7 Quantitative Composition of Compounds

Black pearls are composed of calcium carbonate, $CaCO_3$.
The pearls can be measured by either weighing or counting.

Foundations of College Chemistry, 14th Ed.
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Chapter Outline

The Mole

Individual atoms are tiny and have such a small mass, more convenient units for atoms are needed to be useful on the macroscale.

Analogy

Fruit in a supermarket is "counted" by weighing the mass of fruit.

If the average mass for a piece of fruit is known, the number of pieces of fruit can be calculated.

Example

If one orange has a mass of 186 g, then 75 oranges have what mass?

$$75 \cancel{\text{oranges}} \times \frac{186\text{ g}}{1 \cancel{\text{orange}}} = 13{,}950\text{ g} = 13.95\text{ kg}$$

Chemists count atoms in a similar way, by weighing.

The Mole

The standard unit of measurement for chemistry.

1 mole = 6.022×10^{23} objects

The number represented by 1 mole, 6.022×10^{23}, is also called Avogadro's number.

Such a large number is useful because even the smallest amount of matter contains extremely large numbers of atoms.

The mole is similar to other common units of counting.

Example

1 dozen = 12 objects

Eggs are measured by the dozen.

The Mole

Moles can be used to describe elements, particles or compounds.

Mole is often abbreviated as mol.

1 mol of atoms = 6.022×10^{23} atoms

1 mol of molecules = 6.022×10^{23} molecules

1 mol of electrons = 6.022×10^{23} electrons

Avogadro's number can be used as a conversion factor.

$$\frac{1\text{ mol}}{6.022 \times 10^{23}\text{ objects}} \qquad \frac{6.022 \times 10^{23}\text{ objects}}{1\text{ mol}}$$

The Mole

How does the mol relate to masses of elements?

The atomic mass of 1 mol of any element is defined as the amount of that substance that contains the same number of particles as exactly 12 g of ^{12}C.

1 mol of any element contains the same number of atoms, but can vary greatly in the overall mass.
(Atoms of different elements have different masses)

Molar Mass

Molar Mass: the atomic mass of an element or compound (in grams) which contains Avogadro's number of particles.

Molar masses are expressed to 4 significant figures in the text.

Determining Molar Mass

Convert atomic mass units on the periodic table to grams and sum the masses of the total atoms present.

Example CaF_2

Molar Mass CaF_2 = $\underbrace{40.08\text{ g}}_{Ca} + \underbrace{2(19.00)\text{ g}}_{2F} = 78.08\text{ g}$

Using the Mole and Molar Mass Concepts

We can use both the mol and molar mass as conversion factors.

How many moles of lead does 15.0 g of Pb represent?

Solution Map **g Pb ⟶ mol Pb**

The conversion factor relates g of Pb to moles of Pb.

$$\frac{1\text{ mol Pb}}{207.2\text{ g Pb}} \text{ or } \frac{207.2\text{ g Pb}}{1\text{ mol Pb}}$$

(Obtain molar mass from the periodic table.)

Calculate

$$15.0\ \cancel{\text{g Pb}} \times \frac{1\text{ mol Pb}}{207.2\ \cancel{\text{g Pb}}} = 7.24 \times 10^{-2}\text{ mol Pb}$$

Using the Mole and Molar Mass Concepts

How many moles of mercury does 23.0 g of Hg represent?

a. 4.62×10^{3} mol Hg
b. 1.15×10^{-1} mol Hg
c. 1.15×10^{1} mol Hg
d. 4.62×10^{-3} mol Hg

Solution Map

g Hg ⟶ mol Hg

The conversion factor needed:

$$\frac{1\text{ mol Hg}}{200.6\text{ g Hg}} \text{ or } \frac{200.6\text{ g Hg}}{1\text{ mol Hg}}$$

Calculate

$$23.0\ \cancel{\text{g Hg}} \times \frac{1\text{ mol Hg}}{200.6\ \cancel{\text{g Hg}}} = 1.15 \times 10^{-1}\text{ mol Hg}$$

Using the Mole and Molar Mass Concepts

How many Au atoms are contained in 16.0 g of Au?

Solution Map **g Au ⟶ mol Au ⟶ atoms Au**

Two conversion factors are needed:

$$\frac{1\text{ mol Au}}{197.0\text{ g Au}} \text{ and } \frac{1\text{ mol Au}}{6.022 \times 10^{23}\text{ atoms Au}}$$

Calculate

$$16.0\ \cancel{\text{g Au}} \times \frac{1\ \cancel{\text{mol Au}}}{197.0\ \cancel{\text{g Au}}} \times \frac{6.022 \times 10^{23}\text{ atoms Au}}{1\ \cancel{\text{mol Au}}}$$

$$= 4.89 \times 10^{22}\text{ atoms Au}$$

Using the Mole and Molar Mass Concepts

How many Ti atoms are contained in 7.80 g of Ti?

a. 2.71×10^{-25} atoms Ti
b. 2.25×10^{26} atoms Ti
c. 9.81×10^{22} atoms Ti
d. 6.20×10^{-22} atoms Ti

Solution Map

g Ti ⟶ mol Ti ⟶ atoms Ti

Two conversion factors are needed:

$$\frac{1\text{ mol Ti}}{47.87\text{ g Ti}} \text{ and } \frac{1\text{ mol Ti}}{6.022 \times 10^{23}\text{ atoms Ti}}$$

Calculate

$$7.80\ \cancel{\text{g Ti}} \times \frac{1\ \cancel{\text{mol Ti}}}{47.87\ \cancel{\text{g Ti}}} \times \frac{6.022 \times 10^{23}\text{ atoms Ti}}{1\ \cancel{\text{mol Ti}}}$$

$$= 9.81 \times 10^{22}\text{ atoms Ti}$$

Using the Mole and Molar Mass Concepts

What is the mass of 2.13×10^{18} atoms of Li?

Solution Map **atoms Li ⟶ mol Li ⟶ grams Li**

Two conversion factors are needed:

$$\frac{1\text{ mol Li}}{6.941\text{ g Li}} \text{ and } \frac{1\text{ mol Li}}{6.022 \times 10^{23}\text{ atoms Li}}$$

Calculate

$$2.13 \times 10^{18}\ \cancel{\text{atoms Li}} \times \frac{1\ \cancel{\text{mol Li}}}{6.022 \times 10^{23}\ \cancel{\text{atoms Li}}} \times \frac{6.941\text{ g Li}}{1\ \cancel{\text{mol Li}}}$$

$$= 2.46 \times 10^{-5}\text{ g Li}$$

Using the Mole and Molar Mass Concepts

What is the mass of 1.28×10^8 atoms of Ne?

a. 4.29×10^{-15} g Ne
b. 5.35×10^{32} g Ne
c. 3.06×10^{-17} g Ne
d. 1.11×10^{31} g Ne

Solution Map

atoms Ne $\longrightarrow$ mol Ne $\longrightarrow$ grams Ne

Two conversion factors are needed:

$$\frac{1 \text{ mol Ne}}{20.18 \text{ g Ne}} \quad \text{and} \quad \frac{1 \text{ mol Ne}}{6.022 \times 10^{23} \text{ atoms Ne}}$$

Calculate

$$1.28 \times 10^8 \cancel{\text{atoms Ne}} \times \frac{1 \cancel{\text{mol Ne}}}{6.022 \times 10^{23} \cancel{\text{atoms Ne}}} \times \frac{20.18 \text{ g Ne}}{1 \cancel{\text{mol Ne}}} = 4.29 \times 10^{-15} \text{ g Ne}$$

Using the Mole and Molar Mass Concepts

What is the mass of 1.05 mol of Ag?

Solution Map **mol Ag $\longrightarrow$ grams Ag**

One conversion factor is needed:

$$\frac{1 \text{ mol Ag}}{107.9 \text{ g Ag}}$$

Calculate

$$1.05 \cancel{\text{mol Ag}} \times \frac{107.9 \text{ g Ag}}{1 \cancel{\text{mol Ag}}} = 113. \text{ g Ag}$$

Using the Mole and Molar Mass Concepts

What is the mass of 8.21 mol of K?

a. 321. g K
b. 2.10×10^{-2} g K
c. 113. g K
d. 1.11×10^{12} g K

Solution Map

mol K $\longrightarrow$ grams K

One conversion factor is needed:

$$\frac{1 \text{ mol K}}{39.10 \text{ g K}}$$

Calculate

$$8.21 \cancel{\text{mol K}} \times \frac{39.10 \text{ g K}}{1 \cancel{\text{mol K}}} = 321. \text{ g K}$$

Using the Mole and Molar Mass Concepts

How many hydrogen atoms are in 1.00 moles of H_2 molecules?

Solution Map **mol H_2 $\longrightarrow$ molecules H_2 $\longrightarrow$ atoms H_2**

Two conversion factors are needed:

$$\frac{1 \text{ mol } H_2}{6.022 \times 10^{23} \text{ molecules } H_2} \quad \text{and} \quad \frac{1 \text{ molecule } H_2}{2 \text{ atoms H}}$$

Calculate

$$1.00 \cancel{\text{mol } H_2} \times \frac{6.022 \times 10^{23} \cancel{\text{molecules } H_2}}{1 \cancel{\text{mol } H_2}} \times \frac{2 \text{ atoms H}}{1 \cancel{\text{molecule } H_2}} = 1.20 \times 10^{24} \text{ atoms } H_2$$

Using the Mole and Molar Mass Concepts

How many sulfur atoms are in 2.27 mol of S_8 molecules?

a. 1.33×10^{-23} atoms S_8
b. 7.53×10^{22} atoms S_8
c. 1.09×10^{25} atoms S_8
d. 2.08×10^{-25} atoms S_8

Solution Map

mol S_8 $\longrightarrow$ molecules S_8 $\longrightarrow$ atoms S_8

Two conversion factors are needed:

$$\frac{1 \text{ mol } S_8}{6.022 \times 10^{23} \text{ molecules } S_8} \quad \text{and} \quad \frac{1 \text{ molecule } S_8}{8 \text{ atoms S}}$$

Calculate

$$2.27 \cancel{\text{mol } S_8} \times \frac{6.022 \times 10^{23} \cancel{\text{molecules } S_8}}{1 \cancel{\text{mol } S_8}} \times \frac{8 \text{ atoms S}}{1 \cancel{\text{molecule } S_8}} = 1.09 \times 10^{25} \text{ atoms } S_8$$

Molar Mass of Compounds

Much like an element, molar mass can be defined for a compound.

Molar Mass (MM):
mass of one mole of the formula unit of a compound.

The molar mass of a compound is equal to the sum of the molar masses of all the atoms in the molecule.

Example H_2O

Molar Mass = $MM_O + 2MM_H = 16.00 \text{ g} + 2(1.008 \text{ g}) = 18.02 \text{ g}$

Molar Mass of Compounds

What is the molar mass of aluminum hydroxide, $Al(OH)_3$?

$Al(OH)_3$

Using the atomic masses of each element:

- **a. 43.99 g**
- **b. 78.00 g**
- **c. 75.99 g**
- **d. 46.00 g**

1 Al = 1(26.98 g) = 26.98 g
3 O = 3(16.00 g) = 48.00 g
3 H = 3(1.008 g) = 3.024 g
78.00 g

Molar Mass of Compounds

The molar mass of a compound contains Avogadro's number of formula units/molecules.

H_2	O	H_2O
2 x (6.022 x 10^{23}) H atoms	6.022 x 10^{23} O atoms	6.022 x 10^{23} H_2O molecules
2 mol H atoms	1 mol O atoms	1 mol H_2O molecules
2 x 1.008 g = 2.016 g H	16.00 g O	18.016 g H_2O

Reminder: Pay close attention to whether the desired unit involves atoms or formula units/molecules.

Example Cl_2

Contains 2 mol of Cl atoms but only 1 mol of Cl_2 molecules.

Molar Mass of Compounds

As for elements, we can use both the mol and molar mass of formula units/molecules as conversion factors.

How many moles of NaCl are there in 253 g of NaCl?

Solution Map **g NaCl ⟶ mol NaCl**

To convert between g of NaCl and moles, we must first calculate the molar mass of NaCl.

MM = 22.99 g + 35.45 g = 58.44 g NaCl

Calculate

$$253.\ \cancel{g\ NaCl} \times \frac{1\ mol\ NaCl}{58.44\ \cancel{g\ NaCl}} = 4.33\ mol\ NaCl$$

Molar Mass of Compounds

How many moles of $TiCl_4$ are there in 12.5 g of titanium(IV) chloride?

- **a. 0.0659 mol $TiCl_4$**
- **b. 0.0321 mol $TiCl_4$**
- **c. 2.37 x 10^3 mol $TiCl_4$**
- **d. 1.01 mol $TiCl_4$**

Solution Map

g $TiCl_4$ ⟶ mol $TiCl_4$

MM_{TiCl_4} = 47.87 g + 4(35.45) g = 189.7 g $TiCl_4$

Calculate

$$12.5\ \cancel{g\ TiCl_4} \times \frac{1\ mol\ TiCl_4}{189.7\ \cancel{g\ TiCl_4}} = 0.0659\ mol\ TiCl_4$$

Molar Mass of Compounds

What is the mass of 3.45 mol of Li_2O?

Solution Map **mol Li_2O ⟶ g Li_2O**

To convert between mol of Li_2O and g, we must first calculate the molar mass of Li_2O.

MM_{Li_2O} = 2(6.941) g + 16.00 g = 29.88 g Li_2O

Calculate

$$3.45\ \cancel{mol\ Li_2O} \times \frac{29.88\ g\ Li_2O}{1\ \cancel{mol\ Li_2O}} = 103.\ g\ Li_2O$$

Molar Mass of Compounds

What is the mass of 1.23 mol of PH_3?

- **a. 3.62 x 10^{-2} g PH_3**
- **b. 1.22 x 10^{-6} g PH_3**
- **c. 39.33 g PH_3**
- **d. 41.8 g PH_3**

Solution Map

mol PH_3 ⟶ g PH_3

MM_{PH_3} = 30.97 g + 3(1.008) g = 33.99 g PH_3

Calculate

$$1.23\ \cancel{mol\ PH_3} \times \frac{33.99\ g\ PH_3}{1\ \cancel{mol\ PH_3}} = 41.8\ g\ PH_3$$

Molar Mass of Compounds

How many molecules of H_2S are present in 7.53 g of H_2S? How many atoms of H are present in the sample?

Solution Map g $H_2S \rightarrow$ mol $H_2S \rightarrow$ molecules $H_2S \rightarrow$ atoms H

$$MM_{H_2S} = 2(1.008)\text{ g} + 32.07\text{ g} = 34.09\text{ g } H_2S$$

Calculate

$$7.53\ \cancel{\text{g } H_2S} \times \frac{1\ \cancel{\text{mol } H_2S}}{34.09\ \cancel{\text{g } H_2S}} \times \frac{6.022 \times 10^{23}\text{ molecules } H_2S}{1\ \cancel{\text{mol } H_2S}} = 1.33 \times 10^{23}\text{ molecules } H_2S$$

$$1.33 \times 10^{23}\ \cancel{\text{molecules } H_2S} \times \frac{2\text{ atoms H}}{1\ \cancel{\text{molecule } H_2S}} = 2.66 \times 10^{23}\text{ atoms H}$$

Molar Mass of Compounds

How many molecules of H_2O_2 are there in 0.759 g of the compound?

a. 4.29×10^{-23} molecules H_2O_2
b. 1.55×10^{25} molecules H_2O_2
c. 3.95×10^{24} molecules H_2O_2
d. 1.34×10^{22} molecules H_2O_2

Solution Map
g $H_2O_2 \rightarrow$ mol $H_2O_2 \rightarrow$ molecules H_2O_2

$$MM_{H_2O_2} = 2(1.008)\text{ g} + 2(16.00)\text{ g} = 34.02\text{ g } H_2O_2$$

$$0.759\ \cancel{\text{g } H_2O_2} \times \frac{1\ \cancel{\text{mol } H_2O_2}}{34.02\ \cancel{\text{g } H_2O_2}} \times \frac{6.022 \times 10^{23}\text{ molecules } H_2O_2}{1\ \cancel{\text{mol } H_2O_2}} = 1.34 \times 10^{22}\text{ molecules } H_2O_2$$

Percent Composition of Compounds

Percent = parts per 100 parts

Percent composition: mass percent of each element in a compound

Molar mass: total mass (100%) of a compound

% composition is independent of sample size

% composition can be determined by:
1. Knowing the compound's formula or
2. Using experimental data

Percent Composition from the Compound's Formula

Two Step Strategy

1. Calculate the molar mass of the compound.
2. Divide the total mass of each element by the compound's molar mass and multiply by 100.

$$\%\text{ of the element} = \frac{\text{Total element mass}}{\text{Compound molar mass}} \times 100$$

Percent Composition of Compounds

Calculate the percent composition of K_2S.

Step 1 **Calculate compound molar mass**

$$MM_{K_2S} = 2(39.10)\text{ g} + 32.07\text{ g} = 110.3\text{ g}$$

Step 2 **Calculate % composition of each element.**

$$\%\text{ K} = \frac{2(39.10)\text{ g K}}{110.3\text{ g}} \times 100 = 70.90\ \%\text{ K}$$

$$\%\text{ S} = \frac{32.07\text{ g S}}{110.3\text{ g}} \times 100 = 29.10\ \%\text{ S}$$

Notice the sum of the percentages must equal 100%. This provides another way of determining the % composition of a specific element, if the other %s are known.

Percent Composition of Compounds

Calculate the percent composition of O in H_2O_2.

a. 94.07 %
b. 5.93 %
c. 88.9 %
d. 11.1%

Step 1 **Calculate molar mass**

$$MM_{H_2O_2} = 2(1.008)\text{ g} + 2(16.00)\text{g} = 34.02\text{ g}$$

Step 2 **Calculate % composition O**

$$\%\text{ O} = \frac{2(16.00)\text{ g O}}{34.02\text{ g}} \times 100 = 94.07\ \%\text{ O}$$

Percent Composition of Compounds

Calculate the % composition of K_2CrO_4.

Step 1 Calculate compound molar mass

$$MM_{K_2CrO_4} = 2(39.10)\text{ g} + 52.00\text{ g} + 4(16.00)\text{ g} = 194.2\text{ g}$$

Step 2 Calculate % composition

$$\%\text{ K} = \frac{2(39.10)\text{ g K}}{194.2\text{ g}} \times 100 = 40.27\ \%\text{ K}$$

$$\%\text{ Cr} = \frac{52.00\text{ g Cr}}{194.2\text{ g}} \times 100 = 26.78\ \%\text{ Cr}$$

$$\%\text{ O} = \frac{4(16.00)\text{ g O}}{194.2\text{ g}} \times 100 = 32.95\ \%\text{ O}$$

Percent Composition from Experimental Data

Two Step Strategy

1. Calculate the mass of the compound formed.
2. Divide the mass of each element by the total mass and multiply by 100.

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

Percent Composition from Experimental Data

When heated in air, 1.63 g of Zn reacts with 0.40 g of oxygen to give ZnO. Calculate the percent composition of the compound formed.

Step 1 Calculate the mass of the compound formed.

Mass compound = Mass_{Zn} + Mass_{O} = 1.63 g + 0.40 g = 2.03 g compound

Step 2 Calculate % composition

$$\%\text{ Zn} = \frac{1.63\text{ g Zn}}{2.03\text{ g}} \times 100 = 80.3\ \%\text{ Zn}$$

$$\%\text{ O} = \frac{0.40\text{ g O}}{2.03\text{ g}} \times 100 = \underline{20.\ \%\text{ O}}$$

100.3%

Total should be +/0.5% of 100

Percent Composition from Experimental Data

Aluminum chloride forms by reaction of 13.43 g of Al with 53.18 g of chlorine. What is the percent composition of Cl in the compound?

a. 53.2 %
b. 79.8 %
c. 20.2 %
d. 46.8 %

Step 1 Calculate the mass of the compound formed.

Mass compound = Mass_{Al} + Mass_{Cl} = 13.43 g + 53.18 g = 66.61 g compound

Step 2 Calculate % composition

$$\%\text{ Cl} = \frac{53.18\text{ g Cl}}{66.61\text{ g}} \times 100 = 79.8\ \%$$

Empirical and Molecular Formula

Empirical Formula: smallest whole number ratio of atoms in a compound

Molecular Formula: actual formula of a compound. Represents the total number of atoms in one formula unit of the compound.

Whole number multiple of the empirical formula

Example **Acetylene (C_2H_2) and Benzene (C_6H_6)**

Both have the same empirical formula CH.

Each compound is a multiple of CH.

Acetylene $C_2H_2 = (CH)_2$ **Benzene $C_6H_6 = (CH)_6$**

Empirical and Molecular Formula

Formula	Composition % C	%H	Molar Mass (g/mol)
CH (empirical formula)	92.3	7.7	13.02
C_2H_2 (acetylene)	92.3	7.7	26.04 (2 x 13.02)
C_6H_6(benzene)	92.3	7.7	78.16 (6 x 13.02)

Each compound has very different chemical and physical properties even though they share the same empirical formula.

Compounds with the same empirical formula have the same percent composition.

Molar mass = molar mass of the empirical unit × multiple of the unit

Calculating Empirical Formulas

To calculate an empirical formula, you need to know:

1. The elements present in the compound

2. The atomic masses of each element (from the Periodic Table)

3. The ratio (by mass or %) of the combined elements

Calculating Empirical Formulas

Strategy to Calculate an Empirical Formula:

1. Assume a starting mass of the compound (usually 100.0 g) and express the mass of each element in grams.

2. Convert g of each element to mol using molar mass. (These numbers may or may not be whole numbers.)

3. Divide each of the mole amounts from Step 2 by the smallest mole amount. The new numbers are the subscripts in the empirical formula.

Special Case:

If fractions are encountered, multiply by a common factor to provide whole numbers for each subscript.

Calculating Empirical Formulas

Calculate the empirical formula for a compound that contains 11.19% H and 88.79% O.

Step 1 Find amounts of each element

In a 100.0 g sample, there are

11.19 g H and 88.79 g O

Step 2 Convert g to moles using element molar masses

$$11.19\ \cancel{g\ H} \times \frac{1\text{ mol H}}{1.008\ \cancel{g\ H}} = 11.10\text{ mol H}$$

$$88.79\ \cancel{g\ O} \times \frac{1\text{ mol O}}{16.00\ \cancel{g\ O}} = 5.549\text{ mol O}$$

Calculating Empirical Formulas

Step 3 Convert to whole numbers by dividing by the smallest mole amount.

$$\frac{11.10\text{ mol H}}{5.549\text{ mol O}} = 2.000$$

$$\frac{5.549\text{ mol O}}{5.549\text{ mol O}} = 1.000$$

Empirical formula is H_2O

Calculating Empirical Formulas

Calculate the empirical formula for a compound that contains 56.68% K, 8.68% C and 34.73% O.

a. $K_3C_2O_3$
b. $K_4C_2O_6$
c. K_2CO_3
d. KCO_2

Step 1 Find amounts of each element

In a 100.0 g sample, there are

56.68 g K, 8.68 g C and 34.73 g O

Step 2 Convert g to moles

$$56.69\ \cancel{g\ K} \times \frac{1\text{ mol K}}{39.10\ \cancel{g\ K}} = 1.447\text{ mol K}$$

$$8.68\ \cancel{g\ C} \times \frac{1\text{ mol C}}{12.01\ \cancel{g\ C}} = 0.723\text{ mol C}$$

$$34.73\ \cancel{g\ O} \times \frac{1\text{ mol O}}{16.00\ \cancel{g\ O}} = 2.171\text{ mol O}$$

Calculating Empirical Formulas

Calculate the empirical formula for a compound that contains 56.68% K, 8.68% C and 34.73% O.

Step 3 Convert to whole numbers by dividing by the smallest mole amount.

a. $K_3C_2O_3$
b. $K_4C_2O_6$
c. K_2CO_3
d. KCO_2

$$\frac{1.447\text{ mol K}}{0.723\text{ mol}} = 2.000$$

$$\frac{0.723\text{ mol C}}{0.723\text{ mol}} = 1.000$$

$$\frac{2.171\text{ mol O}}{0.723\text{ mol}} = 3.000$$

Empirical formula is: K_2CO_3

Calculating Empirical Formulas

Calculate the empirical formula for a compound that contains 2.233 g Fe and 1.926 g S?

a. FeS_2
b. Fe_3S_2
c. FeS
d. Fe_2S_3

Step 1 Find amounts of each element
Already provided in problem
2.233 g Fe and 1.926 S

Step 2 Convert g to moles

$$2.233\ \cancel{g\ Fe} \times \frac{1\ \text{mol Fe}}{55.85\ \cancel{g\ Fe}} = 0.03998\ \text{mol Fe}$$

$$1.926\ \cancel{g\ S} \times \frac{1\ \text{mol S}}{32.07\ \cancel{g\ S}} = 0.06006\ \text{mol S}$$

Calculating Empirical Formulas

Step 3 Convert to whole numbers by dividing by the smallest mole amount.

$$\frac{0.03998\ \text{mol Fe}}{0.03998\ \text{mol}} = 1.000 \times 2 = 2.000$$

$$\frac{0.06006\ \text{mol S}}{0.03998\ \text{mol}} = 1.502 \times 2 = 3.000$$

Empirical formula is Fe_2S_3

Common Fractions

Decimal	Fraction
0.25	1/4
0.33…	1/3
0.5	1/2
0.66….	2/3
0.75	3/4

To get a whole number, multiply the decimal by the corresponding number in the denominator of the fraction.
Example After dividing, you get 0.75 (=3/4)
Multiply by the denominator 4(0.75) = 4(3/4) = 3

Calculating the Molecular Formula from the Empirical Formula

If molar mass is known, the molecular formula can be calculated from the empirical formula.

Molecular formula is a multiple of the empirical formula. Need to determine the value of *n*.

Solving for *n*

$$n = \frac{\text{Molar mass}}{\text{Mass of empirical formula}} = \begin{array}{l}\text{number of empirical units}\\ \text{in the molecular formula}\end{array}$$

Calculating the Molecular Formula from the Empirical Formula

The molecular formula can be calculated from the empirical formula if the compound's molar mass is known.

Molecular formula = multiple of the empirical formula

$(EF)_n = MF$

Determining the multiple *n* gives the molecular formula

$$n = \frac{\text{Molar mass}}{\text{Mass of empirical formula}} = \begin{array}{l}\text{number of empirical units}\\ \text{in the molecular formula}\end{array}$$

Calculating the Molecular Formula from the Empirical Formula

A compound with the empirical formula NH_2 was found to have a molar mass of 32.05 g. What is the molecular formula?

$$n = \frac{\text{Molar mass}}{\text{Mass of empirical formula}} = \begin{array}{l}\text{number of empirical units}\\ \text{in the molecular formula}\end{array}$$

$$n = \frac{32.05}{14.01 + 2(1.008)} = 2$$

Molecular formula = $(NH_2)_2 = N_2H_4$

Calculating the Molecular Formula from the Empirical Formula

A compound with the empirical formula NO_2 was found to have a molar mass of 92.00 g. What is the molecular formula?

a) NO_2
b) N_2O_4
c) N_3O6
d) N_4O_8

$$n = \frac{92.00\ \text{g}}{14.01 + 2(16.00)\ \text{g}} = 2$$

Molecular formula = $(NO_2)_2 = N_2O_4$

Calculating the Molecular Formula from the Empirical Formula

Propylene contains 14.3 % H and 85.7 % C and has a molar mass of 42.08 g. What is its molecular formula?

Plan Calculate empirical formula and then determine the molecular formula

Step 1 Find compound masses

In 100.0 g of compound, 14.3 g H and 85.7 g C

Step 2 Convert g to moles

$$14.3\,\cancel{g\,H} \times \frac{1\text{ mol H}}{1.008\,\cancel{g\,H}} = 14.2\text{ mol H}$$

$$85.7\,\cancel{g\,C} \times \frac{1\text{ mol C}}{12.01\,\cancel{g\,C}} = 7.14\text{ mol C}$$

Calculating the Molecular Formula from the Empirical Formula

Step 3 Convert to whole numbers by dividing by the smallest mole amount.

$$\frac{14.2\text{ mol H}}{7.14\text{ mol}} = 1.99$$

$$\frac{7.14\text{ mol C}}{7.14\text{ mol}} = 1.00$$

Empirical formula = CH_2

With EF, calculate the molecular formula

$$n = \frac{42.08\text{ g}}{12.01 + 2(1.008)\text{ g}} = 3$$

Molecular formula = $(CH_2)_3 = C_3H_6$

Calculating the Molecular Formula from the Empirical Formula

Calculate the molecular formula for a compound that contains 80.0% C and 20.0% H with a molar mass of 30.00 g.

a. CH_3

b. CH_2

c. C_2H_6

d. C_2H_4

Plan Calculate empirical and then molecular formula

Step 1 Find compound masses

In 100.0 g of compound, 20.0 g H and 80.0 g C

Step 2 Convert g to moles

$$20.0\,\cancel{g\,H} \times \frac{1\text{ mol H}}{1.008\,\cancel{g\,H}} = 19.8\text{ mol H}$$

$$80.0\,\cancel{g\,C} \times \frac{1\text{ mol C}}{12.01\,\cancel{g\,C}} = 6.66\text{ mol C}$$

Calculating the Molecular Formula from the Empirical Formula

Step 3 Convert to whole numbers by dividing by the smallest mole amount.

$$\frac{19.8\text{ mol H}}{6.66\text{ mol}} = 2.97$$

$$\frac{6.66\text{ mol C}}{6.66\text{ mol}} = 1.00$$

Empirical formula = CH_3

From empirical formula, calculate the molecular formula

$$n = \frac{30.00\text{ g}}{12.01 + 3(1.008)\text{ g}} = 2$$

Molecular formula = $(CH_3)_2 = C_2H_6$

Learning Objectives

7.1 The Mole

Apply the concept of the mole, molar mass, and Avogadro's number to solve chemistry problems.

7.2 Molar Mass of Compounds

Calculate the molar mass of a compound.

7.3 Percent Composition of Compounds

Calculate the percent composition of a compound from its chemical composition and from experimental data.

Learning Objectives

7.4 Calculating Empirical Formulas

Determine the empirical formula for a compound from its percent composition.

7.5 Calculating the Molecular Formula from the Empirical Formula

Compare an empirical formula to a molecular formula and calculate a molecular formula from an empirical formula, using the molar mass.