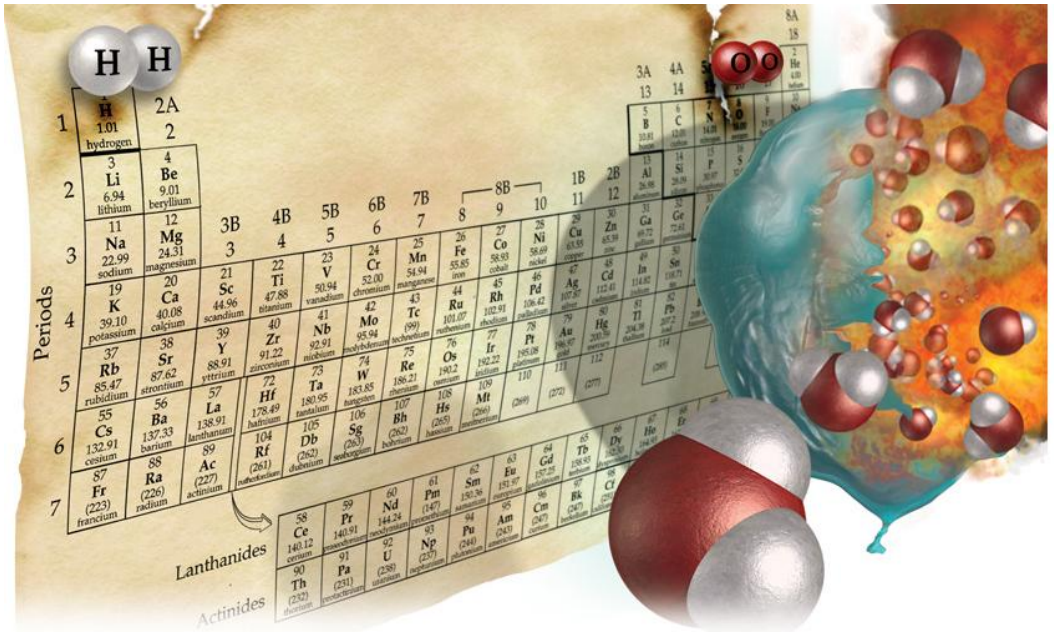
Chemistry: A Molecular Approach, 1st Ed.
Nivaldo Tro

Chapter 3 Molecules, Compounds, and Chemical Equations



Roy Kennedy
Massachusetts Bay Community College
Wellesley Hills, MA

2008, Prentice Hall

Elements and Compounds

- elements combine together to make an almost limitless number of compounds
- the properties of the compound are totally different from the constituent elements

Selected Prope	Selected Properties of Water		Oxygen	
Boiling point, –	Boiling point, 100 °C			
Gas at room ter	Liquid at room temperature		·e	
Explosive	Used to extinguish flame			

Formation of Water from Its Elements

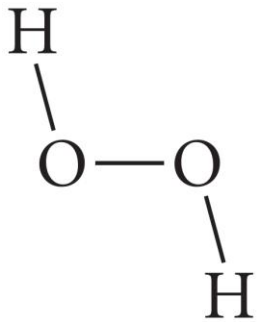
Mixtures and Compounds

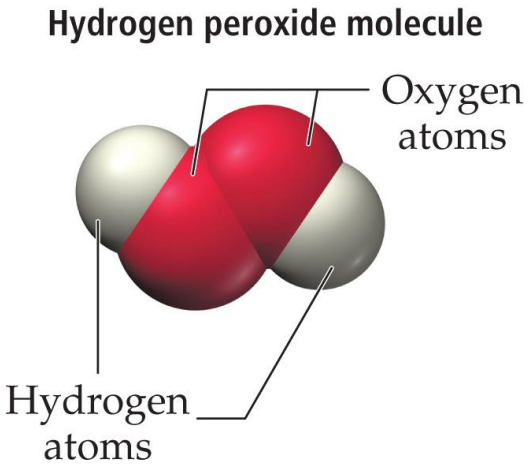
Hydrogen and Oxygen Mixture
Can have any ratio of hydrogen to oxygen.

Water (A Compound)
Water molecules have a fixed ratio of hydrogen (2 atoms) to oxygen (1 atom).

Chemical Bonds

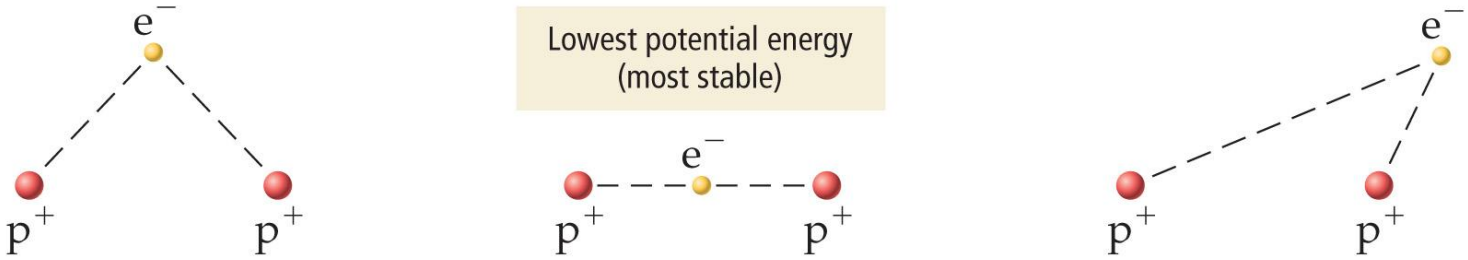
- compounds are made of atoms held together by **chemical bonds**
- bonds are forces of attraction between atoms
- the bonding attraction comes from attractions between protons and electrons



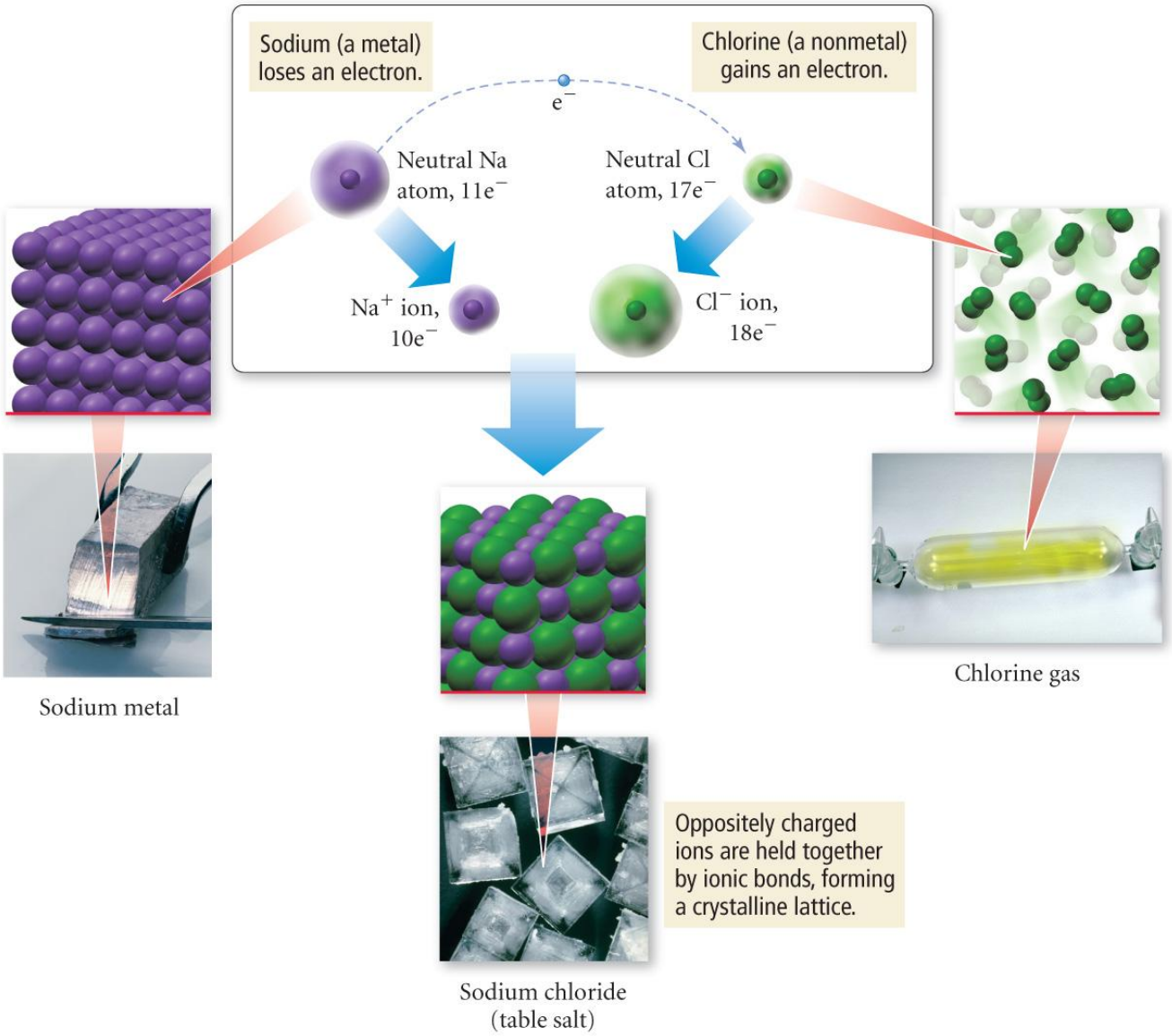


Bond Types

- two general types of bonding between atoms found in compounds, **ionic** and **covalent**
- **ionic bonds** result when electrons have been transferred between atoms, resulting in oppositely charged ions that attract each other
  - ✓ generally found when metal atoms bonded to nonmetal atoms
- **covalent bonds** result when two atoms share some of their electrons
  - ✓ generally found when nonmetal atoms bonded together

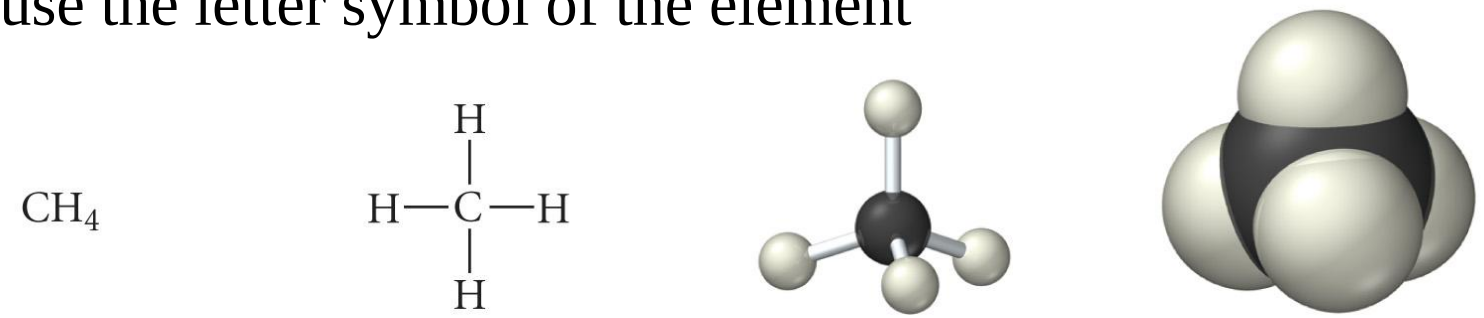


The Formation of an Ionic Compound



Representing Compounds with Chemical Formula

- compounds are generally represented with a **chemical formula**
- the amount of information about the structure of the compound varies with the type of formula
  - ✓ all formula and models convey a limited amount of information – none are perfect representations
- all chemical formulas tell what elements are in the compound
  - ✓ use the letter symbol of the element



Molecular formula | Structural formula | Ball-and-stick model | Space-filling model

Types of Formula
Empirical Formula

- **Empirical Formula** describe the kinds of elements found in the compound and the ratio of their atoms
  - ✓they do not describe how many atoms, the order of attachment, or the shape
  - ✓the formulas for ionic compounds are empirical

Types of Formula
Molecular Formula

- **Molecular Formula** describe the kinds of elements found in the compound and the numbers of their atoms
  - ✓they do not describe the order of attachment, or the shape

Types of Formula
Structural Formula

- **Structural Formula** describe the kinds of elements found in the compound, the numbers of their atoms, order of atom attachment, and the kind of attachment
  - ✓ they do not directly describe the 3-dimensional shape, but an experienced chemist can make a good guess at it
  - ✓ use lines to represent covalent bonds
  - ✓ each line describes the number of electrons shared by the bonded atoms
    - ➢ single line = 2 shared electrons, a single covalent bond
    - ➢ double line = 4 shared electrons, a double covalent bond
    - ➢ triple line = 6 shared electrons, a triple covalent bond

Representing Compounds
Molecular Models

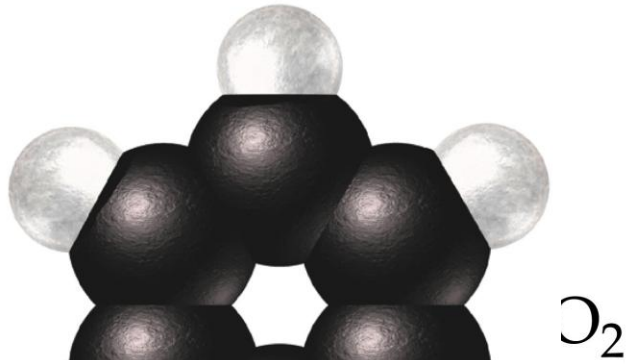
- **Models** show the 3-dimensional structure along with all the other information given in structural formula
- **Ball-and-Stick Models** use balls to represent the atoms and sticks to represent the attachments between them
- **Space-Filling Models** use interconnected spheres to show the electron clouds of atoms connecting together

Chemical Formulas

Hydrogen Peroxide
Molecular Formula = H_2O_2
Empirical Formula = HO

Benzene
Molecular Formula = C_6H_6
Empirical Formula = CH

Glucose
Molecular Formula = $C_6H_{12}O_6$
Empirical Formula = CH_2O



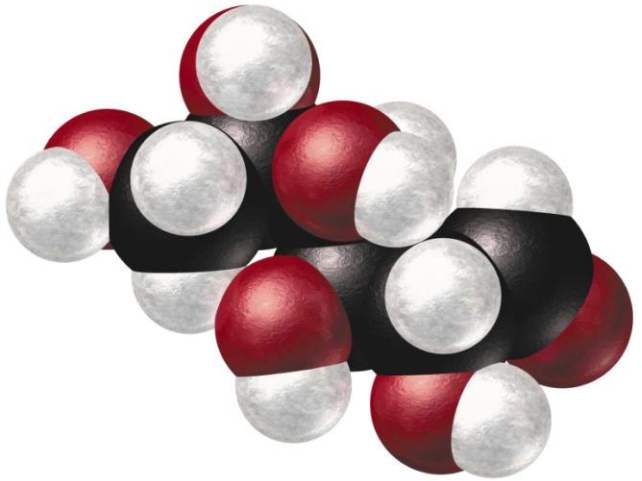
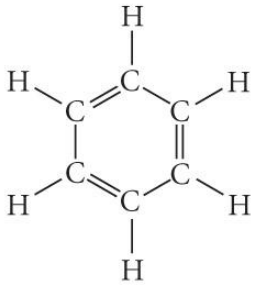
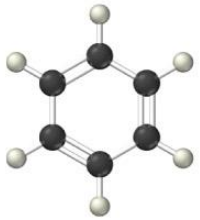
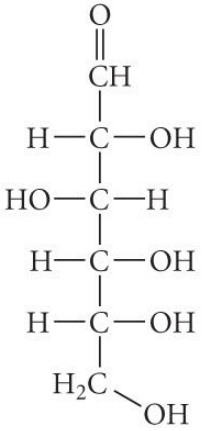
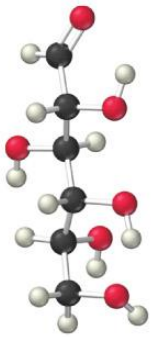
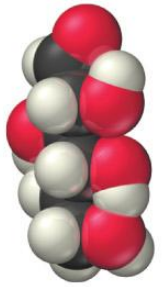
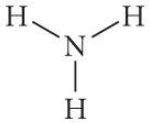
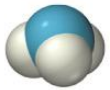
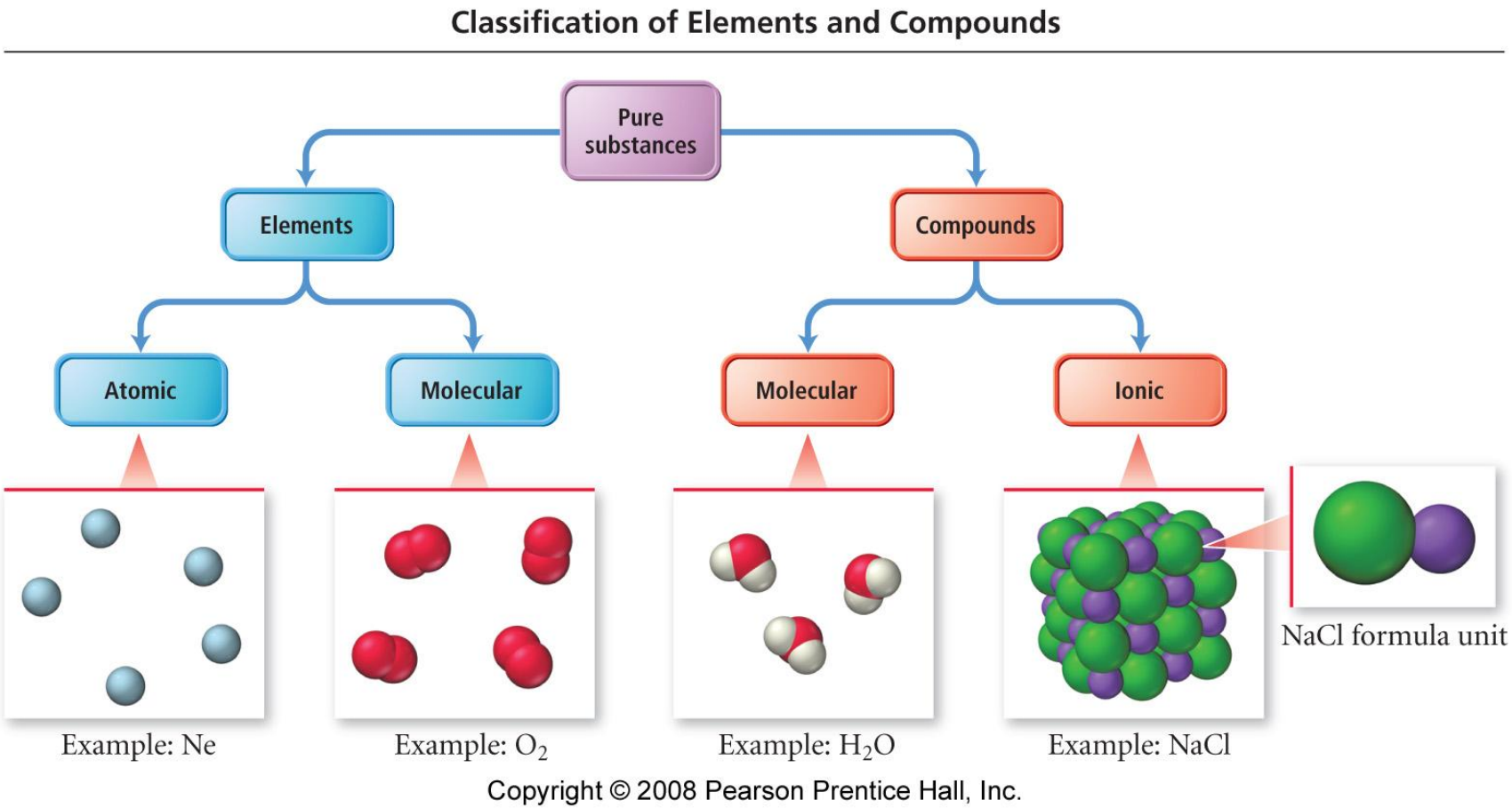


TABLE 3.1 Benzene, Acetylene, Glucose, and Ammonia

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			
Glucose	CH_2O	$C_6H_{12}O_6$			
Ammonia	NH_3	NH_3			

Molecular View of Elements and Compounds

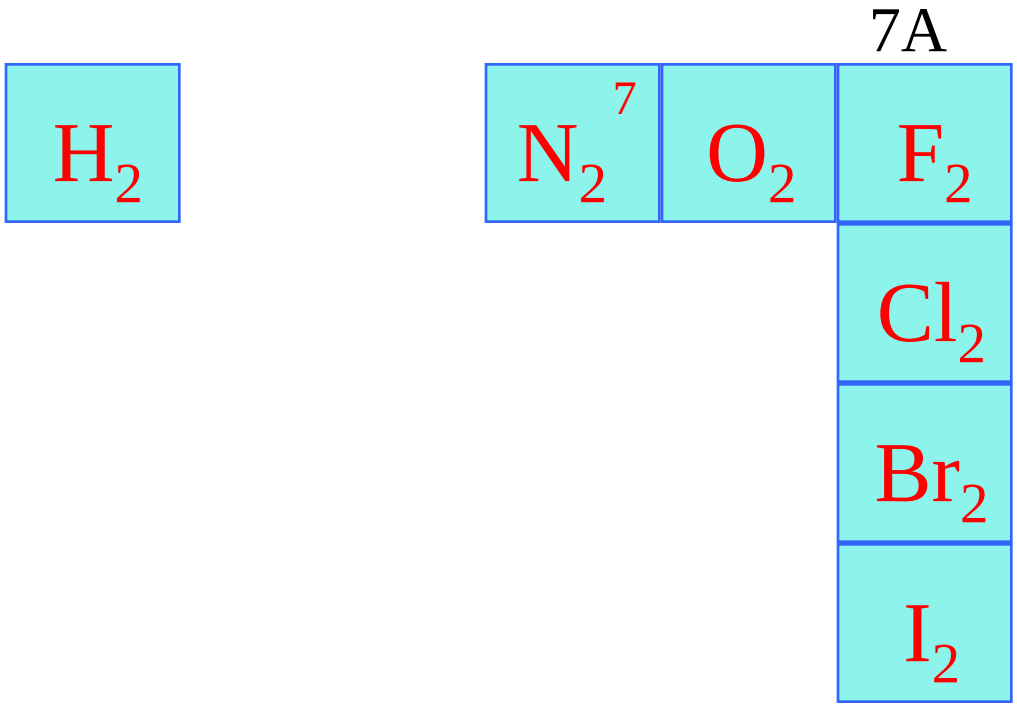


Classifying Materials

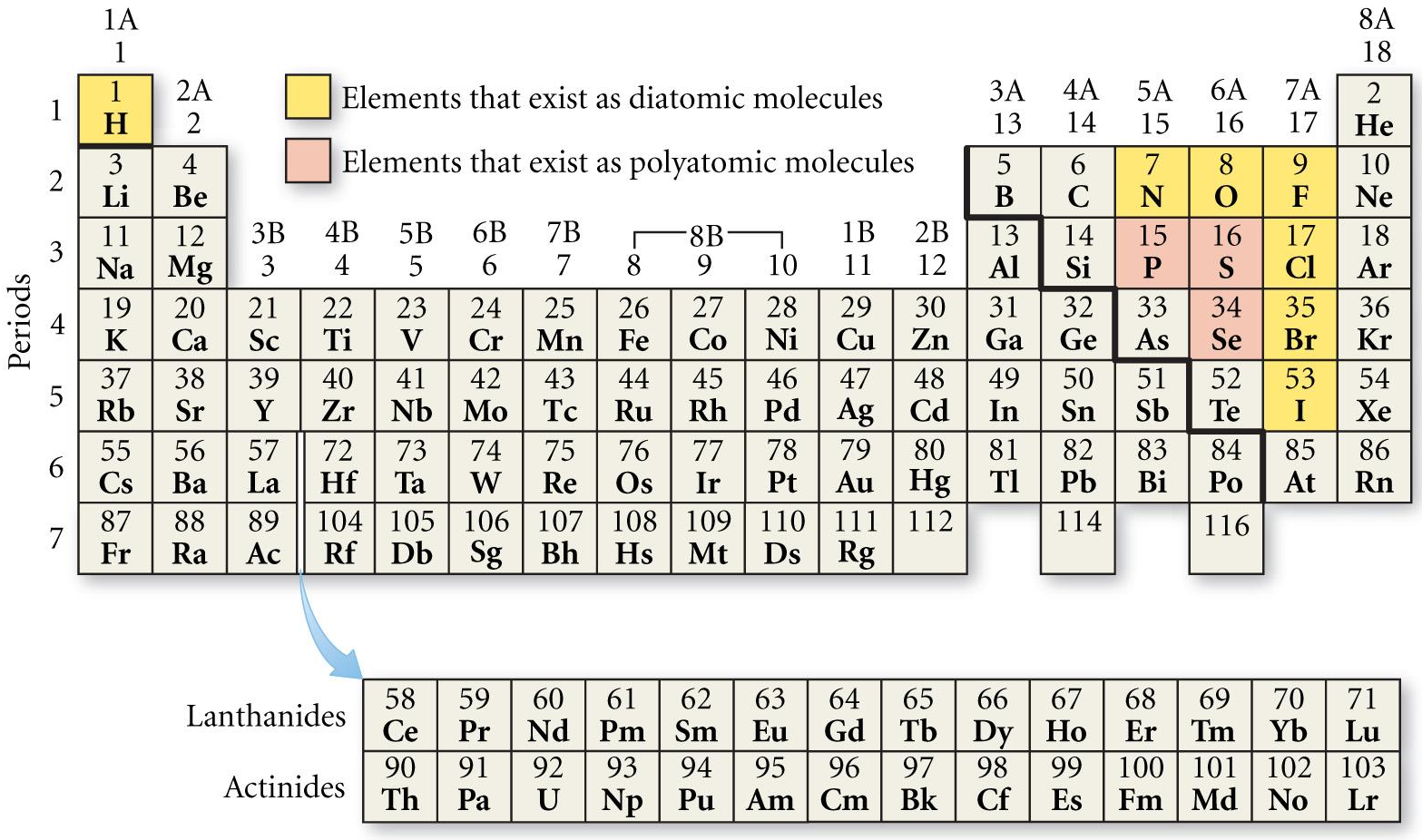
- atomic elements = elements whose particles are single atoms
- molecular elements = elements whose particles are multi-atom molecules
- molecular compounds = compounds whose particles are molecules made of only nonmetals
- ionic compounds = compounds whose particles are cations and anions

Molecular Elements

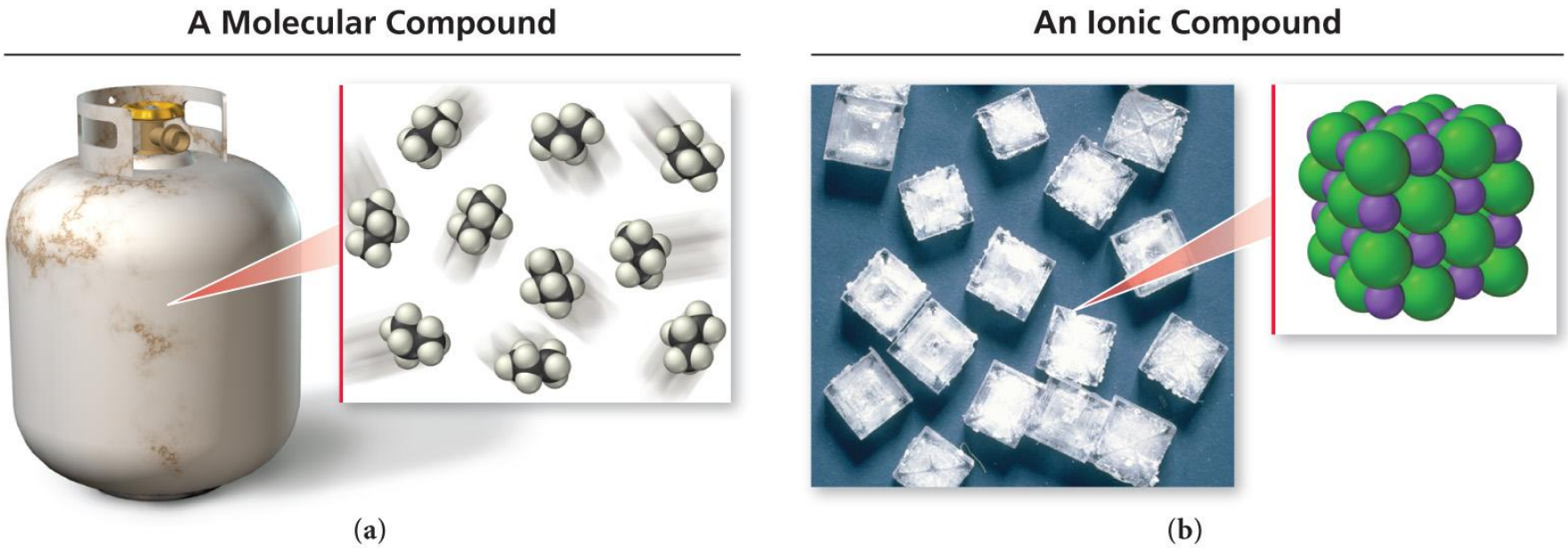
- Certain elements occur as 2 atom molecules
  - ✓ Rule of 7's
- Other elements occur as polyatomic molecules
  - ✓ P_4, S_8, Se_8



Molecular Elements



Ionic vs. Molecular Compounds



Propane – contains individual C_3H_8 molecules

Table salt – contains an array of Na^+ ions and Cl^- ions

Ionic Compounds

- metals + nonmetals
- no individual molecule units, instead have a 3-dimensional array of cations and anions made of **formula units**
- many contain **polyatomic ions**
  - ✓several atoms attached together in one ion

Compounds that Contain Ions

- compounds of metals with nonmetals are made of ions
  - ✓metal atoms form cations, nonmetal atoms for anions
- compound must have no total charge, therefore we must balance the numbers of cations and anions in a compound to get 0 charge
- if Na^+ is combined with S^{2-}, you will need 2 Na^+ ions for every S^{2-} ion to balance the charges, therefore the formula must be Na_2S

Writing Formulas for Ionic Compounds

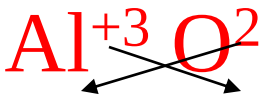
1. Write the symbol for the metal cation and its charge
2. Write the symbol for the nonmetal anion and its charge
3. Charge (without sign) becomes subscript for other ion
4. Reduce subscripts to smallest whole number ratio
5. Check that the sum of the charges of the cation cancels the sum of the anions

Write the formula of a compound made from aluminum ions and oxide ions

1. Write the symbol for the metal cation and its charge
2. Write the symbol for the nonmetal anion and its charge
3. Charge (without sign) becomes subscript for other ion
4. Reduce subscripts to smallest whole number ratio
5. Check that the total charge of the cations cancels the total charge of the anions

Al^{+3} column 3A

O^{2-} column 6A



Al_2O_3

$Al = (2)\cdot(+3) = +6$

$O = (3)\cdot(-2) = -6$

Practice - What are the formulas for compounds made from the following ions?

- potassium ion with a nitride ion
- calcium ion with a bromide ion
- aluminum ion with a sulfide ion

Practice - What are the formulas for compounds made from the following ions?

- K^{+} with N^{3-} K_3N
- Ca^{+2} with Br^{-} $CaBr_2$
- Al^{+3} with S^{2-} Al_2S_3

Formula-to-Name Rules for Ionic Compounds

- made of cation and anion
- some have one or more nicknames that are only learned by experience
  - ✓ NaCl = table salt, $NaHCO_3$ = baking soda
- write systematic name by simply naming the ions
  - ✓ If cation is:
    - ➢ metal with invariant charge = metal name
    - ➢ metal with variable charge = metal name(charge)
    - ➢ polyatomic ion = name of polyatomic ion
  - ✓ If anion is:
    - ➢ nonmetal = stem of nonmetal name + ide
    - ➢ polyatomic ion = name of polyatomic ion

Metal Cations

- Metals with Invariant Charge
  - ✓ metals whose ions can only have one possible charge
    - ➢ Groups $1A^{+1}$ & $2A^{+2}$, Al^{+3}, Ag^{+1}, Zn^{+2}, Sc^{+3}
  - ✓ cation name = metal name

TABLE 3.2 Metals Whose Charge Is Invariant from One Compound to Another

Metal	Ion	Name	Group Number
Li	Li^+	Lithium	1A
Na	Na^+	Sodium	1A
K	K^+	Potassium	1A
Rb	Rb^+	Rubidium	1A
Cs	Cs^+	Cesium	1A
Be	Be^{2+}	Beryllium	2A
Mg	Mg^{2+}	Magnesium	2A
Ca	Ca^{2+}	Calcium	2A
Sr	Sr^{2+}	Strontium	2A
Ba	Ba^{2+}	Barium	2A
Al	Al^{3+}	Aluminum	3A
Zn	Zn^{2+}	Zinc	*
Sc	Sc^{3+}	Scandium	*
Ag**	Ag^+	Silver	*

- Metals with Variable Charges
  - ✓ metals whose ions can have more than onepossible charge
  - ✓ determine charge bycharge on anion
  - ✓ cation name = metal name with Roman numeral charge in parentheses

TABLE 3.4 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

Naming Monatomic Nonmetal Anion

- determine the charge from position on the Periodic Table
- to name anion, change ending on the element name to ***–ide***

4A = -4	5A = -3	6A = -2	7A = -1
C = carbide	N = nitride	O = oxide	F = fluoride
Si = silicide	P = phosphide	S = sulfide	Cl = chloride

Naming Binary Ionic Compounds for Metals with Invariant Charge

- Contain Metal Cation + Nonmetal Anion
- Metal listed first in formula and name

1. name metal cation first, name nonmetal anion second
2. cation name is the metal name
3. nonmetal anion named by changing the ending on the nonmetal name to ***-ide***

Example – Naming Binary Ionic with Invariant Charge Metal

CsF

1. Identify cation and anion

   $Cs = Cs^+$ because it is Group 1A

   $F = F^-$ because it is Group 7A

2. Name the cation

   Cs^+ = cesium

3. Name the anion

   F^- = fluoride

4. Write the cation name first, then the anion name

   cesium fluoride

Name the following compounds

1. KCl

2. $MgBr_2$

3. Al_2S_3

Name the following compounds

1. KCl potassium chloride

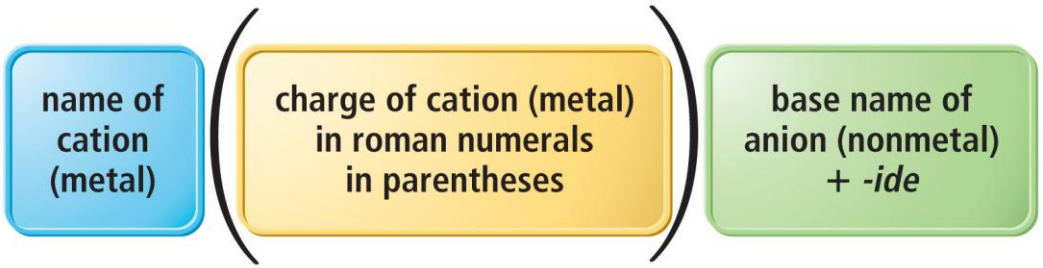
2. $MgBr_2$ magnesium bromide

3. Al_2S_3 aluminum sulfide

Naming Binary Ionic Compounds for Metals with Variable Charge

- Contain Metal Cation + Nonmetal Anion
- Metal listed first in formula and name

1. name metal cation first, name nonmetal anion second
2. metal cation name is the metal name followed by a Roman numeral in parentheses to indicate its charge
   - ✓ determine charge from anion charge
   - ✓ common ions Table 3.4
3. nonmetal anion named by changing the ending on the nonmetal name to ***-ide***



Determining the Charge on a Cation with Variable Charge – Au_2S_3

1. determine the charge on the anion

   Au_2S_3 - the anion is S, since it is in Group 6A, its charge is -2

2. determine the total negative charge

   since there are 3 S in the formula, the total negative charge is -6

3. determine the total positive charge

   since the total negative charge is -6, the total positive charge is +6

4. divide by the number of cations

   since there are 2 Au in the formula and the total positive charge is +6, each Au has a +3 charge

Example – Naming Binary Ionic with Variable Charge Metal

CuF_2

1. Identify cation and anion

   F = F^- because it is Group 7

   Cu = Cu^{2+} to balance the two (-) charges from 2 F^-

2. Name the cation

   Cu^{2+} = copper(II)

3. Name the anion

   F^- = fluoride

4. Write the cation name first, then the anion name

   copper(II) fluoride

Name the following compounds

1. $TiCl_4$

2. $PbBr_2$

3. Fe_2S_3

Name the following compounds

1. $TiCl_4$ titanium(IV) chloride

2. $PbBr_2$ lead(II) bromide

3. Fe_2S_3 iron(III) sulfide

Example – Writing Formula for Binary Ionic Compounds Containing Variable Charge Metal

manganese(IV) sulfide

1. Write the symbol for the cation and its charge — Mn^{+4}
2. Write the symbol for the anion and its charge — S^{2-}
3. Charge (without sign) becomes subscript for other ion — $Mn^{+4}\ S^{2-} \rightarrow Mn_2S_4$
4. Reduce subscripts to smallest whole number ratio — MnS_2
5. Check that the total charge of the cations cancels the total charge of the anions

$$Mn = (1)\cdot(+4) = +4$$
$$S = (2)\cdot(-2) = -4$$

Practice - What are the formulas for compounds made from the following ions?

1. copper(II) ion with a nitride ion

2. iron(III) ion with a bromide ion

Practice - What are the formulas for compounds made from the following ions?

1. Cu^{2+} with N^{3-} Cu_3N_2

2. Fe^{+3} with Br^{-} $FeBr_3$

Compounds Containing Polyatomic Ions

- Polyatomic ions are single ions that contain more than one atom
- Often identified by (ion) in formula
- Name and charge of polyatomic ion do not change
- Name any ionic compound by naming cation first and then anion

Some Common Polyatomic Ions

Name	Formula
acetate	$C_2H_3O_2^-$
carbonate	CO_3^{2-}
hydrogen carbonate (aka bicarbonate)	HCO_3^-
hydroxide	OH^-
nitrate	NO_3^-
nitrite	NO_2^-
chromate	CrO_4^{2-}
dichromate	$Cr_2O_7^{2-}$
ammonium	NH_4^+

Name	Formula
hypochlorite	ClO^-
chlorite	ClO_2^-
chlorate	ClO_3^-
perchlorate	ClO_4^-
sulfate	SO_4^{2-}
sulfite	SO_3^{2-}
hydrogen sulfate (aka bisulfate)	HSO_4^-
hydrogen sulfite (aka bisulfite)	HSO_3^-

Patterns for Polyatomic Ions

1. elements in the same column form similar polyatomic ions
   - ✓ same number of O's and same charge

     ClO_3^- = chlorate ∴ BrO_3^- = bromate
2. if the polyatomic ion starts with H, add ***hydrogen-*** prefix before name and add 1 to the charge

   CO_3^{2-} = carbonate ∴ HCO_3^{-1} = hydrogen carbonate

Periodic Pattern of Polyatomic Ions

-ate groups

3A	4A	5A	6A	7A
BO_3^{-3}	**CO_3^{-2}**	**NO_3^{-1}**		
	SiO_3^{-2}	**PO_4^{-3}**	**SO_4^{-2}**	**ClO_3^{-1}**
		AsO_4^{-3}	SeO_4^{-2}	BrO_3^{-1}
			TeO_4^{-2}	IO_3^{-1}

Patterns for Polyatomic Ions

- -ate ion
  - ✓chlorate = ClO_3^{-1}
- -ate ion + 1 O ⇒ same charge, ***per-*** prefix
  - ✓perchlorate = ClO_4^{-1}
- -ate ion – 1 O ⇒ same charge, ***-ite*** suffix
  - ✓chlorite = ClO_2^{-1}
- -ate ion – 2 O ⇒ same charge, ***hypo-*** prefix, ***-ite*** suffix
  - ✓hypochlorite = ClO^{-1}

Example – Naming Ionic Compounds Containing a Polyatomic Ion

Na_2SO_4

1. Identify the ions

   Na = Na^+ because in Group 1A

   SO_4 = SO_4^{2-} a polyatomic ion

2. Name the cation

   Na^+ = sodium, metal with invariant charge

3. Name the anion

   SO_4^{2-} = sulfate

4. Write the name of the cation followed by the name of the anion

   sodium sulfate

Example – Naming Ionic Compounds Containing a Polyatomic Ion

$Fe(NO_3)_3$

1. Identify the ions

   $NO_3 = NO_3^-$ a polyatomic ion

   $Fe = Fe^{+3}$ to balance the charge of the 3 NO_3^{-1}

2. Name the cation

   Fe^{+3} = iron(III), metal with variable charge

3. Name the anion

   NO_3^- = nitrate

4. Write the name of the cation followed by the name of the anion

   iron(III) nitrate

Name the following

1. NH_4Cl
2. $Ca(C_2H_3O_2)_2$
3. $Cu(NO_3)_2$

Name the following

1. NH_4Cl ammonium chloride

2. $Ca(C_2H_3O_2)_2$ calcium acetate

3. $Cu(NO_3)_2$ copper(II) nitrate

Example – Writing Formula for Ionic Compounds Containing Polyatomic Ion

Iron(III) phosphate

1. Write the symbol for the cation and its charge — Fe^{+3}
2. Write the symbol for the anion and its charge — PO_4^{3-}
3. Charge (without sign) becomes subscript for other ion — $Fe^{+3}\ PO_4^{3-} \rightarrow Fe_3(PO_4)_3$
4. Reduce subscripts to smallest whole number ratio — $FePO_4$
5. Check that the total charge of the cations cancels the total charge of the anions

$$Fe = (1)\cdot(+3) = +3$$
$$PO_4 = (1)\cdot(-3) = -3$$

Practice - What are the formulas for compounds made from the following ions?

1. aluminum ion with a sulfate ion

2. chromium(II) with hydrogen carbonate

Practice - What are the formulas for compounds made from the following ions?

1. Al^{+3} with SO_4^{2-} $Al_2(SO_4)_3$

2. Cr^{+2} with HCO_3^- $Cr(HCO_3)_2$

Hydrates

- hydrates are ionic compounds containing a specific number of waters for each formula unit
- water of hydration often "driven off" by heating
- in formula, attached waters follow ·
  - ✓ $CoCl_2 \cdot 6H_2O$
- in name attached waters indicated by suffix -*hydrate* after name of ionic compound
  - ✓ $CoCl_2 \cdot 6H_2O$ = cobalt(II) chloride hexahydrate
  - ✓ $CaSO_4 \cdot ½H_2O$ = calcium sulfate hemihydrate

Hydrate
$CoCl_2 \cdot 6H_2O$

Anhydrous
$CoCl_2$

Prefix	No. of Waters
hemi	½
mono	1
di	2
tri	3
tetra	4
penta	5
hexa	6
hepta	7
octa	8

Practice

1. What is the formula of magnesium sulfate heptahydrate?

2. What is the name of $NiCl_2 \cdot 6H_2O$?

Practice

1. What is the formula of magnesium sulfate heptahydrate? $MgSO_4 \cdot 7H_2O$

2. What is the name of $NiCl_2 \cdot 6H_2O$?

   nickel(II) chloride hexahydrate

Writing Names of Binary Molecular Compounds of 2 Nonmetals

1. Write name of first element in formula
   - ✓ element furthest left and down on the Periodic Table
   - ✓ use the full name of the element
2. Writes name the second element in the formula with an **-ide** suffix
   - ✓ as if it were an anion, *however, remember these compounds do not contain ions*!
3. Use a prefix in front of each name to indicate the number of atoms
   a) Never use the prefix ***mono-*** on the first element

Subscript - Prefixes

- 1 = mono-
  - ✓ not used on first nonmetal
- 2 = di-
- 3 = tri-
- 4 = tetra-
- 5 = penta-
- 6 = hexa-
- 7 = hepta-
- 8 = octa-
- 9 = nona-
- 10 = deca-

- drop last “a” if name begins with vowel

Example – Naming Binary Molecular

BF_3

1. Name the first element

   boron

2. Name the second element with an ***–ide***

   fluorine ⇒ fluoride

3. Add a prefix to each name to indicate the subscript

   monoboron, trifluoride

4. Write the first element with prefix, then the second element with prefix

   ✓ Drop prefix ***mono*** from first element

   boron trifluoride

Name the following

1. NO_2

2. PCl_5

3. I_2F_7

Name the following

1. NO_2 nitrogen dioxide

2. PCl_5 phosphorus pentachloride

3. I_2F_7 diiodine heptafluoride

Example – Binary Molecular dinitrogen pentoxide

- Identify the symbols of the elements

nitrogen = N

oxide = oxygen = O

- Write the formula using prefix number for subscript

di = 2, penta = 5

N_2O_5

Write formulas for the following

1. dinitrogen tetroxide

2. sulfur hexafluoride

3. diarsenic trisulfide

Write formulas for the following

1. dinitrogen tetroxide N_2O_4

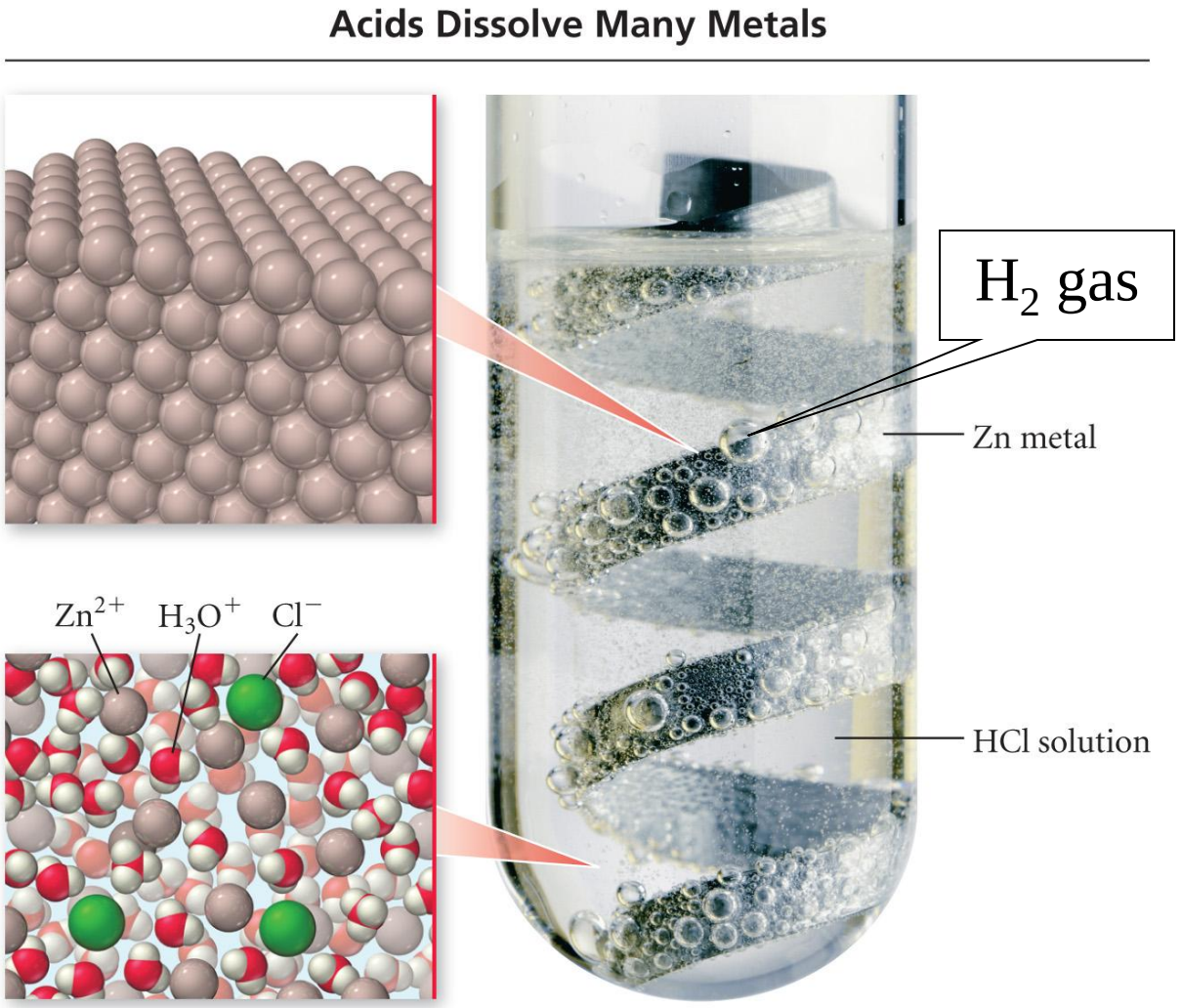
2. sulfur hexafluoride SF_6

3. diarsenic trisulfide As_2S_3

Acids

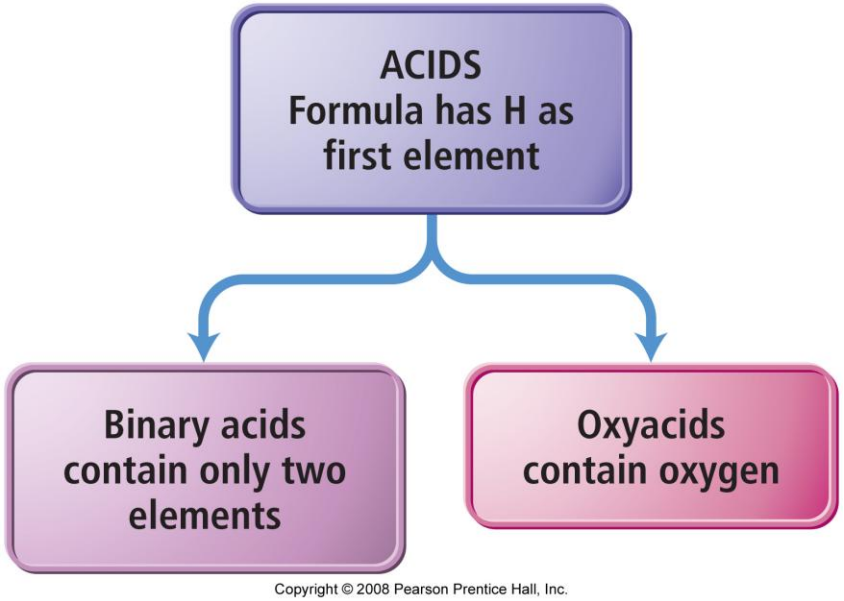
- acids are molecular compounds that form H^+ when dissolved in water
  - ✓to indicate the compound is dissolved in water (*aq*) is written after the formula
    - ➢not named as acid if not dissolved in water
- sour taste
- dissolve many metals
  - ✓like Zn, Fe, Mg; but not Au, Ag, Pt
- formula generally starts with H
  - ✓e.g., HCl, H_2SO_4

Reaction of Acids with Metals



Acids

- Contain H^{+1} cation and anion
  - ✓in aqueous solution
- Binary acids have H^{+1} cation and nonmetal anion
- Oxyacids have H^{+1} cation and polyatomic anion



Naming Binary Acids

- write a **hydro** prefix
- follow with the nonmetal name
- change ending on nonmetal name to *–ic*
- write the word **acid** at the end of the name

Example - Naming Binary Acids – HCl(*aq*)

1. Identify the anion

   Cl = Cl^-, chloride because Group 7A

2. Name the anion with an –***ic*** suffix

   Cl^- = chloride ⇒ chloric

3. Add a ***hydro-*** prefix to the anion name

   hydrochloric

4. Add the word ***acid*** to the end

   hydrochloric acid

Naming Oxyacids

- if polyatomic ion name ends in *–ate*, then change ending to *–ic* suffix
- if polyatomic ion name ends in *–ite*, then change ending to *–ous* suffix
- write word **acid** at end of all names

Example – Naming Oxyacids

$H_2SO_4(aq)$

1. Identify the anion

   $SO_4 = SO_4^{2-}$ = sulfate

2. If the anion has –***ate*** suffix, change it to –***ic***. If the anion has –***ite*** suffix, change it to ***-ous***

   SO_4^{2-} = sulfate $\Rightarrow$ sulfuric

3. Write the name of the anion followed by the word ***acid***

   sulfuric acid

   (kind of an exception, to make it sound nicer!)

Example – Naming Oxyacids

$H_2SO_3(aq)$

1. Identify the anion

   $SO_3 = SO_3^{2-}$ = sulfite

2. If the anion has –***ate*** suffix, change it to –***ic***. If the anion has –***ite*** suffix, change it to ***-ous***

   SO_3^{2-} = sulfite $\Rightarrow$ sulfurous

3. Write the name of the anion followed by the word ***acid***

   sulfurous acid

Name the following

1. H_2S
2. $HClO_3$
3. HNO_2

Name the following

1. H_2S hydrosulfuric acid

2. $HClO_3$ chloric acid

3. HNO_2 nitrous acid

Writing Formulas for Acids

- when name ends in acid, formulas starts with H
- write formulas as if ionic, even though it is molecular
- hydro prefix means it is binary acid, no prefix means it is an oxyacid
- for oxyacid, if ending is –*ic*, polyatomic ion ends in –*ate*; if ending is –*ous*, polyatomic ion ends in –*ous*

Example – Binary Acids
hydrosulfuric acid

1. Write the symbol for the cation and its charge
2. Write the symbol for the anion and its charge
3. Charge (without sign) becomes subscript for other ion
4. Add (*aq*) to indicate dissolved in water
5. Check that the total charge of the cations cancels the total charge of the anions

H^+

in all acids the cation is H^+

S^{2-}

hydro means binary

$H^+ \; S^{2-}$ $\quad H_2S$

$H_2S(aq)$

$H = (2)\cdot(+1) = +2$

$S = (1)\cdot(-2) = -2$

Example – Oxyacids
carbonic acid

1. Write the symbol for the cation and its charge
2. Write the symbol for the anion and its charge
3. Charge (without sign) becomes subscript for other ion
4. Add (*aq*) to indicate dissolved in water
5. Check that the total charge of the cations cancels the total charge of the anions

H^+

in all acids the cation is H^+

CO_3^{2-}

no *hydro* means polyatomic ion

-ic means *-ate* ion

$H^+ \; CO_3^{2-}$ $\quad H_2CO_3$

$H_2CO_3(aq)$

$H = (2)\cdot(+1) = +2$

$CO_3 = (1)\cdot(-2) = -2$

Example – Oxyacids
sulfurous acid

1. Write the symbol for the cation and its charge
2. Write the symbol for the anion and its charge
3. Charge (without sign) becomes subscript for other ion
4. Add (*aq*) to indicate dissolved in water
5. Check that the total charge of the cations cancels the total charge of the anions

H^+

in all acids the cation is H^+

SO_3^{2-}

no *hydro* means polyatomic ion

-ous means *-ite* ion

$H^+ \; SO_3^{2-}$ → H_2SO_3

$H_2SO_3(aq)$

$H = (2)\cdot(+1) = +2$

$SO_3 = (1)\cdot(-2) = -2$

Practice - What are the formulas for the following acids?

1. chlorous acid

2. phosphoric acid

3. hydrobromic acid

Practice - What are the formulas for the following acids?

1. H^+ with ClO_2^- $HClO_2$

2. H^+ with PO_4^{3-} H_3PO_4

3. H^+ with Br^- HBr

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(1.01 amu H) + 16.00 amu O = 18.02 amu

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

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

Molar Mass = 1 mole H_2O

= 2(1.01 g H) + 16.00 g O = 18.02 g

so the Molar Mass of H_2O is 18.02 g/mole

Example – Find the number of CO_2 molecules in 10.8 g of dry ice

Given: **Find:**	10.8 g CO_2 molecules CO_2
Concept Plan:	g CO_2 $\xrightarrow{\frac{1\,\text{mol}}{44.01\,\text{g}}}$ mol CO_2 $\xrightarrow{\frac{6.022\times10^{23}\ \text{molecules}}{1\,\text{mol}}}$ molec CO_2
Relationships:	1 mol CO_2 = 44.01 g, 1 mol = 6.022 x 10^{23}
Solution:	$10.8\,\text{g}\,CO_2 \times \frac{1\,\text{mol}\,CO_2}{44.01\,\text{g}\,CO_2} \times \frac{6.022\times10^{23}\ \text{molecules}}{1\,\text{mol}}$ $= 1.48\times10^{23}$ molecules CO_2
Check:	since the given amount is much less than 1 mol CO_2, the number makes sense

Practice - Converting Grams to Molecules

How many molecules are in 50.0 g of PbO_2?

(PbO_2 = 239.2)

Practice - Converting Grams to Molecules

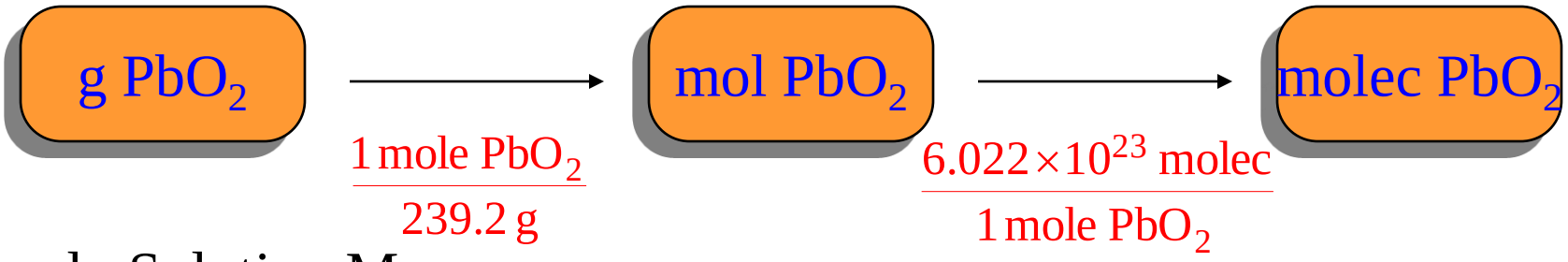
How many molecules are in 50.0 g of PbO_2?

Given: 50.0 g PbO_2

Find: molecules PbO_2

Relationships: 1 mole $PbO_2 \equiv 239.2$ g; 1 mol $\equiv 6.022 \times 10^{23}$ molec

Concept Plan:



Apply Solution Map:

$$50.0\text{ g } PbO_2 \times \frac{1\text{ mole } PbO_2}{239.2\text{ g}} \times \frac{6.022\times10^{23}\text{ molec}}{1\text{ mole } PbO_2} = 1.26\times10^{23}\text{ molec } PbO_2$$

Check Answer:

Units are correct. Number makes sense because given amount less than 1 mole

Percent Composition

- Percentage of each element in a compound
  - ✓ By mass
- Can be determined from

1. the formula of the compound
2. the experimental mass analysis of the compound

- The percentages may not always total to 100% due to rounding

$$\text{Percentage} = \frac{\text{part}}{\text{whole}} \times 100\%$$

Example 3.13 – Find the mass percent of Cl in $C_2Cl_4F_2$

Given:	$C_2Cl_4F_2$
Find:	% Cl by mass
Concept Plan:	$\text{Mass\% Cl} = \dfrac{4 \times \text{molar mass Cl}}{\text{molar mass } C_2Cl_4F_2} \times 100\%$
Relationships:	$\text{Mass\% element } X = \dfrac{\text{mass element } X \text{ in 1 mol}}{\text{mass 1 mol of compound}} \times 100\%$
Solution:	$4 \times \text{molar mass Cl} = 4(35.45\ \text{g/mol}) = 141.8\ \text{g/mol}$ $\text{molar mass } C_2Cl_4F_2 = 2(12.01) + 4(35.45) + 2(19.00) = 203.8\ \text{g/mol}$ $\text{Mass\% Cl} = \dfrac{141.8\ \cancel{\text{g/mol}}}{203.8\ \cancel{\text{g/mol}}} \times 100\% = 69.58\%$
Check:	since the percentage is less than 100 and Cl is much heavier than the other atoms, the number makes sense

Practice - Determine the Percent Composition of the following

- $CaCl_2$

$$\text{Mass \% Ca} = \frac{\text{molar mass Ca}}{\text{molar mass } CaCl_2} \times 100\%$$

$$\text{Mass \% Cl} = \frac{2 \times \text{molar mass Cl}}{\text{molar mass } CaCl_2} \times 100\%$$

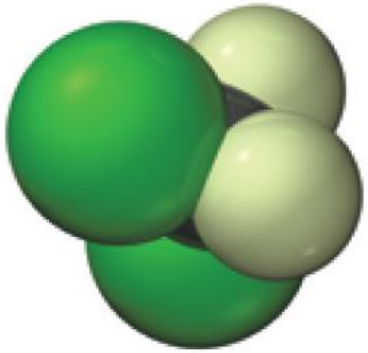
$$2 \times \text{molar mass Cl} = 2(35.45\ \text{g/mol}) = 70.90\ \text{g/mol}$$

$$\text{molar mass } CaCl_2 = 1(40.08) + 2(35.45) = 110.98\ \text{g/mol}$$

$$\text{Mass \% Ca} = \frac{40.08\ \text{g/mol}}{110.98\ \text{g/mol}} \times 100\% = 36.11\%$$

$$\text{Mass \% Cl} = \frac{70.90\ \text{g/mol}}{110.98\ \text{g/mol}} \times 100\% = 63.88\%$$

Mass Percent as a Conversion Factor



- the mass percent tells you the mass of a constituent element in 100 g of the compound
  - ✓the fact that CCl_2F_2 is 58.64% Cl by mass means that 100 g of CCl_2F_2 contains 58.64 g Cl
- this can be used as a conversion factor
  - ✓100 g CCl_2F_2 : 58.64 g Cl

$$\text{g } CCl_2F_2 \times \frac{58.64 \text{ g Cl}}{100 \text{ g } CCl_2F_2} = \text{g Cl} \qquad \text{g Cl} \times \frac{100 \text{ g } CCl_2F_2}{58.64 \text{ g Cl}} = \text{g } CCl_2F_2$$

Example 3.14 – Find the mass of table salt containing 2.4 g of Na

Given: **Find:**	2.4 g Na, 39% Na g NaCl
Concept Plan:	g Na → g NaCl $\frac{100\text{ g NaCl}}{39\text{ g Na}}$
Relationships:	100. g NaCl : 39 g Na
Solution:	$2.4\,\cancel{\text{g Na}} \times \frac{100\text{ g NaCl}}{39\,\cancel{\text{g Na}}} = 6.2\text{ g NaCl}$
Check:	since the mass of NaCl is more than 2x the mass of Na, the number makes sense

Practice – Benzaldehyde is 79.2% carbon. What mass of benzaldehyde contains 19.8 g of C?

Practice – Benzaldehyde is 79.2% carbon. What mass of benzaldehyde contains 19.8 g of C?

Given: **Find:**	19.8 g C, 79.2% C g benzaldehyde
Concept Plan: **Relationships:**	g C → g benzaldehyde $\frac{100\text{ g benzaldehyde}}{79.2\text{ g C}}$ 100. g benzaldehyde : 79.2 g C
Solution:	$19.8\,\text{g C} \times \frac{100\text{ g benzaldehyde}}{79.2\,\text{g C}} = 25.0\text{ g benzaldehyde}$
Check:	since the mass of benzaldehyde is more than the mass of C, the number makes sense

Conversion Factors in Chemical Formulas

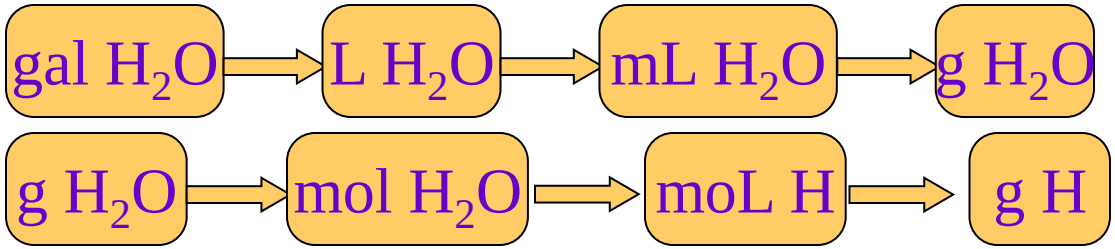
- chemical formulas have inherent in them relationships between numbers of atoms and molecules
  - ✓or moles of atoms and molecules
- these relationships can be used to convert between amounts of constituent elements and molecules
  - ✓like percent composition

Example 3.15 – Find the mass of hydrogen in 1.00 gal of water

Given: 1.00 gal H_2O, d_{H2O} = 1.00 g/ml

Find: g H

Concept Plan:



Relationships: 3.785 L = 1 gal, 1 L = 1000 mL, 1.00 g H_2O = 1 mL, 1 mol H_2O = 18.02 g, 1 mol H = 1.008 g, 2 mol H : 1 mol H_2O

Solution:

$$1.00\ \text{gal}\ H_2O \times \frac{3.785\ \text{L}}{1\ \text{gal}} \times \frac{1000\ \text{mL}}{1\ \text{L}} \times \frac{1.00\ \text{g}}{1\ \text{mL}} = 3.785 \times 10^3\ \text{g}\ H_2O$$

$$3.785 \times 10^3\ \text{g}\ H_2O \times \frac{1\ \text{mol}\ H_2O}{18.02\ \text{g}} \times \frac{2\ \text{mol H}}{1\ \text{mol}\ H_2O} \times \frac{1.008\ \text{g H}}{1\ \text{mol H}} = 4.23 \times 10^2\ \text{g H}$$

Check: since 1 gallon weighs about 3800 g, and H is light, the number makes sense

Practice - How many grams of sodium are in 6.2 g of NaCl? (Na = 22.99; Cl = 35.45)

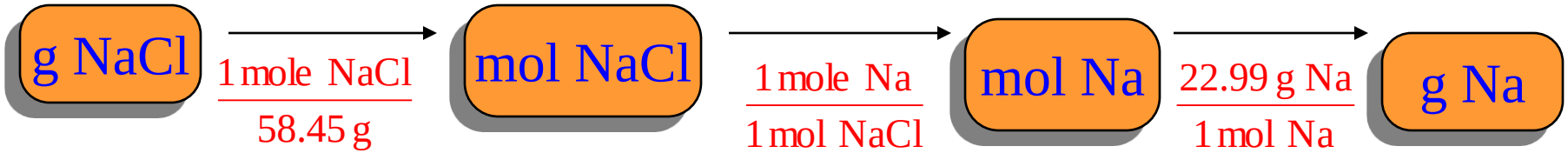
How many grams of sodium are in 6.2 g of NaCl?

Given: 6.2 g NaCl

Find: g Na

Rel: 1 mole NaCl ≡ 58.45 g; 1 mol Na ≡ 1 mol NaCl;
1 mol Na ≡ 22.99 g Na

Concept Plan:



Apply Concept Plan:

$$6.2\,\cancel{g}\ \text{NaCl} \times \frac{1\ \cancel{\text{mole NaCl}}}{58.45\,\cancel{g}} \times \frac{1\ \cancel{\text{mole Na}}}{1\ \cancel{\text{mole NaCl}}} \times \frac{22.99\ \text{g Na}}{1\ \cancel{\text{mole Na}}} = 2.4\ \text{g Na}$$

Check Answer:

Units are correct. Number makes sense because given amount less than 1 mole NaCl.

Empirical Formula

- simplest, whole-number ratio of the atoms of elements in a compound
- can be determined from elemental analysis
  - ✓masses of elements formed when decompose or react compound
    - ➢combustion analysis
  - ✓percent composition

Finding an Empirical Formula

1) convert the percentages to grams
   a) assume you start with 100 g of the compound
   b) skip if already grams
2) convert grams to moles
   a) use molar mass of each element
3) write a pseudoformula using moles as subscripts
4) divide all by smallest number of moles
   a) if result is within 0.1 of whole number, round to whole number
5) multiply all mole ratios by number to make all whole numbers
   a) if ratio ?.5, multiply all by 2; if ratio ?.33 or ?.67, multiply all by 3; if ratio 0.25 or 0.75, multiply all by 4; etc.
   b) skip if already whole numbers

Example 3.17

- Laboratory analysis of aspirin determined the following mass percent composition. Find the empirical formula.

C = 60.00%

H = 4.48%

O = 35.53%

Example:
Find the empirical formula of aspirin with the given mass percent composition.

Write down the given quantity and its units.

Given: C = 60.00%

H = 4.48%

O = 35.53%

Therefore, in 100 g of aspirin there are 60.00 g C, 4.48 g H, and 35.53 g O

Example:
Find the empirical formula of aspirin with the given mass percent composition.

Information
Given: 60.00 g C, 4.48 g H, 35.53 g O

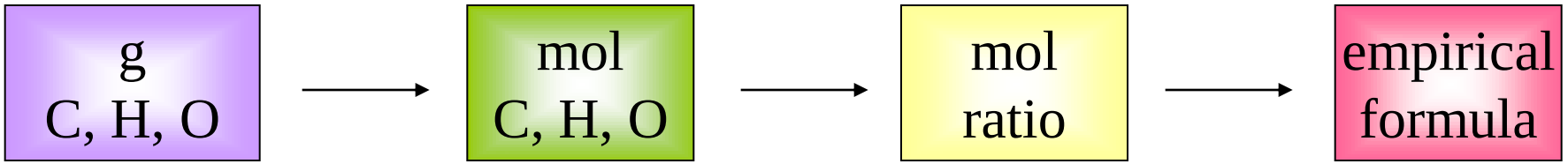
Write down the quantity to find and/or its units.

Find: empirical formula, $C_xH_yO_z$

Example:
Find the empirical formula of aspirin with the given mass percent composition.

Information
Given: 60.00 g C, 4.48 g H, 35.53 g O
Find: Empirical Formula, $C_xH_yO_z$

Write a Concept Plan:



Example:
Find the empirical formula of aspirin with the given mass percent composition.

Information
Given: 60.00 g C, 4.48 g H, 35.53 g O
Find: Empirical Formula, $C_xH_yO_z$
CP: g C,H,O → mol C,H,O → mol ratio → empirical formula

Collect Needed Relationships:

1 mole C = 12.01 g C

1 mole H = 1.008 g H

1 mole O = 16.00 g O

Example:
Find the empirical formula of aspirin with the given mass percent composition.

Information
Given: 60.00 g C, 4.48 g H, 35.53 g O
Find: Empirical Formula, $C_xH_yO_z$
CP: g C,H,O → mol C,H,O → mol ratio → empirical formula
Rel: 1 mol C = 12.01 g;
1 mol H = 1.008 g; 1 mol O = 16.00 g

Apply the Concept Plan:

✓ calculate the moles of each element

$$60.00\,\text{g C} \times \frac{1\,\text{mol C}}{12.01\,\text{g C}} = 4.996\,\text{mol C}$$

$$4.48\,\text{g H} \times \frac{1\,\text{mol H}}{1.008\,\text{g H}} = 4.44\,\text{mol H}$$

$$35.53\,\text{g O} \times \frac{1\,\text{mol O}}{16.00\,\text{g O}} = 2.220\,\text{mol O}$$

Example:
Find the empirical formula of aspirin with the given mass percent composition.

Information

Given: 4.996 mol C, 4.44 mol H, 2.220 mol O

Find: Empirical Formula, $C_xH_yO_z$

CP: g C,H,O → mol C,H,O → mol ratio → empirical formula

Rel: 1 mol C = 12.01 g; 1 mol H = 1.008 g; 1 mol O = 16.00 g

Apply the Concept Plan:

✓ write a pseudoformula

$$C_{4.996}H_{4.44}O_{2.220}$$

Example:
Find the empirical formula of aspirin with the given mass percent composition.

Information

Given: $C_{4.996}H_{4.44}O_{2.220}$

Find: Empirical Formula, $C_xH_yO_z$

CP: g C,H,O → mol C,H,O → mol ratio → empirical formula

Rel: 1 mol C = 12.01 g;
1 mol H = 1.008 g; 1 mol O = 16.00 g

Apply the Concept Plan:

- ✓ find the mole ratio by dividing by the smallest number of moles

$$C_{\frac{4.996}{2.220}}H_{\frac{4.44}{2.220}}O_{\frac{2.220}{2.220}}$$

$$C_{2.25}H_2O_1$$

Example:
Find the empirical formula of aspirin with the given mass percent composition.

Information

Given: $C_{2.25}H_2O_1$

Find: Empirical Formula, $C_xH_yO_z$

CP: g C,H,O → mol C,H,O → mol ratio → empirical formula

Rel: 1 mol C = 12.01 g;
1 mol H = 1.008 g; 1 mol O = 16.00 g

Apply the Concept Plan:

- ✓ multiply subscripts by factor to give whole number

$$\{C_{2.25}H_2O_1\} \times 4$$

$$C_9H_8O_4$$

Practice – Determine the empirical formula of hematite, which contains 72.4% Fe (55.85) and the rest oxygen (16.00)

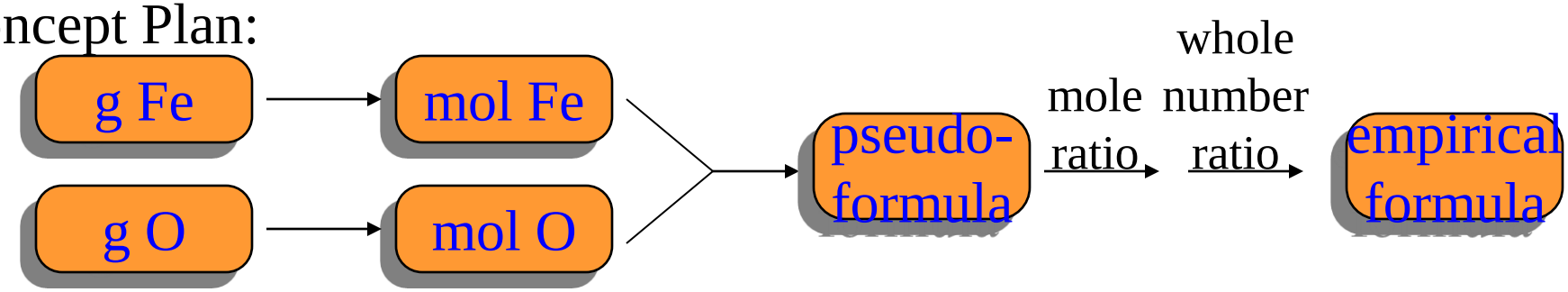
Practice – Determine the empirical formula of hematite, which contains 72.4% Fe (55.85) and the rest oxygen (16.00)

Given: 72.4% Fe, (100 – 72.4) = 27.6% O ∴
in 100 g hematite there are 72.4 g Fe and 27.6 g O

Find: Fe_xO_y

Rel: 1 mol Fe = 55.85 g; 1 mol O = 16.00 g

Concept Plan:



Practice – Determine the empirical formula of hematite, which contains 72.4% Fe (55.85) and the rest oxygen (16.00)

Apply the Concept Plan:

$$72.4\,\cancel{g}\,\text{Fe} \times \frac{1\,\text{mol Fe}}{55.85\,\cancel{g}} = 1.30\,\text{mol Fe}$$

$$26.7\,\cancel{g}\,\text{O} \times \frac{1\,\text{mol O}}{16.00\,\cancel{g}} = 1.73\,\text{mol O}$$

$$\text{Fe}_{1.30}\text{O}_{1.73}$$

$$\text{Fe}_{\frac{1.30}{1.30}}\text{O}_{\frac{1.73}{1.30}} = \text{Fe}_1\text{O}_{1.33}$$

$$\left\{\text{Fe}_1\text{O}_{1.33}\right\} \times 3 = \text{Fe}_3\text{O}_4$$

Molecular Formulas

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

$$\frac{\text{Molar Mass}_{\text{molecular formula}}}{\text{Empirical Formula Mass}} = \text{multiplyin g factor, } n$$

Example 3.18 – Find the molecular formula of butanedione

Given:	emp. form. = C_2H_3O; MM = 86.03 g/mol
Find:	molecular formula
Concept Plan: and Relationships:	$\text{Molecular Form.} = \text{Emp. Form.} \times n$ $n = \dfrac{\text{Molar Mass}}{\text{Emp. Form. Molar Mass}}$

Solution:

$$\text{Molar Mass Emp. Form.} =$$

$$2(12.01\,\text{g/mol}) + 3(1.008\,\text{g/mol}) + 1(16.00\,\text{g/mol}) = 43.04\,\text{g/mol}$$

$$n = \frac{86.09\,\text{g/mol}}{43.04\,\text{g/mol}} = 2 \qquad \text{Molecular Formula} = C_2H_3O \times 2 = C_4H_6O_2$$

Check:	the molar mass of the calculated formula is in agreement with the given molar mass

Practice – Benzopyrene has a molar mass of 252 g/mol and an empirical formula of C_5H_3. What is its molecular formula? (C = 12.01, H=1.01)

Practice – Benzopyrene has a molar mass of 252 g and an empirical formula of C_5H_3. What is its molecular formula? (C = 12.01, H=1.01)

$$C_5 = 5(12.01\ g) = 60.05\ g$$
$$\underline{H_3 = 3(1.01\ g) = 3.03\ g}$$
$$C_5H_3 = 63.08\ g$$

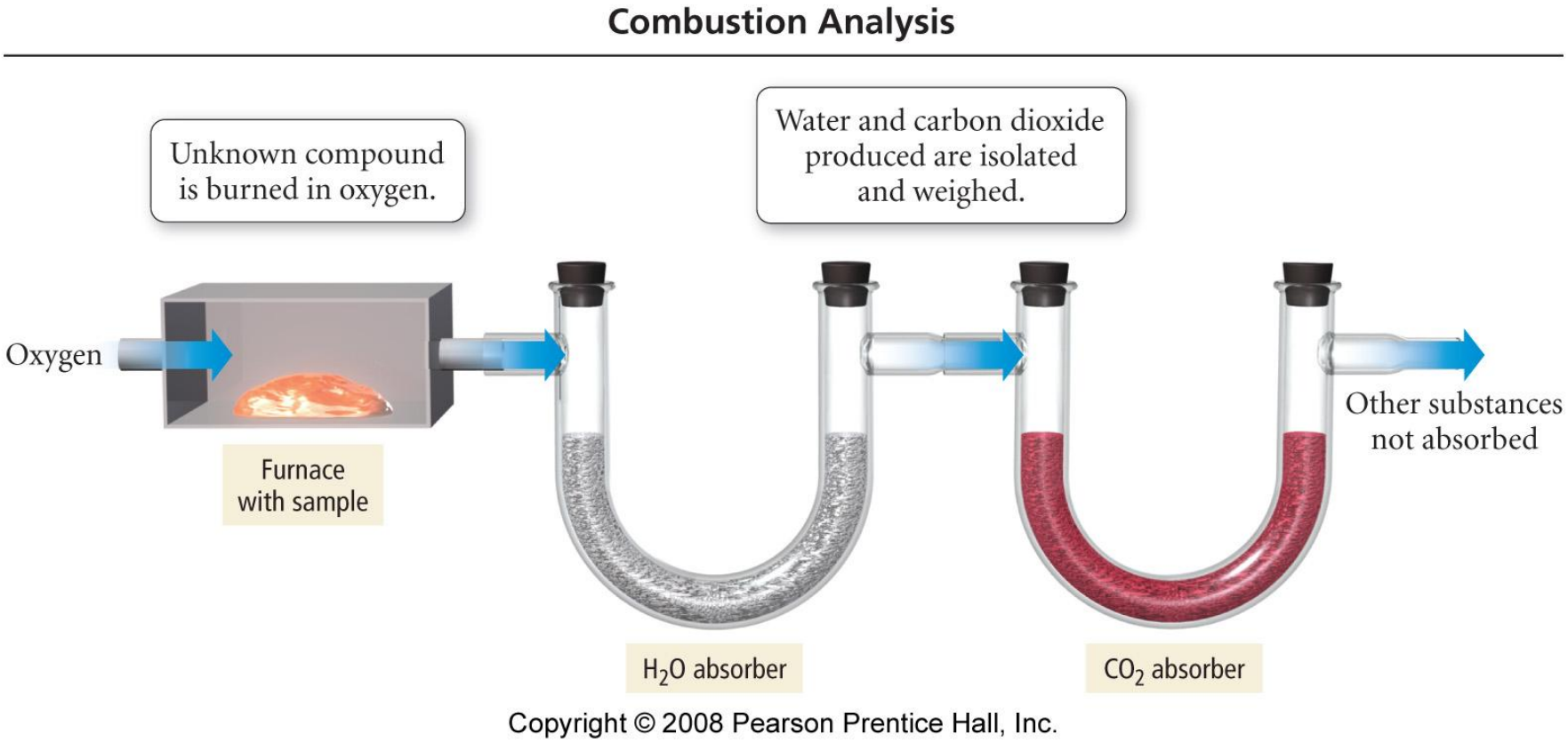
$$n = \frac{252\ \cancel{g/mol}}{63.08\ \cancel{g/mol}} = 4$$

Molecular Formula = $\{C_5H_3\}$ x 4 = $C_{20}H_{12}$

Combustion Analysis

- a common technique for analyzing compounds is to burn a known mass of compound and weigh the amounts of product made
  - ✓ generally used for organic compounds containing C, H, O
- by knowing the mass of the product 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, 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



Example 3.20

- Combustion of a 0.8233 g sample of a compound containing only carbon, hydrogen, and oxygen produced the following:

$$CO_2 = 2.445 \text{ g}$$

$$H_2O = 0.6003 \text{ g}$$

Determine the empirical formula of the compound

Example 3.20:
Find the empirical formula of compound with the given amounts of combustion products

Write down the given quantity and its units.

Given: compound = 0.8233 g

CO_2 = 2.445 g

H_2O = 0.6003 g

Example 3.20:
Find the empirical formula of compound with the given amounts of combustion products

Information
Given: 0.8233 g compound,
2.445 g CO_2, 0.6003 g H

Write down the quantity to find and/or its units.

Find: empirical formula, $C_xH_yO_z$

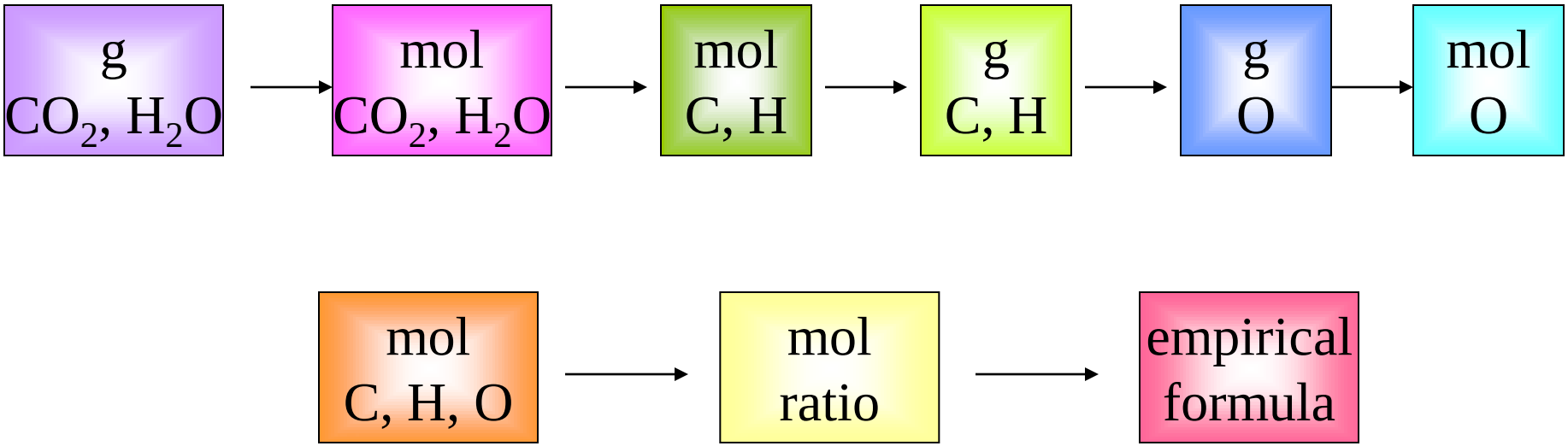
Example 3.20:
Find the empirical formula of compound with the given amounts of combustion products

Information

Given: 0.8233 g compound,
2.445 g CO_2, 0.6003 g H

Find: Empirical Formula, $C_xH_yO_z$

Write a Concept Plan:



Example 3.20:
Find the empirical formula of compound with the given amounts of combustion products

Information

Given: 0.8233 g compound,
2.445 g CO_2, 0.6003 g H_2O

Find: Empirical Formula, $C_xH_yO_z$

CP: g CO_2 & H_2O → mol CO_2 & H_2O → mol C & H → g C & H → g O → mol O → mol ratio → empirical formula

Collect Needed Relationships:

1 mole CO_2 = 44.01 g CO_2

1 mole H_2O = 18.02 g H_2O

1 mole C = 12.01 g C

1 mole H = 1.008 g H

1 mole O = 16.00 g O

1 mole CO_2 = 1 mole C

1 mole H_2O = 2 mole H

Example 3.20:
Find the empirical formula of compound with the given amounts of combustion products

Information

Given: 0.8233 g compound
2.445 g CO_2, 0.6003 g H_2O

Find: Empirical Formula, $C_xH_yO_z$

CP: g CO_2 & H_2O → mol CO_2 & H_2O → mol C & H → g C & H → g O → mol O → mol ratio → empirical formula

Rel: MM of CO_2, H_2O, C, H, O; mol element : 1 mol compound

Apply the Concept Plan:

✓ calculate the moles of C and H

$$2.445\,\text{g}\,CO_2 \times \frac{1\,\text{mol}\,CO_2}{44.01\,\text{g}\,CO_2} \times \frac{1\,\text{mol C}}{1\,\text{mol}\,CO_2} = 0.05556\,\text{mol C}$$

$$0.6003\,\text{g}\,H_2O \times \frac{1\,\text{mol}\,H_2O}{18.02\,\text{g}\,H_2O} \times \frac{2\,\text{mol H}}{1\,\text{mol}\,H_2O} = 0.06662\,\text{mol H}$$

Example 3.20: Find the empirical formula of compound with the given amounts of combustion products

Information

Given: 0.8233 g compound, 2.445 g CO_2, 0.6003 g H_2O, 0.05556 mol C, 0.06662 mol H

Find: Empirical Formula, $C_xH_yO_z$

CP: g CO_2 & H_2O → mol CO_2 & H_2O → mol C & H → g C & H → g O → mol O → mol ratio → emp. formula

Rel: MM of CO_2, H_2O, C, H, O; mol element : 1 mol compound

Apply the Concept Plan:

✓ calculate the grams of C and H

$$0.05556 \text{ mol C} \times \frac{12.01 \text{ g}}{1 \text{ mol C}} = 0.6673 \text{ g C}$$

$$0.06662 \text{ mol H} \times \frac{1.008 \text{ g}}{1 \text{ mol H}} = 0.06715 \text{ g H}$$

Example 3.20: Find the empirical formula of compound with the given amounts of combustion products

Information

Given: 0.8233 g compound, 2.445 g CO_2, 0.6003 g H_2O, 0.05556 mol C, 0.6673 g C, 0.06662 mol H, 0.06715 g H,

Find: Empirical Formula, $C_xH_yO_z$

CP: g CO_2 & H_2O → mol CO_2 & H_2O → mol C & H → g C & H → g O → mol O → mol ratio → emp. formula

Rel: MM of CO_2, H_2O, C, H, O; mol element : 1 mol compound

Apply the Concept Plan:

✓ calculate the grams and moles of O

$$0.8233\,\text{g compound} - (0.6673\,\text{g C} + 0.06715\,\text{g H}) = 0.0889\,\text{g O}$$

$$0.0889\,\text{g O} \times \frac{1\,\text{mol O}}{16.00\,\text{g}} = 0.00556\,\text{mol O}$$

Example 3.20: Find the empirical formula of compound with the given amounts of combustion products

Information

Given: 0.8233 g compound, 2.445 g CO_2, 0.6003 g H_2O, 0.05556 mol C, 0.6673 g C, 0.06662 mol H, 0.06715 g H, 0.0889 g O, 0.00556 mol O

Find: Empirical Formula, $C_xH_yO_z$

CP: g CO_2 & H_2O → mol CO_2 & H_2O → mol C & H → g C & H → g O → mol O → mol ratio → emp. formula

Rel: MM of CO_2, H_2O, C, H, O; mol element : 1 mol compound

Apply the Concept Plan:

✓ write a pseudoformula

$$C_{0.05556}H_{0.06662}O_{0.00556}$$

Example 3.20: Find the empirical formula of compound with the given amounts of combustion products

Information

Given: 0.8233 g compound, 2.445 g CO_2, 0.6003 g H_2O, 0.05556 mol C, 0.6673 g C, 0.06662 mol H, 0.06715 g H, 0.0889 g O, 0.00556 mol O

Find: Empirical Formula, $C_xH_yO_z$

CP: g CO_2 & H_2O → mol CO_2 & H_2O → mol C & H → g C & H → g O → mol O → mol ratio → emp. formula

Rel: MM of CO_2, H_2O, C, H, O; mol element : 1 mol compound

Apply the Concept Plan:

- ✓ find the mole ratio by dividing by the smallest number of moles

$$C_{\frac{0.05556}{0.00556}}H_{\frac{0.06662}{0.00556}}O_{\frac{0.00556}{0.00556}}$$

$$C_{10}H_{12}O_1$$

Example 3.20: Find the empirical formula of compound with the given amounts of combustion products

Information

Given: 0.8233 g compound, 2.445 g CO_2, 0.6003 g H_2O, 0.05556 mol C, 0.6673 g C, 0.06662 mol H, 0.06715 g H, 0.0889 g O, 0.00556 mol O

Find: Empirical Formula, $C_xH_yO_z$

CP: g CO_2 & H_2O → mol CO_2 & H_2O → mol C & H → g C & H → g O → mol O → mol ratio → emp. formula

Rel: MM of CO_2, H_2O, C, H, O;
mol element : 1 mol compound

Apply the Concept Plan:

- ✓ multiply subscripts by factor to give whole number, if necessary
- ✓ write the empirical formula

$$C_{10}H_{12}O_1$$

The smell of dirty gym socks is caused by the compound caproic acid. Combustion of 0.844 g of caproic acid produced 0.784 g of H_2O and 1.92 g of CO_2. If the molar mass of caproic acid is 116.2 g/mol, what is the molecular formula of caproic acid? (MM C = 12.01, H = 1.008, O = 16.00)

Combustion of 0.844 g of caproic acid produced 0.784 g of H_2O and 1.92 g of CO_2. If the molar mass of caproic acid is 116.2 g/mol, what is the molecular formula of caproic acid?

$$1.92\,\text{g}\,CO_2 \times \frac{1\,\text{mol}\,CO_2}{44.01\,\text{g}\,CO_2} \times \frac{1\,\text{mol C}}{1\,\text{mol}\,CO_2} = 0.0436\,\text{mol C}$$

$$0.784\,\text{g}\,H_2O \times \frac{1\,\text{mol}\,H_2O}{18.02\,\text{g}\,H_2O} \times \frac{2\,\text{mol H}}{1\,\text{mol}\,H_2O} = 0.0870\,\text{mol H}$$

$$0.0436\,\text{mol C} \times \frac{12.01\,\text{g}}{1\,\text{mol C}} = 0.524\,\text{g C}$$

$$0.0870\,\text{mol H} \times \frac{1.008\,\text{g}}{1\,\text{mol H}} = 0.0877\,\text{g H}$$

	C	H	O
g	0.524	0.0877	0.232
moles	0.0436	0.0870	0.0145

$$0.844\text{ g compound} - (0.524\text{ g C} + 0.0877\text{ g H}) = 0.232\text{ g O}$$

$$0.232\text{ g O} \times \frac{1\text{ mol O}}{16.00\text{ g}} = 0.0145\text{ mol O}$$

$$C_{0.0436}H_{0.0870}O_{0.0145}$$

$$C_{\frac{0.0436}{0.0145}}H_{\frac{0.0870}{0.0145}}O_{\frac{0.0145}{0.0145}}$$

$$C_3H_6O_1$$

$$C_3 = 3(12.01\,g) = 36.03\,g$$
$$H_6 = 6(1.008\,g) = 6.048\,g$$
$$O_1 = 1(16.00\,g) = 16.00\,g$$
$$C_3H_6O = 58.08\,g$$

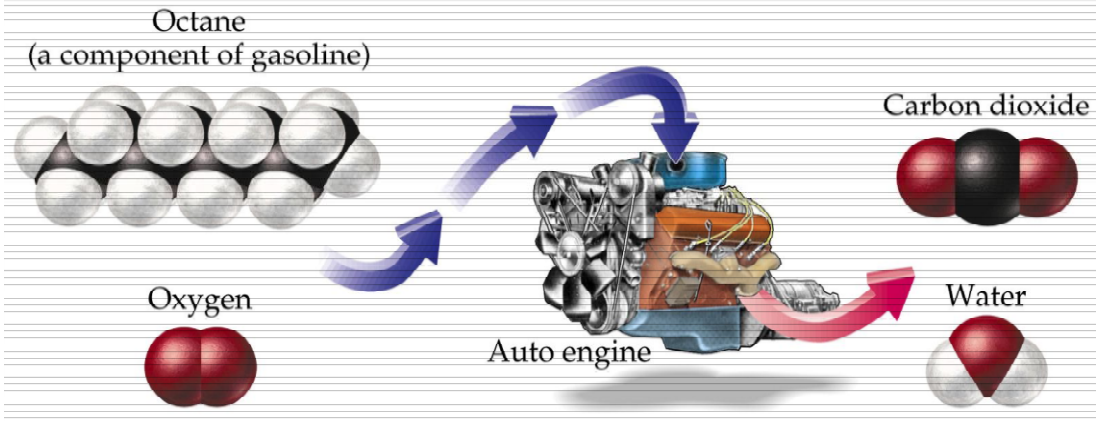
$$n = \frac{116.2\ \cancel{g/mol}}{58.08\ \cancel{g/mol}} = 2$$

Molecular Formula = $\{C_3H_6O\}$ x 2 = $C_6H_{12}O_2$

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

Reactants → Products



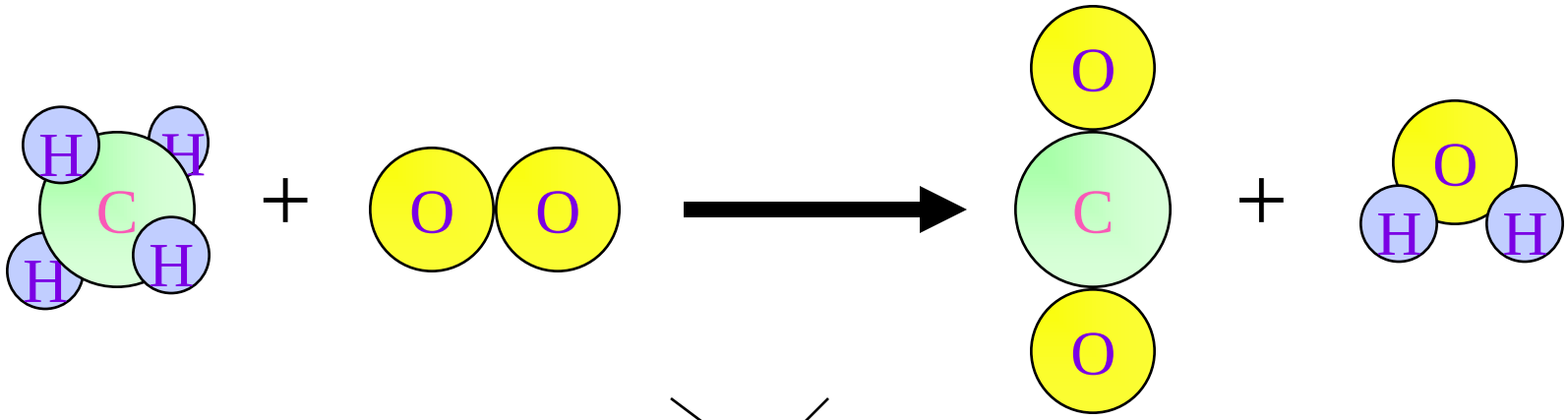
Chemical Equations

- Shorthand way of describing a reaction
- Provides 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

Combustion of Methane

- methane gas burns to produce carbon dioxide gas and gaseous water
  - ✓ whenever something burns it combines with $O_2(g)$

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



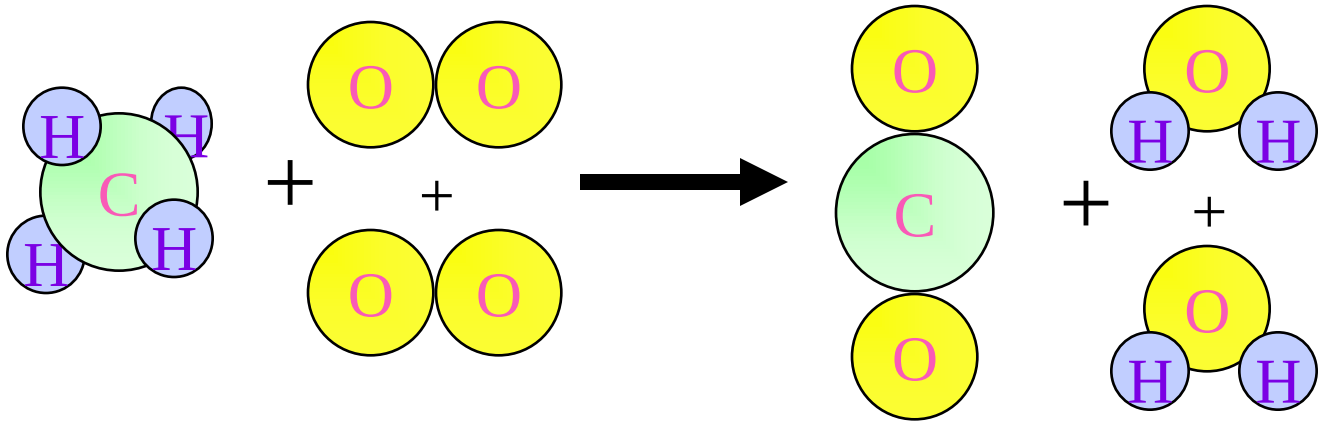
1 C + 4 H + 2 O → 1 C + 2 O + 2 H + O

1 C + 2 H + 3 O

Combustion of Methane Balanced

- to show the reaction obeys the Law of Conservation of Mass, it must be **balanced**

$$CH_4(g) + \mathbf{2}\ O_2(g) \rightarrow CO_2(g) + \mathbf{2}\ H_2O(g)$$



$$1\,C + 4\,H + 4\,O \longrightarrow 1\,C + 4\,H + 4\,O$$

Chemical Equations

$$CH_4(g) + \mathbf{2}\ O_2(g) \rightarrow CO_2(g) + \mathbf{2}\ H_2O(g)$$

- CH_4 and O_2 are the reactants, and CO_2 and H_2O are the products
- the (g) after the formulas tells us the state of the chemical
- the number in front of each substance tells us the numbers of those molecules in the reaction
  - ✓ called the **coefficients**

Chemical Equations

$$CH_4(g) + \mathbf{2}\ O_2(g) \rightarrow CO_2(g) + \mathbf{2}\ H_2O(g)$$

- this equation is balanced, meaning that there are equal numbers of atoms of each element on the reactant and product sides
  - ✓ to obtain the number of atoms of an element, multiply the subscript by the coefficient

$$1 \leftarrow C \rightarrow 1$$

$$4 \leftarrow H \rightarrow 4$$

$$4 \leftarrow O \rightarrow 2 + 2$$

Symbols Used in Equations

- symbols used to indicate state after chemical
  - ✓(*g*) = gas; (*l*) = liquid; (*s*) = solid
  - ✓(*aq*) = aqueous = dissolved in water
- energy symbols used above the arrow for decomposition reactions
  - ✓ Δ = heat
  - ✓ hν = light
  - ✓shock = mechanical
  - ✓elec = electrical

Example 3.22 Write a balanced equation for the combustion of butane, C_4H_{10}

Write a skeletal equation	$C_4H_{10}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
Balance atoms in complex substances first	$4 \Leftarrow C \Rightarrow 1 \times 4$ $C_4H_{10}(l) + O_2(g) \rightarrow 4\ CO_2(g) + H_2O(g)$ $10 \Leftarrow H \Rightarrow 2 \times 5$ $C_4H_{10}(l) + O_2(g) \rightarrow 4\ CO_2(g) + 5\ H_2O(g)$
Balance free elements by adjusting coefficient in front of free element	$13/2 \times 2 \Leftarrow O \Rightarrow 13$ $C_4H_{10}(l) + 13/2\ O_2(g) \rightarrow 4\ CO_2(g) + 5\ H_2O(g)$
If fractional coefficients, multiply thru by denominator	$\{C_4H_{10}(l) + 13/2\ O_2(g) \rightarrow 4\ CO_2(g) + 5\ H_2O(g)\} \times 2$ $2\ C_4H_{10}(l) + 13\ O_2(g) \rightarrow 8\ CO_2(g) + 10\ H_2O(g)$
Check	$8 \Leftarrow C \Rightarrow 8;\ 20 \Leftarrow H \Rightarrow 20;\ 26 \Leftarrow O \Rightarrow 26$

Practice

when aluminum metal reacts with air, it produces a white, powdery compound aluminum oxide

- ✓reacting with air means reacting with O_2

aluminum(*s*) + oxygen(*g*) → aluminum oxide(*s*)

$$Al(s) + O_2(g) \rightarrow Al_2O_3(s)$$

Practice

when aluminum metal reacts with air, it produces a white, powdery compound aluminum oxide

- ✓reacting with air means reacting with O_2

aluminum(*s*) + oxygen(*g*) → aluminum oxide(*s*)

$$Al(s) + O_2(g) \rightarrow Al_2O_3(s)$$

$$4\ Al(s) + 3\ O_2(g) \rightarrow 2\ Al_2O_3(s)$$

Practice

Acetic acid reacts with the metal aluminum to make aqueous aluminum acetate and gaseous hydrogen

- ✓acids are always aqueous
- ✓metals are solid except for mercury

Practice

Acetic acid reacts with the metal aluminum to make aqueous aluminum acetate and gaseous hydrogen

- ✓acids are always aqueous
- ✓metals are solid except for mercury

$$Al(s) + HC_2H_3O_2(aq) \rightarrow Al(C_2H_3O_2)_3(aq) + H_2(g)$$

$$2\ Al(s) + 6\ HC_2H_3O_2(aq) \rightarrow 2\ Al(C_2H_3O_2)_3(aq) + 3\ H_2(g)$$

Classifying Compounds Organic vs. Inorganic

- in the18th century, compounds from living things were called organic; compounds from the nonliving environment were called inorganic
- organic compounds easily decomposed and could not be made in 18th century lab
- inorganic compounds very difficult to decompose, but able to be synthesized

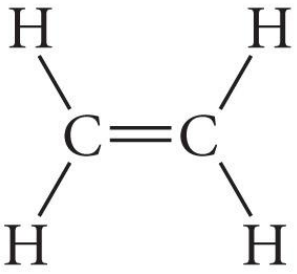
Modern Classifying Compounds Organic vs. Inorganic

- today we commonly make organic compounds in the lab and 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 main element that is the focus of organic chemistry is **carbon**

Carbon Bonding

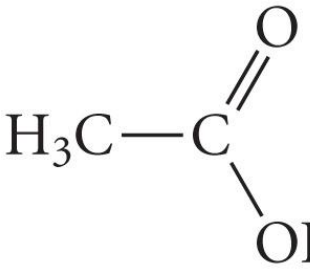
- carbon atoms bond almost exclusively covalently
  - ✓compounds with ionic bonding C are generally inorganic
- when C bonds, it forms 4 covalent bonds
  - ✓4 single bonds, 2 double bonds, 1 triple + 1 single, etc.
- carbon is unique in that it can form limitless chains of C atoms, both straight and branched, and rings of C atoms

Carbon Bonding

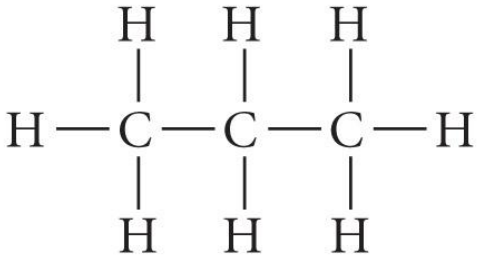


Ethene (C_2H_4)

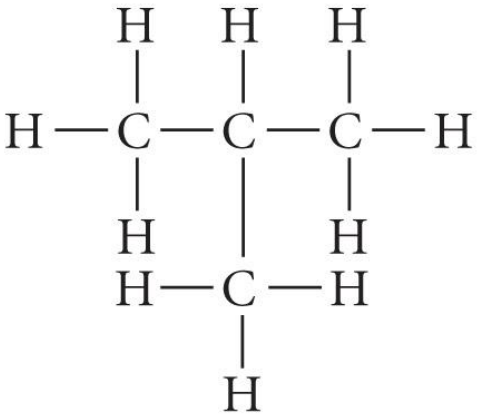
Ethyne (C_2H_2)



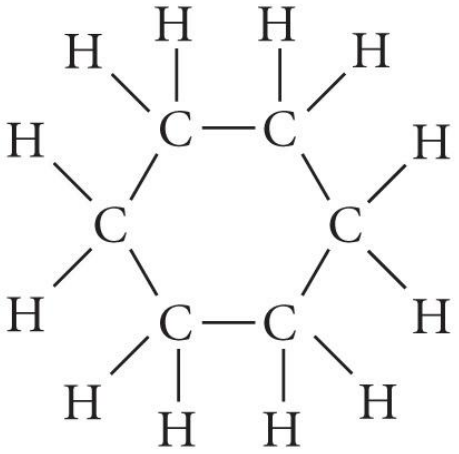
Acetic acid (CH_3COOH)



Propane (C_3H_8)



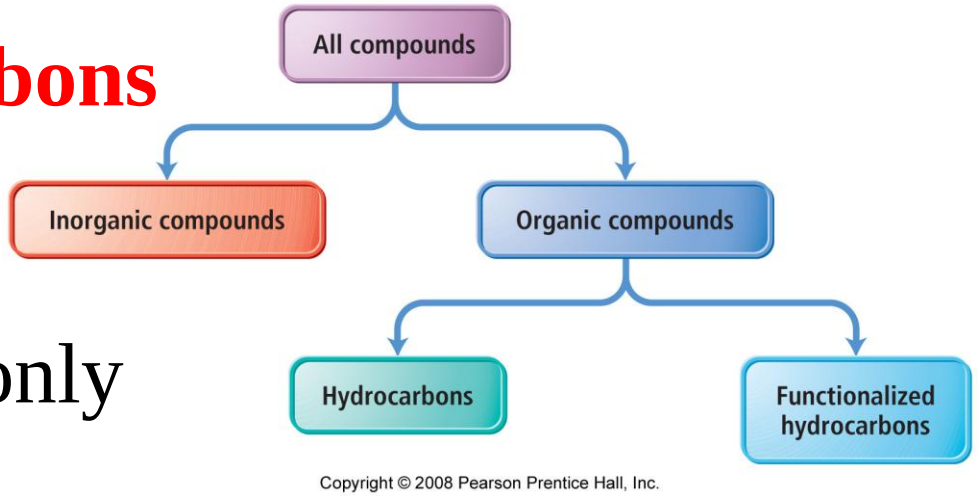
Isobutane (C_4H_{10})



Cyclohexane (C_6H_{12})

Classifying Organic Compounds

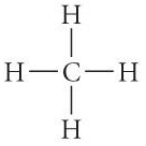
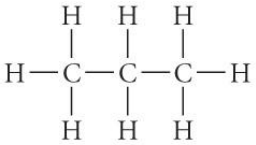
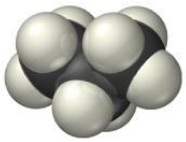
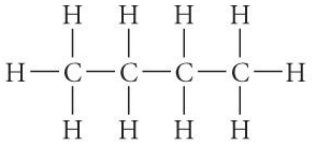
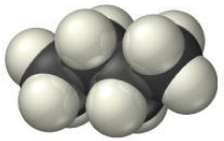
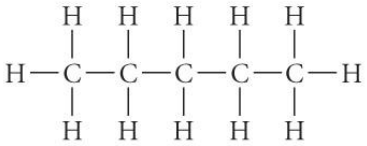
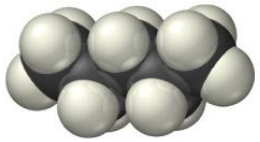
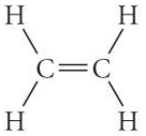
- there are two main categories of organic compounds, **hydrocarbons** and **functionalized hydrocarbons**
- hydrocarbons contain only C and H
- most fuels are mixtures of hydrocarbons



Classifying Hydrocarbons

- hydrocarbons containing only single bonds are called **alkanes**
- hydrocarbons containing one or more C=C are called **alkenes**
- hydrocarbons containing one or more C≡C are called **alkynes**
- hydrocarbons containing C_6 "benzene" ring are called **aromatic**

TABLE 3.7 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.

Naming Straight Chain Hydrocarbons

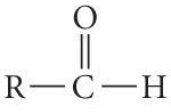
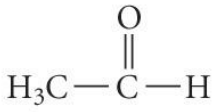
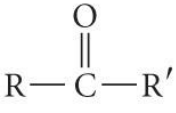
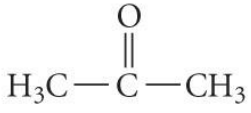
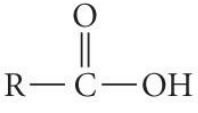
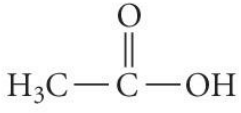
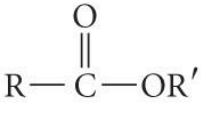
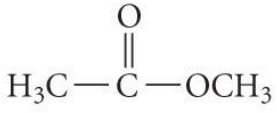
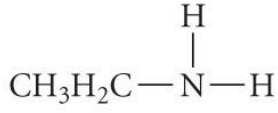
- consists of a base name to indicate the number of carbons in the chain, with a suffix to indicate the class and position of multiple bonds
  - ✓ suffix –ane for alkane, –ene for alkene, –yne for alkyne

Base Name	No. of C	Base Name	No. of C
meth-	1	hex-	6
eth-	2	hept-	7
prop-	3	oct-	8
but-	4	non-	9
pent-	5	dec-	10

Functionalized Hydrocarbons

- **functional groups** are non-carbon groups that are on the molecule
- substitute one or more functional groups replacing H's on the hydrocarbon chain
- generally, the chemical reactions of the compound are determined by the kinds of functional groups on the molecule

TABLE 3.8 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_3H_2C—O—CH_2CH_3	Diethyl ether	Anesthetic; laboratory solvent
Aldehydes	-al			Ethanal (acetaldehyde)	Perfumes; flavors
Ketones	-one			Propanone (acetone)	Fingernail polish remover
Carboxylic acids	acid			Acetic acid	Vinegar
Esters	-ate			Methyl acetate	Laboratory solvent
Amines	amine	RNH_2		Ethyl amine	Smell of rotten fish