

Number of moles of an element in one mole of the compound is equal to the number of atoms of the element in one molecule (i.e. the subscript in the molecular formula)

We bypass the potentially cumbersome process of finding the empirical formula first

Molar mass of the compound	Mass % of element (or its mass in a given sample mass)	→	Mass of element in 1 mol of the compound	→	Moles of element in 1 mol of the compound	=	Subscripts in molecular formula

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Practice

The composition of adipic acid is 49.3% C, 6.9% H, and 43.8% O (by mass), and has a molar mass of 146 g/mol. Find the *molecular formula without finding the empirical formula first.*

We consider the mass of 1 mol of the compound: 146 g

We then calculate the mass and moles of each element

Moles of element in 1 mol of compound = Subscript in formula

$$C_6H_{10}O_4$$

C: $(0.493)(146\text{ g}) = 72.0\text{ g} \times \dfrac{1\text{ mol}}{12.01\text{ g}} = 6\text{ mol}$

H: $(0.069)(146\text{ g}) = 10.1\text{ g} \times \dfrac{1\text{ mol}}{1.01\text{ g}} = 10\text{ mol}$

O: $(0.438)(146\text{ g}) = 64.0\text{ g} \times \dfrac{1\text{ mol}}{16.0\text{ g}} = 4\text{ mol}$

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Given molar mass and % mass compositions, we can find the empirical formula by continuing the alternative method:

Find molecular formula first

Simplify it to the smallest integer subscripts

- By dividing the subscripts in the molecular formula by their "greatest common factor"

- In practice it amounts to dividing them by small integers like 2 or 3 and stop when doing so would yield fractional numbers

For example:

Given the molecular formula of $C_6H_{10}O_4$

Divide by 2

Empirical formula: $C_3H_5O_2$

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By the way:

Remember, ionic compounds don’t have molecules.

The formula for an ionic compound already has the simplest ratios of the ions.

$NaCl$ **not** Na_2Cl_2

not Na_3Cl_3

The chemical formula of an ionic compound is typically the empirical formula.

Typically? 🤔

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Sometimes the formula of an ionic compound is not its empirical formula.

If one or both of the ions are polyatomic, the simplest ratio of ions might not correspond to the empirical formula.

$Na_2C_2O_4$ is the formula for sodium oxalate and it is different from its empirical formula $NaCO_2$

The formula for compounds of mercury (I) ion, Hg_2^{2+}, must contain Hg in pairs:

Hg_2Cl_2 and not HgCl

$(Hg_2)_3(PO_4)_2$ and not Hg_3PO_4

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So, the formula of an ionic compound may still be a multiple of its empirical formula, with the right polyatomic ion, with the right charge, combined with the right ion of the opposite charge.

But we generally avoid dealing with ionic compounds when it comes to “molecular formula” versus empirical formula, since they don’t have actual molecules.

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The empirical formula of styrene is CH; its molar mass is 104.1 g/mol. What is the molecular formula of styrene?

a) C_2H_4
b) C_8H_8
c) $C_{10}H_{10}$
d) C_6H_6

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Naming Binary Compounds (Nomenclature)

Binary Compounds
(Composed of 2 elements)

metal—nonmetal
➡ **Ionic** Compound

nonmetal—nonmetal
➡ **Covalent** Compound

Assumed ionic, and named accordingly.
Sometimes it's not really ionic, but even then it is named as if it were ionic.

Naming Binary (two elements) Ionic Compounds

- If the first element in the formula is a metal, it is an ionic compound
- The metal is the cation.
- Cation name first, anion name second.

{cation name} {anion name}

If the metal forms only one kind of cation:

name of the parent element

sodium, calcium, zinc, etc.

root of the element name and adding ***–ide***

*Chlorine becomes chlor**ide***

*Oxygen becomes ox**ide***

What if the metal can form cations with different charges?

See next page

Naming Binary Ionic Compounds (continued)

What if the metals in compound can form cations with more than one charge?

Similar to the simple, "one charge" case, except:

- Charge on the metal ion must be specified with a Roman numeral in parentheses
- **Transition metal cations** (except for Ag, Zn, Cd) **usually require a Roman numeral**.
- **Lead** (Pb) and **Tin** (Sn) require a Roman numeral
- Metals that form only one cation should not be identified by a roman numeral

Examples:

CuBr	Copper(I) bromide
FeS	Iron(II) sulfide
PbO_2	Lead(IV) oxide

Polyatomic Ions

- Ions containing more than one atom ("poly"=many)
  - like molecules, except that they carry a charge.
  - each has a definite, characteristic charge.
- Almost all (here) are anions, except NH_4^+ and Hg_2^{2+}
- Naming ionic compounds of polyatomic ions are just like binary ionic compounds.
- The polyatomic anion name follows the cation name (which may itself be polyatomic)

Common Polyatomic Ions

Hg_2^{2+}	mercury(I)
NH_4^+	ammonium
NO_2^-	nitrite
NO_3^-	nitrate
SO_3^{2-}	sulfite
HSO_3^{2-}	hydrogen sulfite; bisulfite
SO_4^{2-}	sulfate
HSO_4^-	hydrogen sulfate; bisulfate
$S_2O_3^{2-}$	thiosulfate
OH^-	hydroxide
CN^-	cyanide
OCN^-	cyanate
SCN^-	thiocyanate

PO_4^{3-}	phosphate
HPO_4^{2-}	hydrogen phosphate
$H_2PO_4^-$	dihydrogen phosphate
CO_3^{2-}	carbonate
HCO_3^-	hydrogen carbonate; bicarbonate
ClO^-	hypochlorite
ClO_2^-	chlorite
ClO_3^-	chlorate
ClO_4^-	perchlorate
$C_2H_3O_2^-$	acetate
MnO_4^-	permanganate
CrO_4^{2-}	chromate
$Cr_2O_7^{2-}$	dichromate
O_2^{2-}	peroxide
$C_2O_4^{2-}$	oxalate

- Don’t rush to memorize the names and formulas
- Memorize polyatomic ion names/formulas only if the connections and rules we will see are not enough to derive them

We can figure out polyatomic ion charges:

odd number of **odd-**group atoms gives **odd** charge

Goes with odd atomic number

If the charge is odd:

- Almost always –1 for anions (except –3 for phosphate)
- +1 for ammonium)

If the charge is not odd (i.e. even)

- –2 for anions
- +2 for the Hg_2^{2+} cation

Even/Odd part of the charge is "always" correct

For ions not encountered in General Chemistry, the charge magnitude can be (far) beyond –3 or +3

odd number of **odd-**group atoms gives **odd** charge

In the examples below, following elements have odd atomic numbers

H: atomic no.=1 (Group **1**)	**N**: atomic no.=7 (Group **15**; or **5**A)	**Cl**: atomic no.=17 (Group **17**; or **7**A)	**Mn**: atom no.=25 (Group **7**)

ClO_3^-	**one** atom with odd atomic number	➡	charge should be **odd**	➡	−1
CO_3^{2-}	**zero** atoms with odd atomic number	➡	charge should be **even**	➡	−2
NO_3^-	**one** atom with odd atomic number	➡	charge should be **odd**	➡	−1
SO_4^{2-}	**zero** atoms with odd atomic number	➡	charge should be **even**	➡	−2
HPO_4^{2-}	**two** atoms with odd atomic number	➡	charge should be **even**	➡	−2
HCO_3^-	**one** atom with odd atomic number	➡	charge should be **odd**	➡	−1
$Cr_2O_7^{2-}$	**zero** atoms with odd atomic number	➡	charge should be **even**	➡	−2
MnO_4^-	**one** atom with odd atomic number	➡	charge should be **odd**	➡	−1
$C_2O_4^{2-}$	**zero** atoms with odd atomic number	➡	charge should be **even**	➡	−2
$C_2H_3O_2^-$	**three** atoms with odd atomic number	➡	charge should be **odd**	➡	−1

Can you predict the charges on the following <u>anions</u> not on your usual list of polyatomic ions?

Only hint: the magnitude of the charge will not exceed 2

$CF_3SO_3^{x}$

$C_4O_4^{x}$

Naming of Oxyanions

- Almost all of the anions we deal with here are "oxyanions"
  - They contain oxygen in addition to a non-oxygen "central" element
- If there is only <u>one kind</u> of common oxyanion of an element (it combines with only a certain number of oxygen atoms to form its <u>one</u> oxyanion):

  {root of element name} ate

- If the element can form <u>two kinds</u> of oxyanions (each with a different number of oxygens), then:
  - the **low-oxygen** one ends with –ite

    {root of element name} ite

  - the **high-oxygen** one again ends with –ate

    {root of element name} ate

NO_2^-	Nitrite	Low oxygen content
NO_3^-	Nitrate	High oxygen content

SO_3^{2-}	Sulfite	Low oxygen content
SO_4^{2-}	Sulfate	High oxygen content

- Unfortunately there is no fixed number of oxygens corresponding to –ite and –ate
- We only know that –ite goes with less oxygen, and –ate goes with more oxygen, when there are two choices
  - ➢ 2 and 3 for oxyanions of nitrogen, but …
  - ➢ 3 and 4 for oxyanions of sulfur, etc.

Naming of Oxyanions of Halogens

- **Halogens** can form **four** kinds of oxyanions
  - With 1, 2, 3, 4 oxygens (pretty simple to remember)
- Middle two are the “regular” low-oxygen and high-oxygen ions
- The ion with even less oxygen than the “low-oxygen” ion end with the same –ite ending, but takes a prefix of **“hypo”**
- The ion with even more oxygen than the “high-oxygen” ion end with the same –ate ending, but takes a prefix of **“per”**

Higher than “high”

ClO_4^-	Perchlorate	Highest oxygen content (4)
ClO_3^-	Chlorate	High oxygen content (3)
ClO_2^-	Chlorite	Low oxygen content (2)
ClO^- or OCl^-	Hypochlorite	Lowest oxygen content (1)

Lower than “low”

Chlorine, Bromine, and Iodine form analogous oxyanions:

BrO_3^-	Bromate	
IO_4^-	Periodate	
IO^-	Hypoiodite	
BrO_2^-	Bromite	
FO^-	Hypofluorite	The only oxyanion of fluorine that actually exists!

Examples of ionic compounds with polyatomic ions:

$NaOH$	Sodium hydroxide
$Mg(NO_3)_2$	Magnesium nitrate
$(NH_4)_2SO_4$	Ammonium sulfate

If multiple polyatomic ions are needed in the formula, they are enclosed in parentheses before putting their count as subscript

An element's most stable ion forms an ionic compound with chlorine having the formula XCl_2. If the ion has 36 electrons, what is the element that produces the ion?

a) Kr
b) Se
c) Sr
d) Rb
e) None of these

______ form ions with a 2+ charge when they react with nonmetals.

a) Alkali metals
b) Alkaline earth metals
c) Halogens
d) Noble gases
e) None of these

Which is *not* the correct chemical formula for the compound named?

a) potassium phosphate, K_3PO_4
b) iron(II) oxide, FeO
c) calcium carbonate, $CaCO_3$
d) sodium sulfide, NaS
e) lithium nitrate, $LiNO_3$

Naming <u>Binary</u> <u>Covalent</u> Compounds

Formed between <u>two nonmetals</u>.

- Naming scheme modeled after ionic compounds
- But we need to indicate the number of atoms of each element because they can't be predicted

The first element in the formula is named first, using the full element name (just like for the cations)

The second element is named as if it were an anion, with an *–ide* added to the root name of the element

up to this point, same as for ionic compounds, but …

- <u>Prefixes</u> are used to denote the <u>numbers</u> of atoms present.
- The prefix *mono-* is never used for naming the first element.

Prefixes Used to Indicate Number in Chemical Names

Prefix	Number
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

Binary Covalent Compounds Examples:

CO	Carbon monoxide
CO_2	Carbon dioxide
SF_6	Sulfur hexafluoride
N_2O_4	Dinitrogen tetroxide

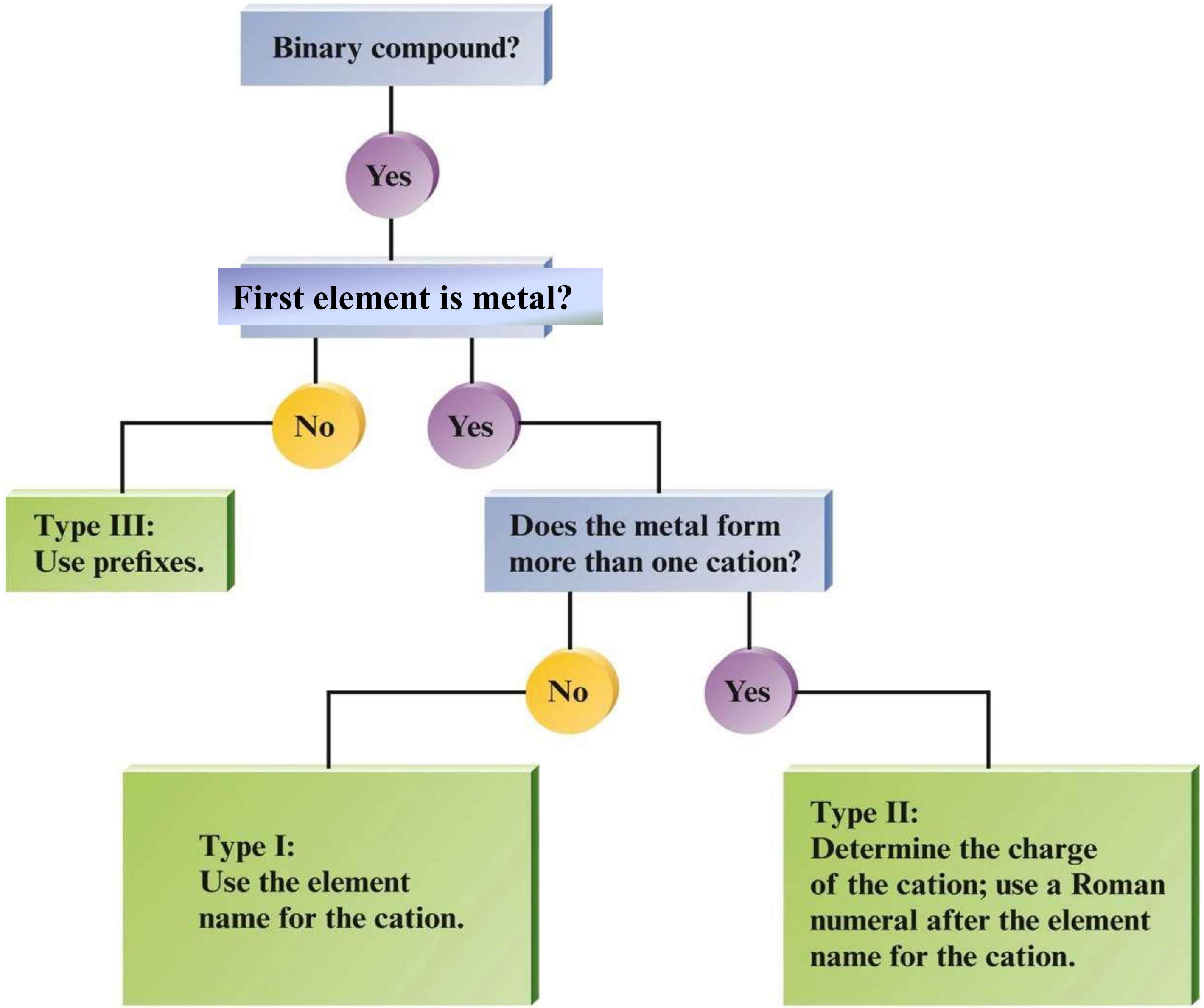
"a" or "o" at the end of the prefix is dropped if followed by "o":

monooxide ➔ monoxide

pentaoxide ➔ pentoxide

This is basically relevant only for prefixing oxygen

Simplified Flowchart for Naming Binary Compounds



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Which of the following is named incorrectly?

a. Li_2O, lithium oxide
b. $FePO_4$, iron(III) phosphate
c. HF, hydrogen fluoride
d. $BaCl_2$, barium dichloride
e. Mg_3N_2, magnesium nitride

Why didn’t we consider “hydrogen fluoride” wrong?

Why didn’t we say it should be
“hydrogen monofluoride”?

Covalent compounds whose formula starts with hydrogen (H) are “acidic”, and when they are named as an ordinary compound rather than as an acid, the prefixes are not used.

Acidic molecules “lose” the acidic hydrogens as H^+ cations.

What remains is an anion whose charge is known.

There is one anion, and the number of H atoms in the formula is equal to the charge of the anion produced.

The number of atoms of each element is known and fixed.

So:
We skip prefixes when naming acidic molecules as "compounds" (acid names are a whole different thing that we will look at later)

Which is the correct formula for copper(I) sulfide?

a) CuS
b) Cu_2S
c) CuS_2
d) Cu_2S_2
e) None of these

Which of the following is the correct chemical formula for iron(III) oxide?

a) FeO
b) Fe_3O
c) FeO_3
d) Fe_2O_3
e) Fe_3O_2

What is the correct name for the compound with the formula $Mg_3(PO_4)_2$?

a) Trimagnesium diphosphate
b) Magnesium(II) phosphate
c) Magnesium phosphate
d) Magnesium(II) diphosphate
e) Magnesium(III) diphosphate

Acids

- Acids can be recognized by:

  the **hydrogen that appears <u>first</u>** in the formula.
  For example, HCl or $HC_2H_3O_2$

- Molecule with one or more ionizable H atoms

- When the molecule acts as an acid, the acidic, ionizable H becomes an H^+, leaving behind an anion.
  - The acid **molecule** is producing ions, but it is not an ionic compound; it is a **molecular** compound.

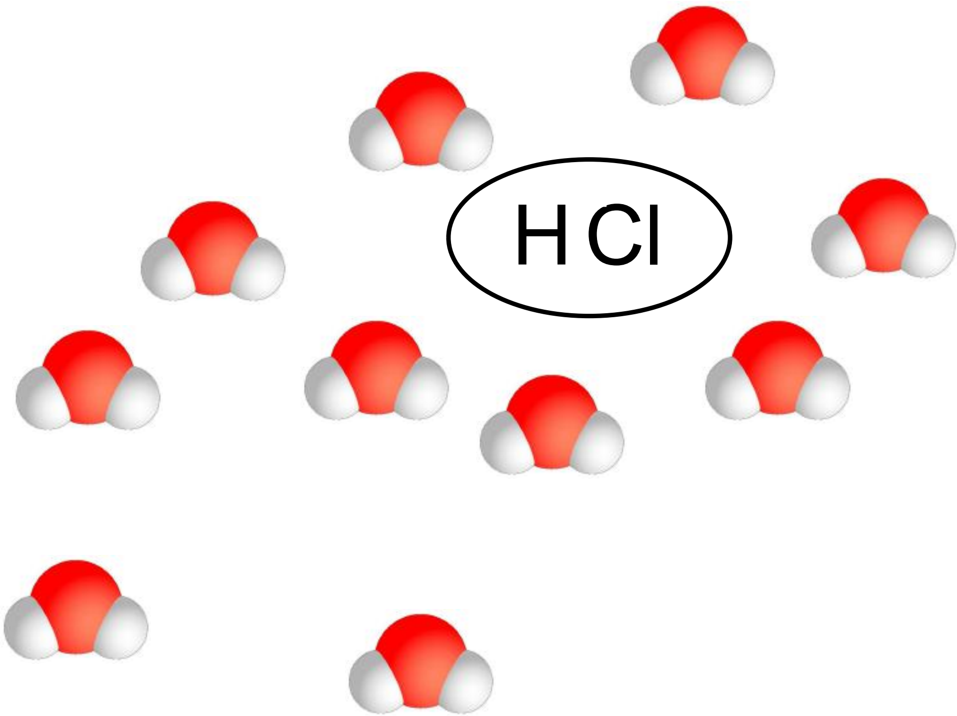
$$HA \longrightarrow H^+ + A^-$$

acidic hydrogen

anion produced by the acid

A relevant historical note

- Lavoisier (**<u>incorrectly</u>**) thought oxygen was what made a substance acidic.
  - "Oxygen" means "acid generator"
- When faced with acids that refused to reveal any oxygen in them (such as HCl), chemists of the time thought that the original substance somehow got oxygenated by water when dissolved, and <u>then</u> became acidic.
- Therefore, acids with no oxygen were regarded and named as "acid" only in aqueous solution.
- **That naming tradition survives to this day**.



Hydrogen Chloride (g)

Molecular H-Cl *bond is covalent, not ionic*

Hydrochloric acid (aq)

HCl *in water:*
in the form of ions H^+ *and* Cl^-
No chemical bond between H^+ *and* Cl^-

Of course what "hydrogen chloride" does in water has nothing to do with oxygen.
But the tradition of naming acids without oxygen as "acids" only when in aqueous solution survives.

More on the presence and absence of oxygen in acids …

- Monatomic (single atom) anion names all end with ***–ide***
  - Sulfide, chloride, etc.
  - Oxide (obviously has oxygen), combined with H^+ forms water, which is not named as an acid; so not relevant
- Cyanide CN^- also ends with ***–ide***, even though it is not monatomic
  - It acts like the anion of a halogen in some ways, so early chemists could have lumped it with the halides and named similarly
- Anions with names ending with ***–ide*** indeed don't have oxygen
- But there are some anions without oxygen whose names don't end in ***–ide***

More on the presence and absence of oxygen in acids …

Some anions without oxygen have names that don’t end with ***–ide***

e.g. thiocyan**ate**, SCN^-

Why are we talking about the endings of anion names?

Because acid naming scheme ultimately cares about:

the **anion name ending**

not the presence/absence of oxygen in the molecule

— Many textbooks and sources focus on the oxygen

— Understandable, and almost right; but not quite right

In short …

Acid naming scheme only cares about the anion name ending.

Three possibilities for anion name ending:

-ide

-ate

-ite

Naming Acids

If the anion name ends with *–ide*

Its acid is named with the prefix ***hydro–*** and the suffix ***–ic***.

{Hydro} {root} {ic} acid

root name of the anion formed by the acid

For acids whose anions end with *–ide* **only:**

Named as acid only if they are in aqueous solution

Chloride	(Cl^-)	➡	$HCl(aq)$	Hydrochloric acid
Cyanide	(CN^-)	➡	$HCN(aq)$	Hydrocyanic acid
Sulfide	(S^{2-})	➡	$H_2S(aq)$	Hydrosulfuric acid

means "water";
related to naming these compounds as acid only in water

Naming Acids

If the anion name ends with *–ide*

The <u>pure compound is named as a binary covalent compound</u> (or *as if it <u>were</u>,* if the anion has more than one atom like CN^-)

$HCl(g)$ Hydrogen chloride

$HCN(g)$ Hydrogen cyanide

Except we don't use Greek prefixes because we know how many hydrogens are required to make a neutral compound:

As many as the charge the anion alone would have

For example, $H_2S(g)$ is "hydrogen sulfide", not "dihydrogen sulfide"

<u>Remember</u>: this is only when the anion name ends with *-ide*

Naming Acids

If the anion name ends with ***–ate***

The suffix ***–ic*** is added to the root name

{root}{ic} {acid}

Examples:

Nitr**ate** (NO_3^-)	⟶	HNO_3	Nitr**ic** acid
Acet**ate** ($C_2H_3O_2^-$)	⟶	$HC_2H_3O_2$	Acet**ic** acid
Sulf**ate** (SO_4^{2-})	⟶	H_2SO_4	Sulf*ur***ic** acid

Always named as an acid, aqueous solution or not.

No "hydrogen nitrate" or "hydrogen acetate"!

Naming Acids

If the anion name ends with ***–ite***

The suffix ***–ous*** is added to the root name

{root}{ous} {acid}

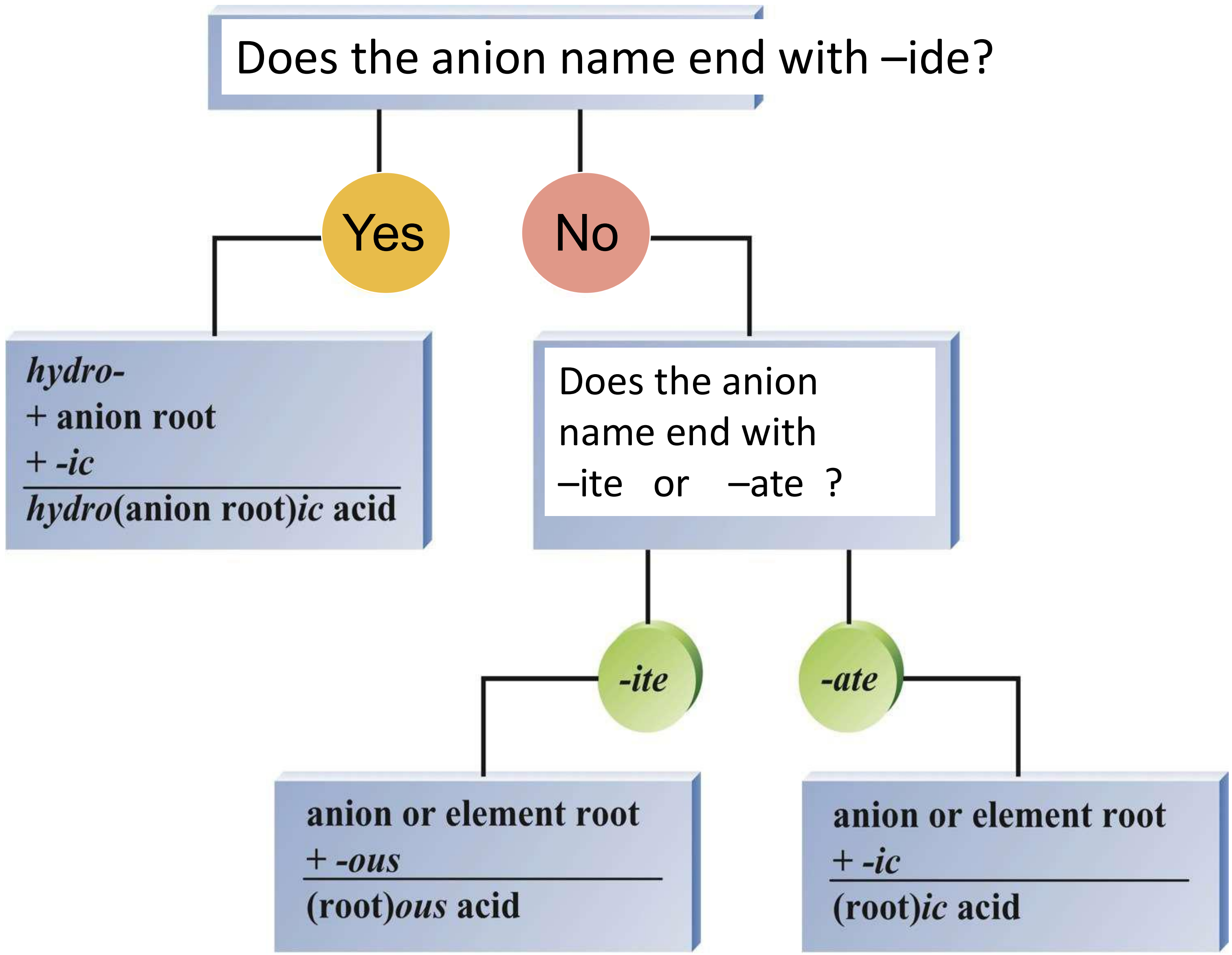
Examples:

Nitr**ite** (NO_2^-)	⟶	HNO_2	Nitr**ous** acid
Chlor**ite** (ClO_2^-)	⟶	$HClO_2$	Chlor**ous** acid
Sulf**ite** (SO_3^{2-})	⟶	H_2SO_3	Sulf*ur***ous** acid

Always named as an acid, aqueous solution or not.

No “hydrogen nitrite” or “hydrogen chlorite”!

Flowchart for Naming Acids



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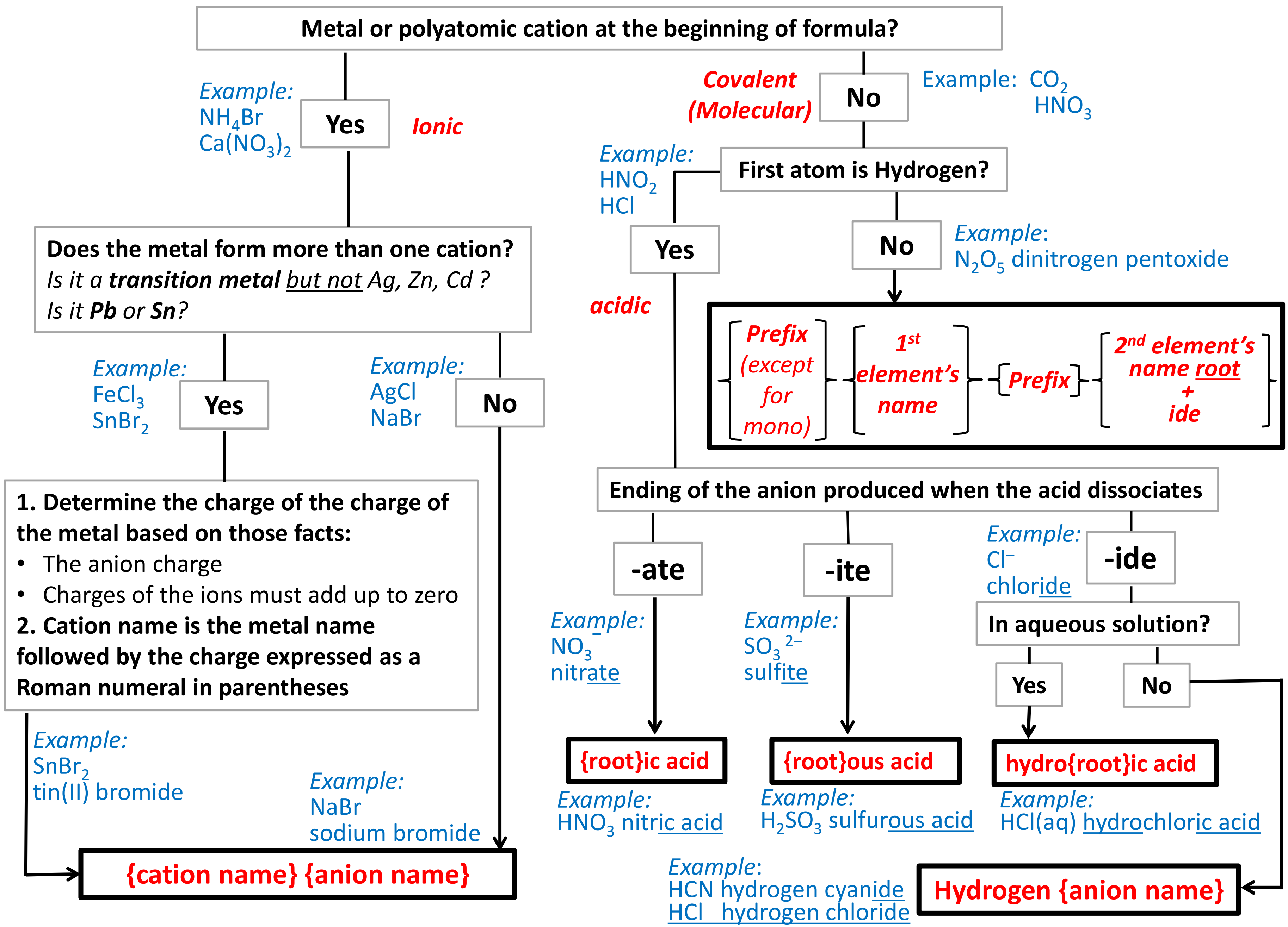
Complete Flowchart for Naming Simple Inorganic Compounds (may include polyatomic ions)



What is the correct name for the acid with the formula HFO?

a) Fluoric Acid
b) Hydrofluoric Acid
c) Hydrofluorous Acid
d) Hypofluorous Acid
e) Perfluoric Acid

Which of the following compounds is named incorrectly?

a) KNO_3 potassium nitrate
b) TiO_2 titanium(II) oxide
c) $Sn(OH)_4$ tin(IV) hydroxide
d) PBr_5 phosphorus pentabromide
e) $CaCrO_4$ calcium chromate