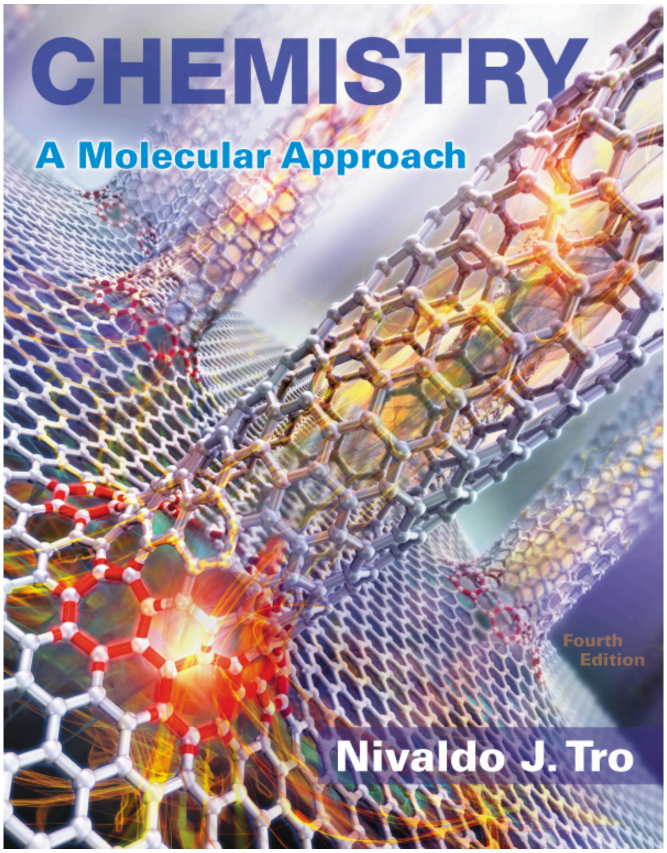


Lecture Presentation

Chapter 3

Molecules, Compounds, and Chemical Equations

How Many Different Substances Exist?

- Elements combine with each other to form *compounds*.
- The great diversity of substances that we find in nature is a direct result of the ability of elements to form compounds.

Hydrogen, Oxygen, and Water

- The dramatic difference between the elements hydrogen and oxygen and the compound water is typical of the differences between elements and the compounds that they form.
- *When two or more elements combine to form a compound, an entirely new substance results.*

Hydrogen, Oxygen, and Water

Selected Properties	Hydrogen 	Oxygen 	Water 
Boiling Point	−253 °C	−183 °C	100 °C
State at Room Temperature	Gas	Gas	Liquid
Flammability	Explosive	Necessary for combustion	Used to extinguish flame

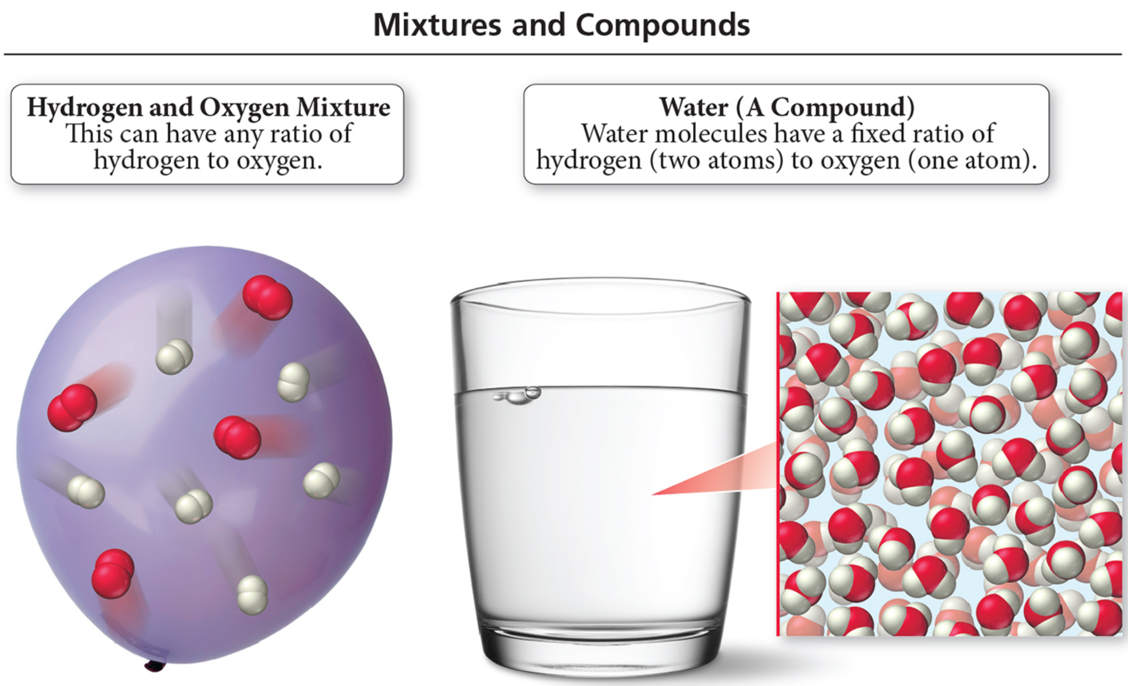
Mixtures and Compounds

- In a mixture, elements can mix in any proportions whatsoever (hydrogen, H_2, and oxygen, O_2).
- In a compound, elements combine in fixed, definite proportions (water, H_2O).

Definite Proportion

- A hydrogen–oxygen mixture can have any proportions of hydrogen and oxygen gas.
- Water, by contrast, is composed of water molecules that always contain two hydrogen atoms to every one oxygen atom.
- Water has a definite proportion of hydrogen to oxygen.

Definite Proportion



Chemical Bonds

- Compounds are composed of atoms held together by ***chemical bonds***.
- Chemical bonds result from the attractions between the charged particles (the electrons and protons) that compose atoms.
- Chemical bonds are broadly classified into two types:
  - ionic and
  - covalent.

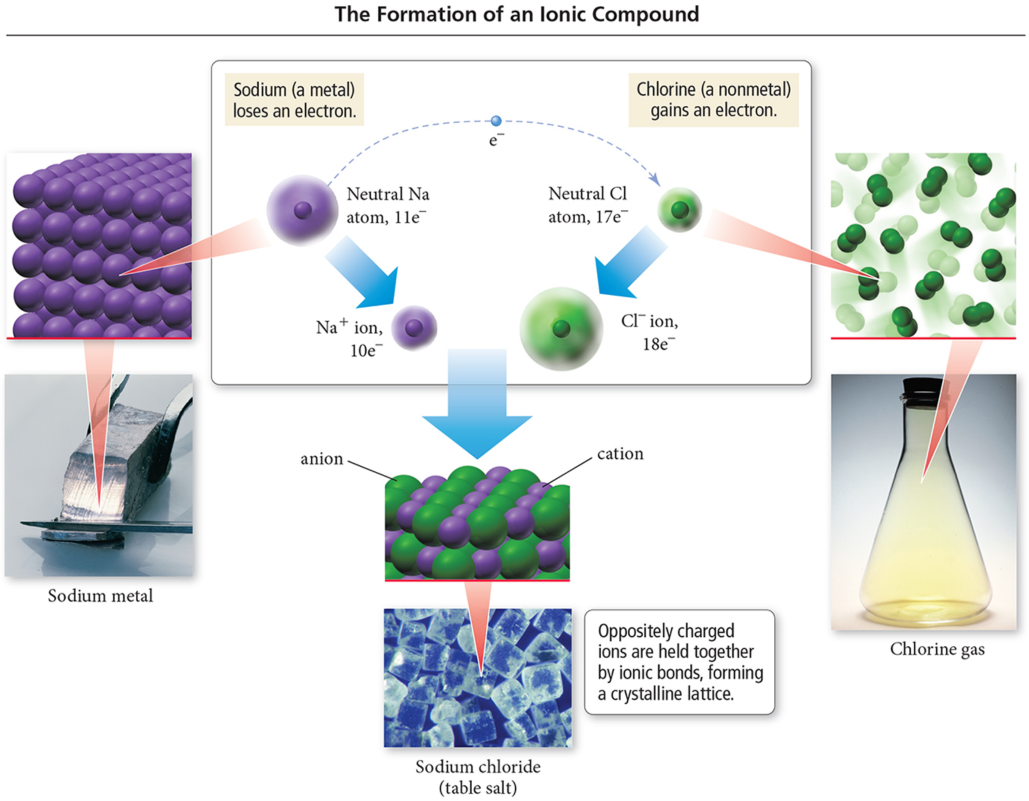
Ionic Bonds

- *Ionic bonds*, which occur between metals and nonmetals, involve the *transfer* of electrons from the metal atom to the nonmetal atom.
- The metal atom then becomes a *cation* while the nonmetal atom becomes an *anion.*
- These oppositely charged ions attract one another by electrostatic forces and form an **ionic bond**.

Ionic Compounds in Solid Phase

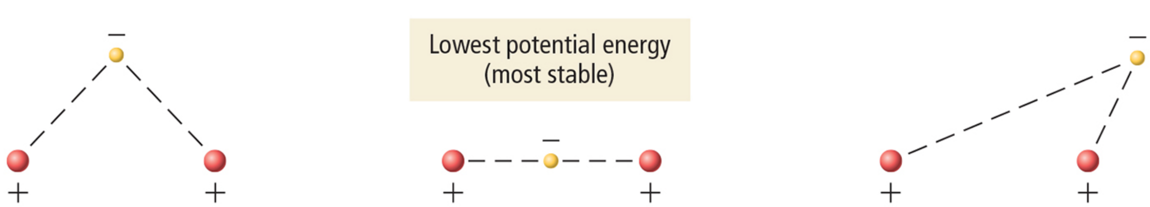
- In the solid phase, the ionic compound is composed of a lattice—a regular three-dimensional array—of alternating cations and anions.

The Formation of Ionic Compounds



Covalent Bonds

- ***Covalent bonds*** occur between two or more nonmetals. The two atoms *share* electrons between them, composing a *molecule*.
- Covalently bonded compounds are also called **molecular compounds**.



Representing Compounds: Chemical Formulas and Molecular Models

- A compound's **chemical formula** indicates the elements present in the compound and the relative number of atoms or ions of each.
  - Water is represented as H_2O.
  - Sodium Chloride is represented as NaCl.
  - Carbon dioxide is represented as CO_2.
  - Carbon tetrachloride is represented as CCl_4.

Types of Chemical Formulas

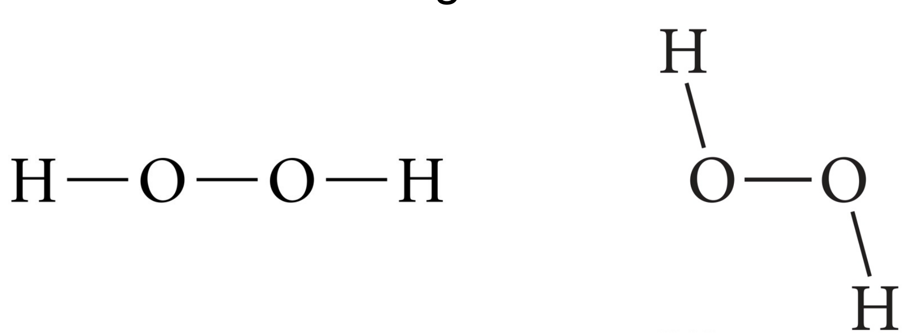
- Chemical formulas can generally be categorized into three different types:
  - empirical formulas,
  - molecular formulas, and
  - structural formulas.

Types of Chemical Formulas

- An **empirical formula** gives the *relative* number of atoms of each element in a compound.
- A **molecular formula** gives the *actual* number of atoms of each element in the molecule of a compound.
  - (a) For H_2O_2, the greatest common factor is 2. The empirical formula is therefore HO.
  - (b) For B_2H_6, the greatest common factor is 2. The empirical formula is therefore BH_3.
  - (c) For CCl_4, the only common factor is 1, so the empirical formula and the molecular formula are identical.

Types of Chemical Formulas

- A **structural formula** uses lines to represent covalent bonds and shows how atoms in a molecule are connected or bonded to each other.
- It can also show the molecule's geometry.
- The structural formula for H_2O_2 can be shown as either of the following.

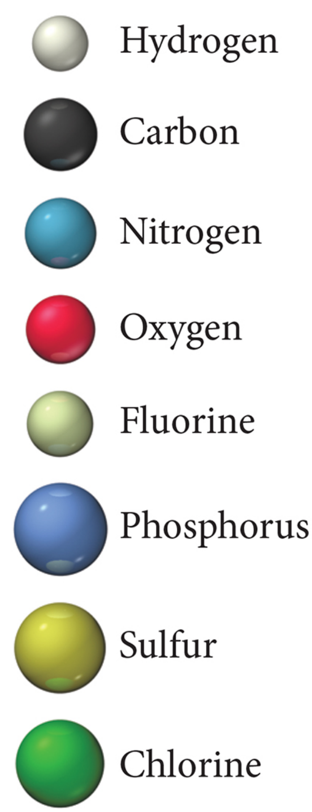


Types of Chemical Formulas

- The type of formula we use depends on how much we know about the compound and how much we want to communicate.
- A structural formula communicates the most information. An empirical formula communicates the least.

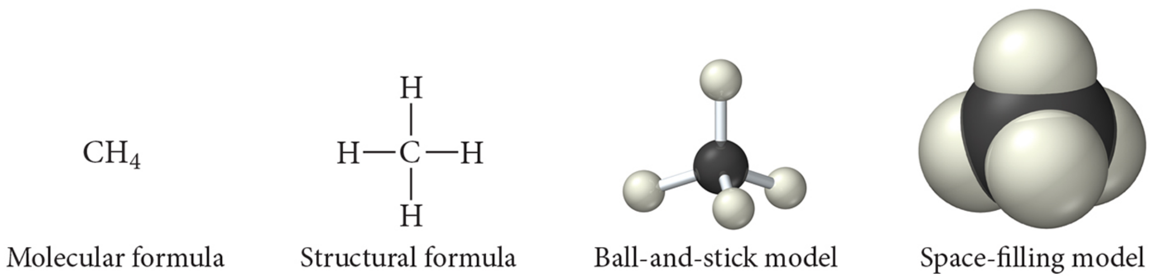
Molecular Models

- A *molecular model* is a more accurate and complete way to specify a compound.
- A **ball-and-stick molecular model** represents atoms as balls and chemical bonds as sticks; how the two connect reflects a molecule’s shape.
- The balls are typically color-coded to specific elements.



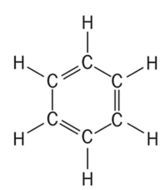
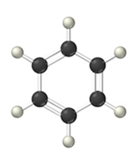
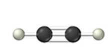
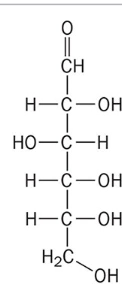
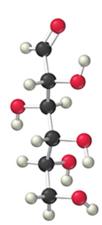
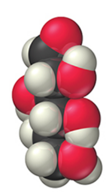
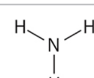
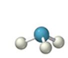
Molecular Models

- In a **space-filling molecular model**, atoms fill the space between each other to more closely represent a best estimate for how a molecule might appear if scaled to visible size.



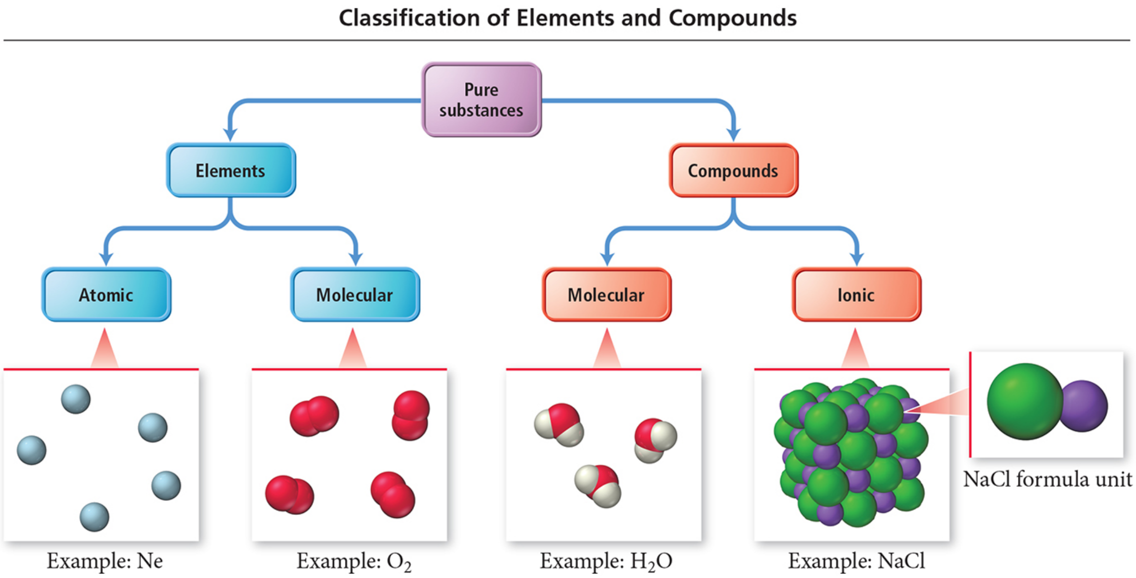
Ways of Representing a Compound

TABLE 3.1 Benzene, Acetylene, Glucose, and Ammonia

Name of Compound	Empirical Formula	Molecular Formula	Structural Formula	Ball-and-Stick Model	Space-Filling Model
Benzene	CH	C_6H_6			
Acetylene	CH	C_2H_2	H—C≡C—H		
Glucose	CH_2O	$C_6H_{12}O_6$			
Ammonia	NH_3	NH_3			

An Atomic-Level View of Elements and Compounds

- Elements may be either atomic or molecular.
- Compounds may be either molecular or ionic.

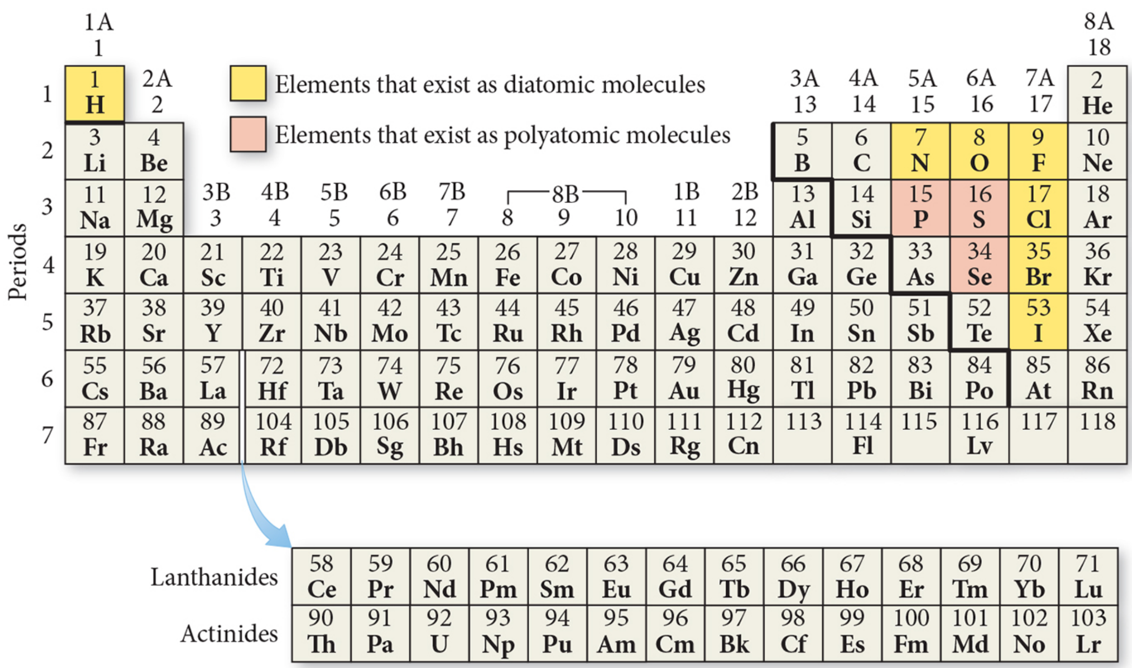


View of Elements and Compounds

- **Atomic elements** exist in nature with single atoms as their basic units. Most elements fall into this category.
  - Examples are Na, Ne, K, Mg, etc.
- **Molecular elements** do not normally exist in nature with single atoms as their basic units; instead, they exist as molecules—two or more atoms of the element bonded together.
  - There are only seven diatomic elements and they are H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2.
  - Also, P_4 and S_8 are polyatomic elements.

Molecular Elements

Molecular Elements



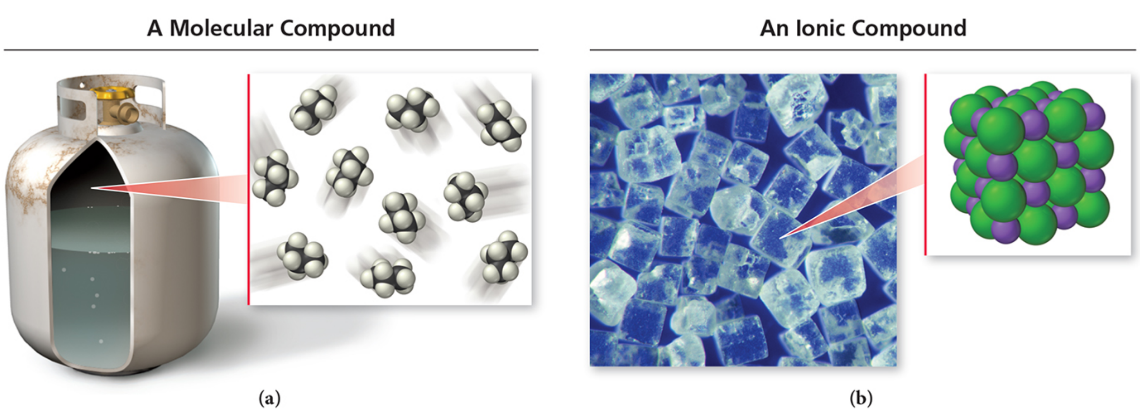
Molecular Compounds

- **Molecular compounds** are usually composed of two or more covalently bonded nonmetals.
- The basic units of molecular compounds are **molecules** composed of the constituent atoms.
  - Water is composed of H_2O molecules.
  - Dry ice is composed of CO_2 molecules.
  - Propane (often used as a fuel for grills) is composed of C_3H_8 molecules.

Ionic Compounds

- **Ionic compounds** are composed of cations (usually a metal) and anions (usually one or more nonmetals) bound together by ionic bonds.
- The basic unit of an ionic compound is the **formula unit**, the smallest, electrically neutral collection of ions.
- Table salt is an ionic compound with the formula unit NaCl, which is composed of Na^+ and Cl^- ions in a one-to-one ratio.

Molecular and Ionic Compounds



Polyatomic Ions

- Many common ionic compounds contain ions that are themselves composed of a group of covalently bonded atoms with an overall charge.
- This group of charged species is called polyatomic ions.
  - $NaNO_3$ contains Na^+ and NO_3^-.
  - $CaCO_3$ contains Ca^{2+} and CO_3^{2-}.
  - $Mg(ClO_3)_2$ contains Mg^{2+} and ClO_3^-.

Ionic Compounds: Formulas and Names

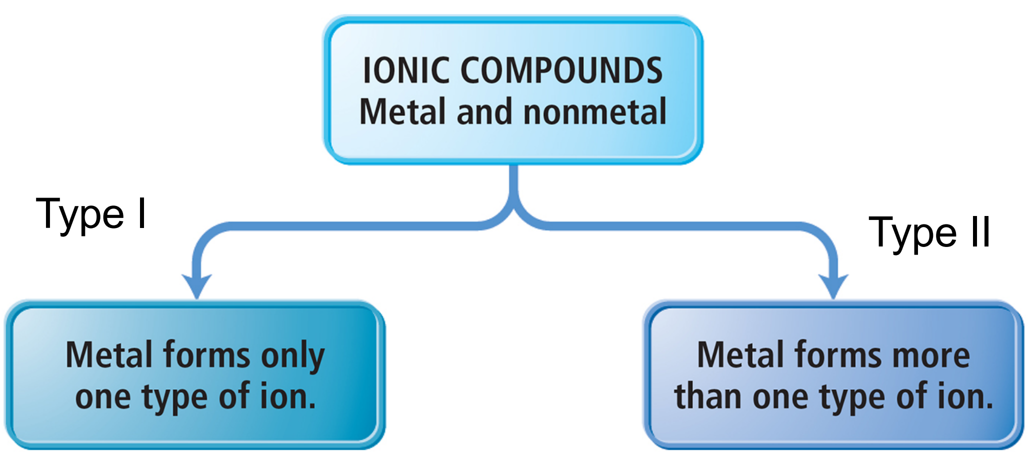
- ***Summarizing Ionic Compound Formulas***
  - Ionic compounds always contain positive and negative ions.
  - In a chemical formula, the sum of the charges of the cations must equal the sum of the charges of the anions.
  - The formula of an ionic compound reflects the smallest whole-number ratio of ions.

Ionic Compounds: Formulas and Names

- The charges of the representative elements can be predicted from their group numbers.
- The representative elements form only one type of charge.
- Transition metals tend to form multiple types of charges.
- Hence, their charges cannot be predicted as in the case of most representative elements.

Naming Ionic Compounds

- Ionic compounds can be categorized into two types, depending on the metal in the compound.



Naming Type I Ionic Compounds

- Type I ionic compounds contain a metal whose charge is invariant from one compound to another when bonded with a nonmetal anion.
- The metal ion always has the same charge.

Metals Whose Charge Are Invariant from One Compound to Another

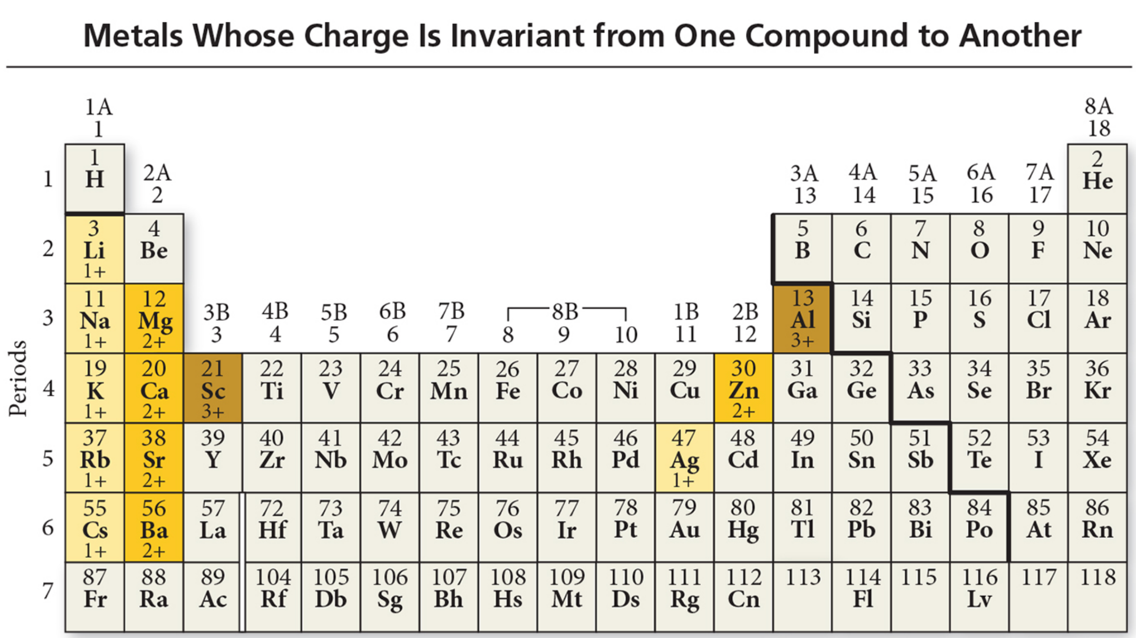


TABLE 3.2 Some Common Monoatomic Anions

Nonmetal	Symbol for Ion	Base Name	Anion Name
Fluorine	F^-	fluor	Fluoride
Chlorine	Cl^-	chlor	Chloride
Bromine	Br^-	brom	Bromide
Iodine	I^-	iod	Iodide
Oxygen	O^{2-}	ox	Oxide
Sulfur	S^{2-}	sulf	Sulfide
Nitrogen	N^{3-}	nitr	Nitride
Phosphorus	P^{3-}	phosph	Phosphide

Naming Binary Ionic Compounds of Type I Cations

- **Binary compounds** contain only two different elements. The names of binary ionic compounds take the following form:

name of cation (metal)

base name of anion (nonmetal) + *-ide*

Examples: Type I Binary Ionic Compounds

- The name for KCl consists of the name of the cation, *potassium*, followed by the base name of the anion, *chlor*, with the ending -*ide*.
  - KCl is *potassium chloride*.
- The name for CaO consists of the name of the cation, *calcium*, followed by the base name of the anion, *ox*, with the ending *-ide.*
  - CaO is *calcium oxide*.

Naming Type II Ionic Compounds

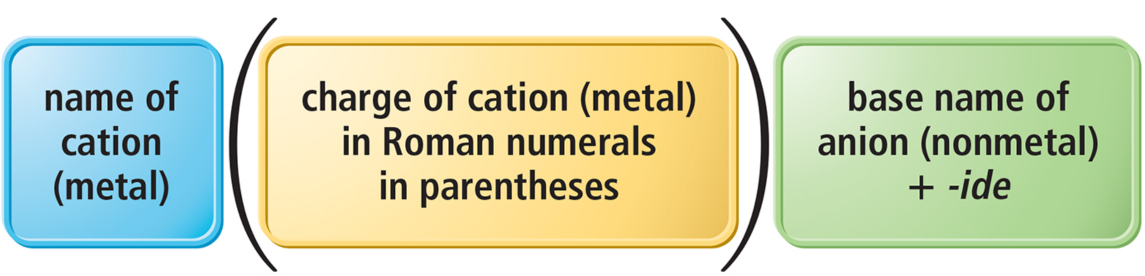
- The second type of ionic compound contains a metal that can form more than one kind of cation, depending on the compound, bonded to a nonmetal anion.
- The metal's charge must be specified for a given compound.
- The proportion of metal cation to nonmetal anion helps us determine the charge on the metal ion.

Type II Ionic Compounds

- Iron, for instance, forms a 2+ cation in some of its compounds and a 3+ cation in others.
- Metals of this type are often *transition metals*.
  - FeS: Here, iron is +2 cation (Fe^{2+}).
  - Fe_2S_3: Here, iron is +3 cation (Fe^{3+}).
  - Cu_2O: Here, copper is +1 cation (Cu^{+}).
  - CuO: Here, copper is +2 cation (Cu^{2+}).
- Some main group metals, such as Pb, Ti, and Sn, form more than one type of cation.

Naming Type II Binary Ionic Compounds

- The full name of compounds containing metals that form more than one kind of cation have the following form:



- The charge of the metal cation can be determined by inference from the sum of the charges of the nonmetal.

Naming Type II Binary Ionic Compounds

- For these types of metals, the name of the cation is followed by a roman numeral (in parentheses) that indicates the charge of the metal in that particular compound.
  - For example, we distinguish between Fe^{2+} and Fe^{3+} as follows:
    - Fe^{2+} Iron(II)
    - Fe^{3+} Iron(III)

TABLE 3.3 Some Metals That Form Cations with Different Charges

Metal	Ion	Name	Older Name*
Chromium	Cr^{2+}	Chromium(II)	Chromous
	Cr^{3+}	Chromium(III)	Chromic
Iron	Fe^{2+}	Iron(II)	Ferrous
	Fe^{3+}	Iron(III)	Ferric
Cobalt	Co^{2+}	Cobalt(II)	Cobaltous
	Co^{3+}	Cobalt(III)	Cobaltic
Copper	Cu^{+}	Copper(I)	Cuprous
	Cu^{2+}	Copper(II)	Cupric
Tin	Sn^{2+}	Tin(II)	Stannous
	Sn^{4+}	Tin(IV)	Stannic
Mercury	Hg_2^{2+}	Mercury(I)	Mercurous
	Hg^{2+}	Mercury(II)	Mercuric
Lead	Pb^{2+}	Lead(II)	Plumbous
	Pb^{4+}	Lead(IV)	Plumbic

*An older naming system substitutes the names found in this column for the name of the metal and its charge. Under this system, chromium(II) oxide is named chromous oxide. Additionally, the suffix *-ous* indicates the ion with the lesser charge, and *-ic* indicates the ion with the greater charge. We will *not* use the older system in this text.

Example: Type II Binary Ionic Compounds

- To name $CrBr_3$, determine the charge on the chromium.
  - Total charge on cation + total anion charge = 0.
  - Cr charge + 3(Br^- charge) = 0.
  - Since each Br has a –1 charge, then:
    - Cr charge + 3(–1) = 0.
    - Cr charge –3 = 0.
    - Cr = +3.
  - Hence, the cation Cr^{3+} is called chromium(III), while Br^- is called bromide.
- Therefore, $CrBr_3$ is **chromium(III) bromide.**

Naming Ionic Compounds Containing Polyatomic Ions

- We name ionic compounds that contain a polyatomic ion in the same way as other ionic compounds, except that we use the name of the polyatomic ion whenever it occurs.
- For example, $NaNO_2$ is named according to its cation, Na^+, *sodium*, and its polyatomic anion, NO_2^-, *nitrite*.
- Hence, $NaNO_2$ is *sodium nitrite*.

TABLE 3.4 Some Common Polyatomic Ions

Name	Formula	Name	Formula
Acetate	$C_2H_3O_2^-$	Hypochlorite	ClO^-
Carbonate	CO_3^{2-}	Chlorite	ClO_2^-
Hydrogen carbonate (or bicarbonate)	HCO_3^-	Chlorate	ClO_3^-
Hydroxide	OH^-	Perchlorate	ClO_4^-
Nitrite	NO_2^-	Permanganate	MnO_4^-
Nitrate	NO_3^-	Sulfite	SO_3^{2-}
Chromate	CrO_4^{2-}	Hydrogen sulfite (or bisulfite)	HSO_3^-
Dichromate	$Cr_2O_7^{2-}$	Sulfate	SO_4^{2-}
Phosphate	PO_4^{3-}	Hydrogen sulfate (or bisulfate)	HSO_4^-
Hydrogen phosphate	HPO_4^{2-}	Cyanide	CN^-
Dihydrogen phosphate	$H_2PO_4^-$	Peroxide	O_2^{2-}
Ammonium	NH_4^+		

Oxyanions

- Most polyatomic ions are **oxyanions**, anions containing oxygen and another element.
- Notice that when a series of oxyanions contains different numbers of oxygen atoms, they are named according to the number of oxygen atoms in the ion.
- If there are two ions in the series,
  - the one with more oxygen atoms has the ending *-ate*, and
  - the one with fewer has the ending *-ite*.
- For example,
  - NO_3^- is ***nitrate***.
  - SO_4^{2-} is ***sulfate***.
  - NO_2^- is ***nitrite***.
  - SO_3^{2-} is ***sulfite***.

Oxyanions

- If there are more than two ions in the series, then the prefixes *hypo*-, meaning *less than*, and *per*-, meaning *more than*, are used.

ClO^-	*hypo*chlor*ite*	BrO^-	*hypo*brom*ite*
ClO_2^-	chlor*ite*	BrO_2^-	brom*ite*
ClO_3^-	chlor*ate*	BrO_3^-	brom*ate*
ClO_4^-	*per*chlor*ate*	BrO_4^-	*per*brom*ate*

Hydrated Ionic Compounds

- **Hydrates** are ionic compounds containing a specific number of water molecules associated with each formula unit.
  - For example, the formula for epsom salts is $MgSO_4 \cdot 7H_2O$.
  - Its systematic name is magnesium sulfate heptahydrate.
  - $CoCl_2 \cdot 6H_2O$ is cobalt(II) chloride hexahydrate.

Common Hydrate Prefixes

- Common hydrate prefixes

hemi = ½	tri = 3	hexa = 6
mono = 1	tetra = 4	hepta = 7
di = 2	penta = 5	octa = 8

- Other common hydrated ionic compounds and their names are as follows:
  - $CaSO_4 \cdot ½ H_2O$ is called calcium sulfate hemihydrate.
  - $BaCl_2 \cdot 6H_2O$ is called barium chloride hexahydrate.
  - $CuSO_4 \cdot 6H_2O$ is called copper sulfate hexahydrate.

Molecular Compounds: Formulas and Names

- The formula for a molecular compound *cannot* readily be determined from its constituent elements because the same combination of elements may form many different molecular compounds, each with a different formula.
  - Nitrogen and oxygen form all of the following unique molecular compounds: NO, NO_2, N_2O, N_2O_3, N_2O_4, and N_2O_5.

Molecular Compounds: Formulas and Names

- *Molecular compounds are composed of two or more nonmetals*.
- Generally, write the name of the element with the smallest group number first.
- If the two elements lie in the same group, then write the element with the greatest row number first.
  - The prefixes given to each element indicate the number of atoms present.

Binary Molecular Compounds

- These prefixes are the same as those used in naming hydrates:

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

- If there is only one atom of the *first element* in the formula, the prefix *mono*- is normally omitted.

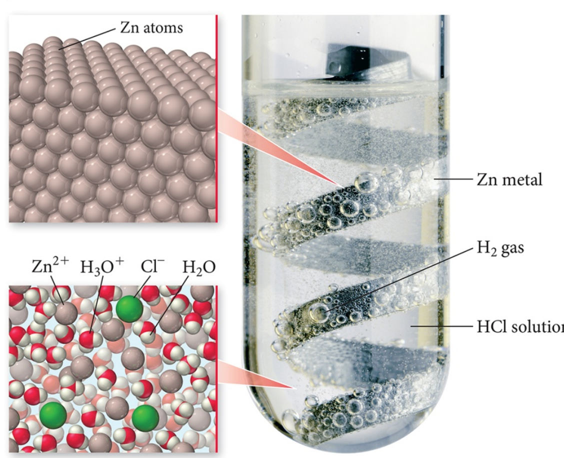
Acids

- Acids are molecular compounds that release hydrogen ions (H^+) when dissolved in water.
- Acids are composed of hydrogen, usually written first in their formulas, and one or more nonmetals, written second.

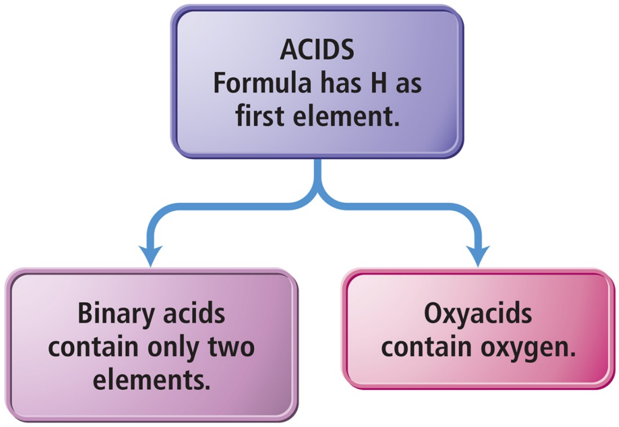
Acids

- Sour taste
- Dissolve many metals
  - such as Zn, Fe, and Mg; but not Au, Ag, or Pt
- Formulas generally start with H,
  - e.g., HCl, H_2SO_4
- HCl is a molecular compound that, when dissolved in water, forms $H^+_{(aq)}$ and $Cl^-_{(aq)}$ ions, where *aqueous* (*aq*) means *dissolved in water*.

Acids Dissolve Many Metals



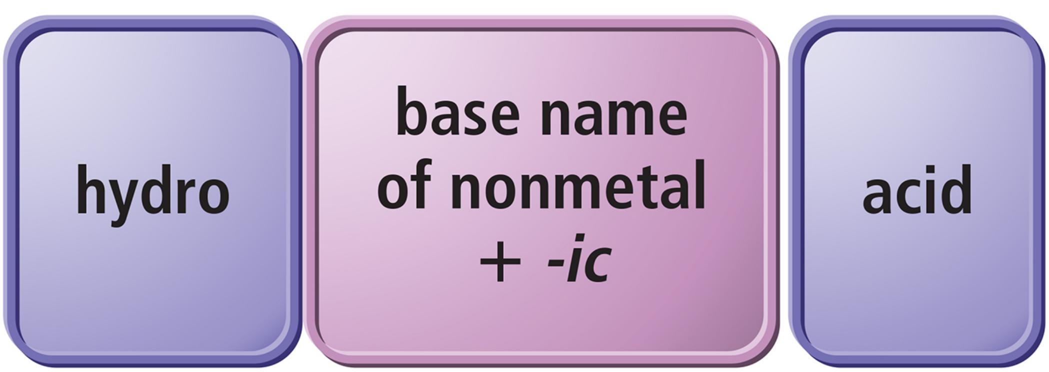
Acids



- Binary acids have H^+ cation and nonmetal anion.
- Oxyacids have H^+ cation and polyatomic anion.

Naming Binary Acids

- Write a ***hydro-*** prefix.
- Follow with the nonmetal base name.
- Add –*ic*.
- Write the word **acid** at the end of the name.



Naming Oxyacids

- If the polyatomic ion name ends in –*ate,* change ending to –*ic*.
- If the polyatomic ion name ends in –*ite*, change ending to –*ous*.
- Write word **acid** at the end of all names.

oxyanions ending with -*ate* | base name of oxyanion + -*ic* | acid

oxyanions ending with -*ite* | base name of oxyanion + -*ous* | acid

Practice: Name the Acid

1. H_2S

2. $HClO_3$

3. $HC_2H_3O_2$

Practice: Name the Acid

1. H_2S **hydrosulfuric acid**

2. $HClO_3$ **chloric acid**

3. $HC_2H_3O_2$ **acetic acid**

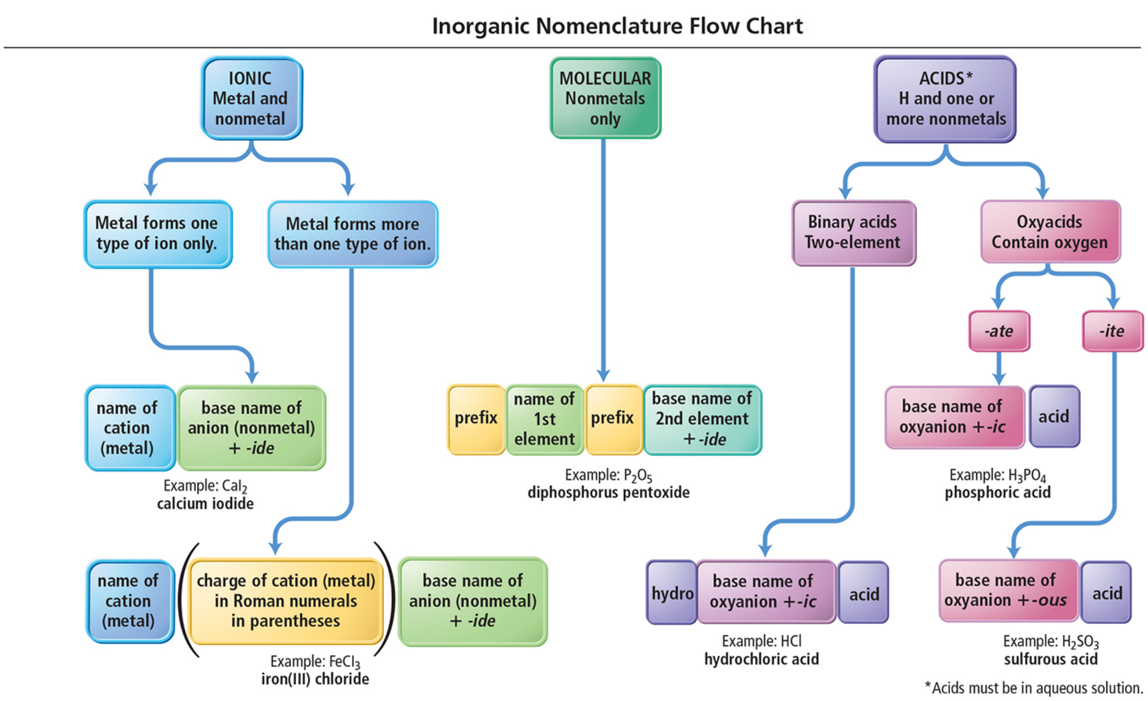
Writing Formulas for Acids

- When the name ends in **acid**, the formula starts with **H** followed by an anion.
- Write the formula as if it is ionic, even though it is molecular.
- *Hydro-* prefix means it is binary acid; no prefix means it is an oxyacid.
- For an oxyacid,
  - if the ending is ***–ic***, the polyatomic ion ends in *–ate*.
  - if the ending is ***–ous***, the polyatomic ion ends in *–ous*.

Acid Rain

- Certain pollutants, such as NO, NO_2, SO_2, and SO_3, form acids when mixed with water, resulting in acidic rainwater.
- Acid rain can fall or flow into lakes and streams, making these bodies of water more acidic.

Inorganic Nomenclature Flow Chart



Formula Mass

- The mass of an individual molecule or formula unit
  - also known as molecular mass or molecular weight
- Sum of the masses of the atoms in a single molecule or formula unit
  - whole = sum of the parts!

Mass of 1 molecule of H_2O =

(2 atoms H)(1.01 amu/H atom) + (1 atom O)(16.00 amu/atom O)
= 18.02 amu

$$\text{Formula mass} = \left(\boxed{\begin{array}{c}\text{Number of atoms}\\ \text{of 1st element in}\\ \text{chemical formula}\end{array}} \times \boxed{\begin{array}{c}\text{Atomic mass}\\ \text{of}\\ \text{1st element}\end{array}}\right) + \left(\boxed{\begin{array}{c}\text{Number of atoms}\\ \text{of 2nd element in}\\ \text{chemical formula}\end{array}} \times \boxed{\begin{array}{c}\text{Atomic mass}\\ \text{of}\\ \text{2nd element}\end{array}}\right) + \ldots$$

Molar Mass of Compounds

- ***The molar mass of a compound*—the mass, in grams, of 1 mol of its molecules or formula units—is numerically equivalent to its formula mass with units of g/mol.**

Molar Mass of Compounds

- The relative masses of molecules can be calculated from atomic masses:

  formula mass = 1 molecule of H_2O = 2(1.01 amu H) + 16.00 amu O = 18.02 amu

- 1 mole of H_2O contains 2 moles of H and 1 mole of O:

  molar mass = 1 mole H_2O

  = 2mol(1.01 g/1 mol H) + 1mol(16.00 g/1 mol O) = 18.02 g/1 mol H_2O

  So the molar mass of H_2O is 18.02 g/mole.

- Molar mass = formula mass (in g/mole)

Using Molar Mass to Count Molecules by Weighing

- Molar mass in combination with Avogadro's number can be used to determine the number of atoms in a given mass of the element.
  - Use molar mass to convert to the amount in moles. Then use Avogadro's number to convert to number of molecules.

$$\#\ molecule\ A = \frac{gram\ A}{1} * \frac{1\ mole\ A}{molar\ mass\ A} * \frac{6.02\times10^{23}\ molecules\ A}{1\ mole\ A}$$

Composition of Compounds

- **A chemical formula, in combination with the molar masses of its constituent elements, indicates the relative quantities of each element in a compound.**

Composition of Compounds

- Percentage by mass of each element in a compound. Can be determined from
  1. the formula of the compound and
  2. the experimental mass analysis of the compound.
- The percentages may not always total to 100% due to rounding.

$$\text{Mass percent of element X} = \frac{\text{mass of element X in 1 mol of compound}}{\text{mass of 1 mol of compound}} \times 100\%$$

Conversion Factors from Chemical Formulas

- Chemical formulas show the relationship between numbers of atoms and molecules.
  - Or moles of atoms and molecules

$$58.64 \text{ g Cl} : 100 \text{ g } CCl_2F_2$$

$$1 \text{ mol } CCl_2F_2 : 2 \text{ mol Cl}$$

- These relationships can be used to determine the amounts of constituent elements and molecules.
  - Like percent composition

Determining a Chemical Formula from Experimental Data

Empirical Formula

- Simplest, whole-number ratio of the atoms or moles of elements in a compound, *not a ratio of masses*
- Can be determined from elemental analysis
  - Percent composition
  - Masses of elements formed when a compound is decomposed, or that react together to form a compound

Finding an Empirical Formula

1. Convert the percentages to grams.
   (a) Assume you start with 100 g of the compound.
   (b) Skip if it is already in grams.
2. Convert grams to moles.
   (a) Use the molar mass of each element.
3. Write a pseudoformula using moles as subscripts.

Finding an Empirical Formula (continued)

4. Divide all by the smallest number of moles.
   (a) If the result is within 0.1 of a whole number, round to the whole number.
5. Multiply all mole ratios by a number to make all whole numbers.
   (a) If ratio is .5, multiply all by 2.
   (b) If ratio is .33 or .67, multiply all by 3.
   (c) If ratio is 0.25 or 0.75, multiply all by 4, etc.
   (d) Skip if ratios are already whole numbers.

Molecular Formulas for Compounds

- The molecular formula is a whole-number multiple of the empirical formula.
- To determine the molecular formula, you need to know the empirical formula and the molar mass of the compound.

Molecular formula = (empirical formula)***n***,

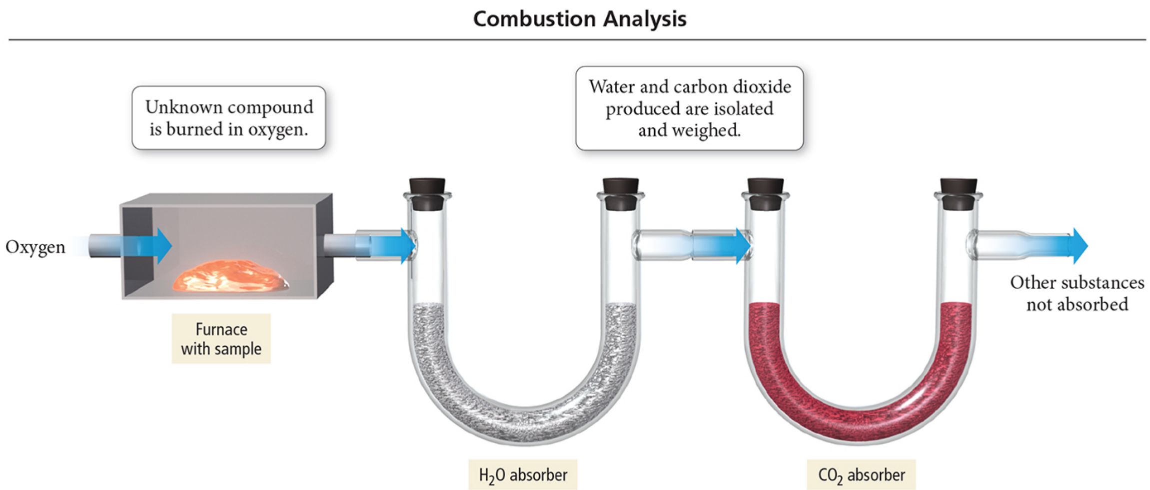
*where **n** is a positive integer.*

$$n = \frac{\text{molar mass}}{\text{empirical formula molar mass}}$$

Combustion Analysis

- A common technique for analyzing compounds is to burn a known mass of compound and weigh the amounts of products.
  - This is generally used for organic compounds containing C, H, and O.
- By knowing the mass of the products and composition of constituent element in the product, the original amount of constituent element can be determined.
  - All the original C forms CO_2, the original H forms H_2O, and the original mass of O is found by subtraction.
- Once the masses of all the constituent elements in the original compound have been determined, the empirical formula can be found.

Combustion Analysis



Chemical Reactions

- Reactions involve chemical changes in matter resulting in new substances.
- Reactions involve rearrangement and exchange of atoms to produce new molecules.
  - Elements are not transmuted during a reaction.

$$\textbf{Reactants} \longrightarrow \textbf{Products}$$

Chemical Equations

- Shorthand way of describing a reaction
- Provide information about the reaction
  - Formulas of reactants and products
  - States of reactants and products
  - Relative numbers of reactant and product molecules that are required can be used to determine weights of reactants used and products that can be made

TABLE 3.5 States of Reactants and Products in Chemical Equations

Abbreviation	State
(*g*)	Gas
(*l*)	Liquid
(*s*)	Solid
(*aq*)	Aqueous (water solution)

Combustion of Methane

- Methane gas burns to produce carbon dioxide gas and gaseous water.
  - Whenever something burns it combines with $O_2(g)$.

$$CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

- If you look closely, you should immediately spot a problem.

$$CH_4(g) + \underbrace{O_2(g)}_{\text{2 O atoms}} \longrightarrow CO_2(g) + H_2O(g)$$

2 O atoms + 1 O atom = **3 O atoms**

Combustion of Methane

- Notice also that the left side has four hydrogen atoms while the right side has only two.

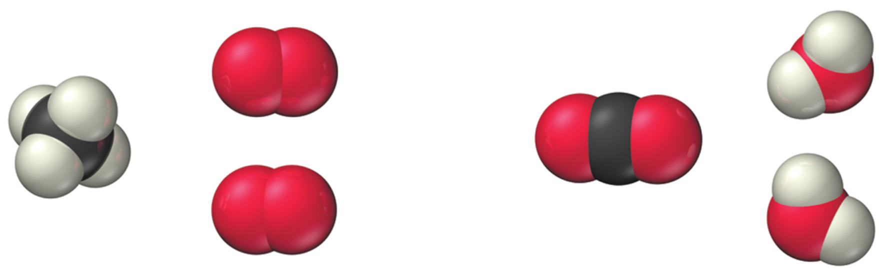
$$\underbrace{CH_4(g)}_{\text{4 H atoms}} + O_2(g) \longrightarrow CO_2(g) + \underbrace{H_2O(g)}_{\text{2 H atoms}}$$

- To correct these problems, we must **balance** the equation by changing the coefficients, not the subscripts.

Combustion of Methane: Balanced

- To show that the reaction obeys the Law of Conservation of Mass, the equation must be **balanced**.
  - We adjust the numbers of molecules so there are equal numbers of atoms of each element on both sides of the arrow.

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$

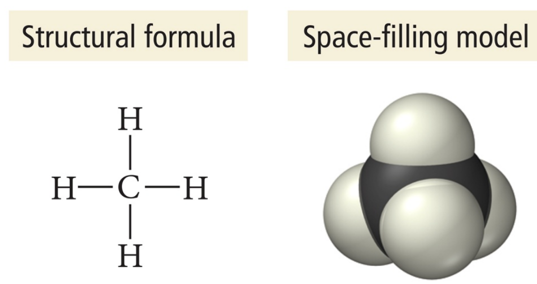


Organic Compounds

- Early chemists divided compounds into two types: organic and inorganic.
- Compounds originating from living things were called organic; compounds originating from the earth were called inorganic.
- Organic compounds were easily decomposed and could not be made in the lab.
- Inorganic compounds were very difficult to decompose but could be synthesized.

Modern Organic Compounds

- Today, organic compounds are commonly made in the lab and we find them all around us.
- Organic compounds are mainly made of C and H, sometimes with O, N, P, S, and trace amounts of other elements.
- The key element of organic chemistry is **carbon**.

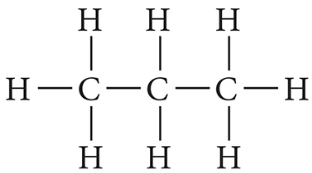


Methane, CH_4

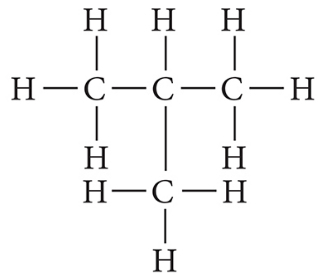
Carbon Bonding

- Carbon atoms bond almost exclusively covalently.
  - Compounds with ionic bonding C are generally inorganic.
- When C bonds, it forms four covalent bonds, including single, double, and triple bonds.
- Carbon is unique in that it can bond with itself and form limitless chains of C atoms, both straight and branched, as well as rings of C atoms.

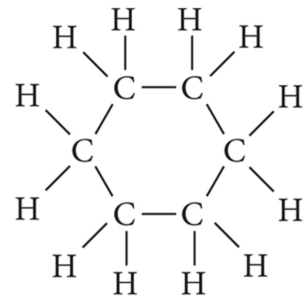
Carbon Bonding



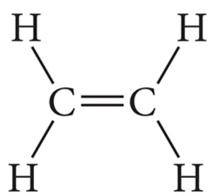
Propane (C_3H_8)



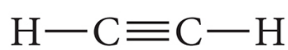
Isobutane (C_4H_{10})



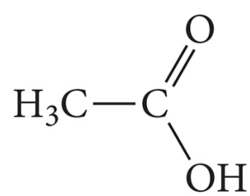
Cyclohexane (C_6H_{12})



Ethene (C_2H_4)



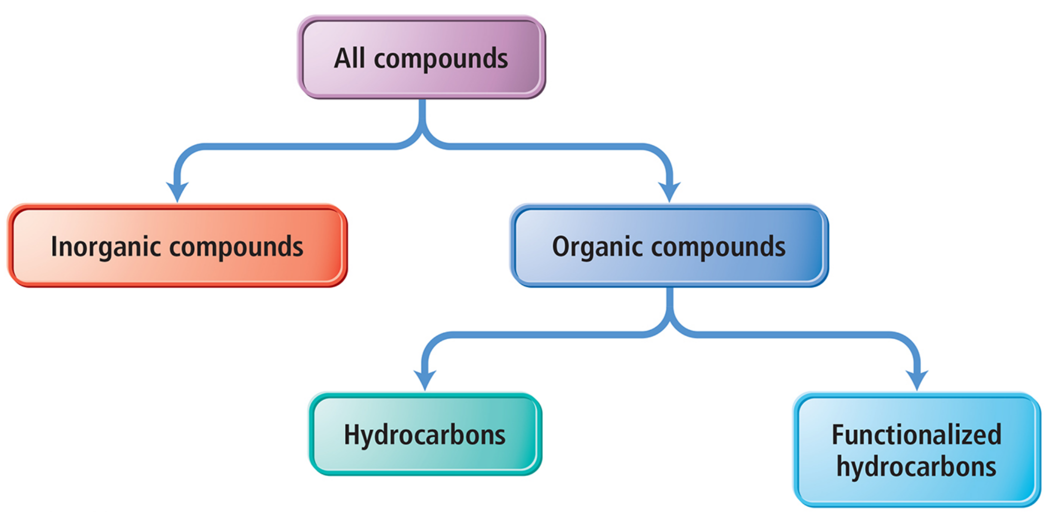
Ethyne (C_2H_2)



Acetic acid (CH_3COOH)

Hydrocarbons

- Organic compounds can be categorized into two types: hydrocarbons and functionalized hydrocarbons.



Hydrocarbons

- **Hydrocarbons** are organic compounds that contain only carbon and hydrogen.
- Hydrocarbons compose common fuels such as
  - oil,
  - gasoline,
  - liquid propane gas,
  - and natural gas.

Naming of Hydrocarbons

- Hydrocarbons containing only single bonds are called **alkanes**.
- Those containing double or triple bonds are **alkenes** and **alkynes**, respectively.
- Hydrocarbons consist of a base name and a suffix.
  - alkane (-*ane*)
  - alkene (-*ene*)
  - alkyne (-*yne*)

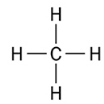
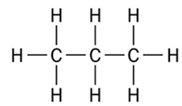
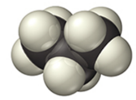
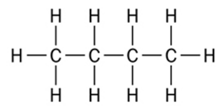
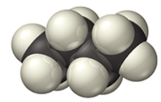
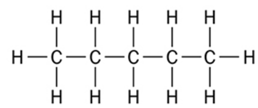
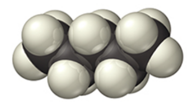
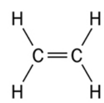
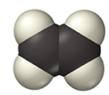
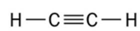
- The base names for a number of hydrocarbons are listed here:

1 meth-	6 hex-
2 eth-	7 hept-
3 prop-	8 oct-
4 but-	9 non-
5 pent-	10 dec-

Base name determined by number of C atoms

Suffix determined by presence of multiple bonds

TABLE 3.6 Common Hydrocarbons

Name	Molecular Formula	Structural Formula	Space-filling Model	Common Uses
Methane	CH_4			Primary component of natural gas
Propane	C_3H_8			LP gas for grills and outdoor stoves
n-Butane*	C_4H_{10}			Common fuel for lighters
n-Pentane*	C_5H_{12}			Component of gasoline
Ethene	C_2H_4			Ripening agent in fruit
Ethyne	C_2H_2			Fuel for welding torches

*The "*n*" in the names of these hydrocarbons stands for "normal," which means straight chain.

Functionalized Hydrocarbons

- The term ***functional group*** derives from the functionality or chemical character that a specific atom or group of atoms imparts to an organic compound.
  - Even a carbon–carbon double or triple bond can justifiably be called a "functional group."
- A group of organic compounds with the same functional group forms a **family.**

Functionalized Hydrocarbons

Hydrocarbon (R) group

CH_3OH

Methanol

OH functional group

Hydrocarbon (R) group

CH_3
|
CH_3CHOH

Isopropanol
(2-propanol)

OH functional group

TABLE 3.7 Families of Organic Compounds

Family	Name Ending	General Formula	Example	Name	Occurrence/Use
Alcohols	-ol	R—OH	CH_3CH_2—OH	Ethanol (ethyl alcohol)	Alcohol in fermented beverages
Ethers	ether	R—O—R′	CH_3CH_2—O—CH_2CH_3	Diethyl ether	Anesthetic; laboratory solvent
Aldehydes	-al	R—C(=O)—H	H_3C—C(=O)—H	Ethanal (acetaldehyde)	Perfumes; flavors
Ketones	-one	R—C(=O)—R′	H_3C—C(=O)—CH_3	Propanone (acetone)	Fingernail polish remover
Carboxylic acids	acid	R—C(=O)—OH	H_3C—C(=O)—OH	Acetic acid	Vinegar
Esters	-ate	R—C(=O)—OR′	H_3C—C(=O)—OCH_3	Methyl acetate	Laboratory solvent
Amines	amine	RNH_2	CH_3CH_2—N(H)—H	Ethyl amine	Smell of rotten fish