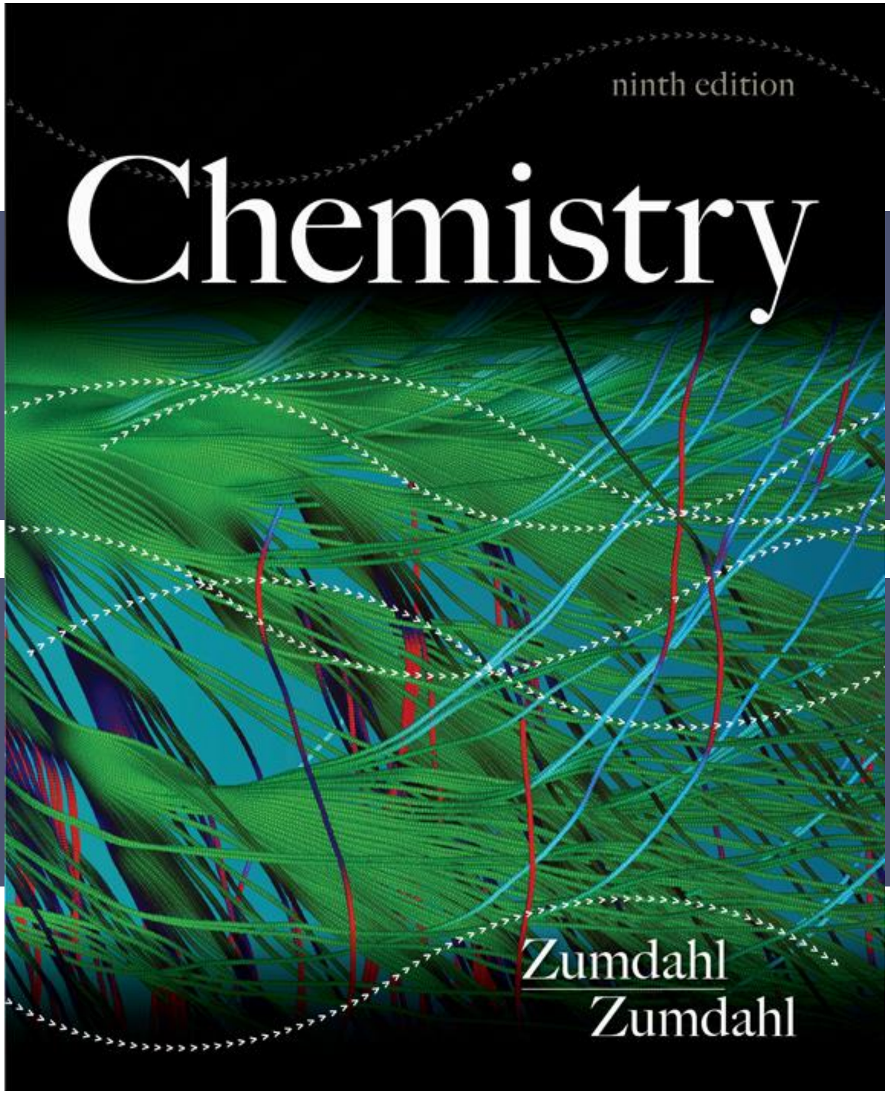


Chapter 3

Stoichiometry

Chemical Stoichiometry

- Stoichiometry – The study of quantities of materials consumed and produced in chemical reactions.

- Objects behave as though they were all identical.
- Atoms are too small to count.
- Need average mass of the object.

EXERCISE!

$10 \rightarrow 37.60$
$x \rightarrow 394.80$

$=$

$3948 = x\,37.60$

$x = \dfrac{3948}{37.60} = 105.0$

A pile of marbles weigh 394.80 g. 10 marbles weigh 37.60 g. How many marbles are in the pile?

$$\text{Avg. Mass of 1 Marble} = \frac{37.60\text{ g}}{10\text{ marbles}} = 3.76\text{ g / marble}$$

$$\frac{394.80\text{ g}}{3.76\text{ g}} = 105\text{ marbles}$$

- Atomic mass unit: a mass unit equal exactly (1/12) the mass of carbon-12
- ^{12}C is the standard for atomic mass, with a mass of exactly 12 atomic mass units (u).
- The masses of all other atoms are given relative to this standard.
- Elements occur in nature as mixtures of isotopes.
- Carbon = 98.89% ^{12}C
  1.11% ^{13}C
  < 0.01% ^{14}C

Average Atomic Mass for Carbon

98.89% of 12 u + 1.11% of 13.0034 u =

exact number

(0.9889)(12 u) + (0.0111)(13.0034 u) =

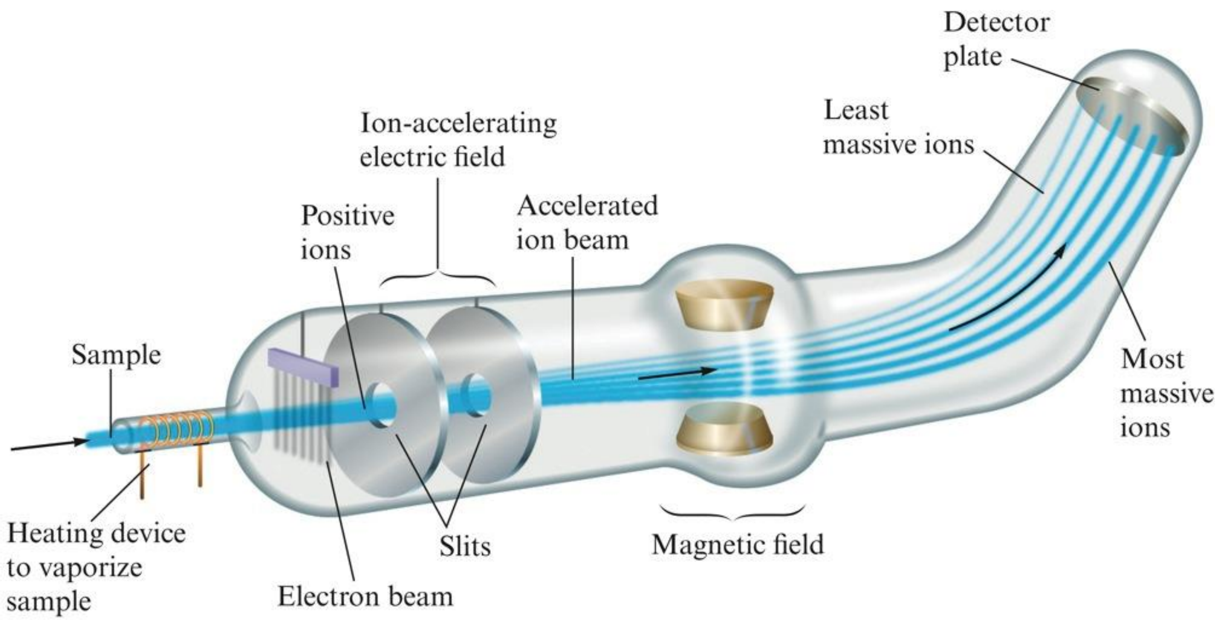
12.01 u

Average Atomic Mass for Carbon

- Even though natural carbon does not contain a single atom with mass 12.01, for stoichiometric purposes, we can consider carbon to be composed of only one type of atom with a mass of 12.01.
- This enables us to count atoms of natural carbon by weighing a sample of carbon.

Schematic Diagram of a Mass Spectrometer

Geoff Tompkinson/Photo Researchers, Inc.



Exercise 2.2 Chlorine consists of the following isotopes:

Isotope	Isotopic Mass (amu)		Fractional Abundance
Chlorine-35	34.96885	x	0.75771
Chlorine-37	36.96590	x	0.24229

x

What is the atomic weight of chlorine?

(34.96885 x 0.75771) + (36.96590 x 0.24229)
=35.45 amu

Mass number	Isotopic mass (amu)	Fractional abundance
Cr-50	49.9461	0.0435
Cr-52	51.9405	0.8379
Cr-53	52.9407	0.0950
Cr-54	53.9389	0.0236

What is the atomic weigh of chromium?

EXERCISE!

An element consists of 62.60% of an isotope with mass 186.956 u and 37.40% of an isotope with mass 184.953 u.

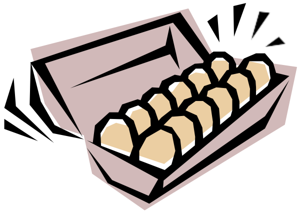
- Calculate the average atomic mass and identify the element.

186.2 u

Rhenium (Re)

3.2 The Mole Concept

Dozen = 12



The Mole (mol): A unit to count numbers of particles

Pair = 2

A mole (symbol mol) is defined as *the quantity of a given substance that contains as many molecules or formula units as the number of atoms in exactly 12 g of carbon-12*

1 mol = N_A = 6.0221415 x 10^{23} = Avogadro's number

1 mole of Na_2CO_3 contains 6.02 x 10^{23} Na_2CO_3 units 1 mole of Na_2CO_3 contains 2x 6.02 x 10^{23} Na^+ ions 1 mole of Na_2CO_3 contains 6.02 x 10^{23} CO_3^{2-} ions

molar mass of a substance is *the mass of one mole of the substance.*

C has a molar mass of exactly 12 g/mol,

C_2H_5OH has a molar mass of exactly 46.1 g/mol

- The number equal to the number of carbon atoms in exactly 12 grams of pure ^{12}C.
- 1 mole of something consists of 6.022×10^{23} units of that substance (Avogadro's number).
- 1 mole C = 6.022×10^{23} C atoms = 12.01 g C

For any element

atomic mass (amu) = molar mass (grams)

- Mass in grams of one mole of the substance:

  Molar Mass of N = 14.01 g/mol

  Molar Mass of H_2O = 18.02 g/mol

  $(2 \times 1.008\ \text{g}) + 16.00\ \text{g}$

  Molar Mass of $Ba(NO_3)_2$ = 261.35 g/mol

  $137.33\ \text{g} + (2 \times 14.01\ \text{g}) + (6 \times 16.00\ \text{g})$

EXERCISE!

Calculate the number of iron atoms in a 4.48 mole sample of iron.

2.70×10^{24} Fe atoms

CONCEPT CHECK!

Calculate the number of copper atoms in a 63.55 g sample of copper.

6.022×10^{23} Cu atoms

(Q) How much, in grams, do 8.85 × 10^{24} atoms of zinc weigh?

A. 3.49 × 10^{49} g

B. 961 g

C. 4.45 g

D. 5.33 × 10^{47} g

E. 1.47 g

$$8.85 \times 10^{24} \times \frac{1\ \cancel{\text{mol}}}{6.022 \times 10^{23}} \times \frac{65.41\text{ g Zn}}{1\ \cancel{\text{mol}}}$$

= 961 g of Zn

$$8.85\times10^{24}\cancel{\text{atoms}}\times\left(\frac{1\ \cancel{\text{mol}}}{6.022\times10^{23}\ \cancel{\text{atoms}}}\right)\times\left(\frac{65.41\text{ g Zn}}{1\ \cancel{\text{mol}}}\right)$$

?

= 961 g Zn

CONCEPT CHECK!

Atomic Mass

Which of the following is closest to the average mass of one atom of copper?

Average weight of Isotopes

a) 63.55 g
b) 52.00 g
c) 58.93 g
d) 65.38 g
e) $\boxed{1.055 \times 10^{-22} \text{ g}}$

Example 3.3 Calculating the Mass of an Atom or Molecule

a. What is the mass in grams of one chlorine atom, Cl?

b. What is the mass in grams of one HCl molecule?

Solution

a. The atomic weight of Cl is 35.5 amu, so the molar mass of Cl is 35.5 g/mol. Dividing 35.5 g (per mole) by 6.02×10^{23} (Avogadro's number) gives the mass of one atom.

$$\text{Mass of a Cl atom} = \frac{35.5\ \text{g}}{6.02 \times 10^{23}} = \mathbf{5.90 \times 10^{-23}\ g}$$

b. The molecular weight of HCl equals the AW of H plus the AW of Cl, or 1.01 amu + 35.5 amu = 36.5 amu. Therefore, 1 mol HCl contains 36.5 g HCl and

$$\text{Mass of an HCl molecule} = \frac{36.5\ \text{g}}{6.02 \times 10^{23}} = \mathbf{6.06 \times 10^{-23}\ g}$$

A silicon chip used in an integrated circuit of a microcomputer has a mass of 5.68 mg. How many silicon (Si) atoms are present in the chip?

$$5.68\ \cancel{\text{mg Si}} \times \frac{1\ \text{g Si}}{1000\ \cancel{\text{mg Si}}} = 5.68 \times 10^{-3}\ \text{g Si}$$

$$5.68 \times 10^{-3}\ \cancel{\text{g Si}} \times \frac{1\ \text{mol Si}}{28.09\ \cancel{\text{g Si}}} = 2.02 \times 10^{-4}\ \text{mol Si}$$

$$2.02 \times 10^{-4}\ \cancel{\text{mol Si}} \times \frac{6.022 \times 10^{23}\ \text{atoms}}{1\ \cancel{\text{mol Si}}} = 1.22 \times 10^{20}\ \text{atoms}$$

(Q) How many S atoms are there in 16.3 g of S?

$$16.3\ \cancel{\text{g}} \times \frac{1\ \cancel{\text{mol}}}{32.065\ \cancel{\text{g}}} \times \frac{6.022 \times 10^{23}\ \text{atom}}{1\ \cancel{\text{mol}}}$$

$$= 3.06 \times 10^{23}\ \text{atoms}$$

g $\rightarrow$ mol x Atomic Mass = الجواب

- **Mole Calculations**

(Q) A chemist determines from the amounts of elements that 0.0654 mol ZnI_2 can form. How many grams of zinc iodide is this? molar mass of ZnI_2 is 319 g/mol

Number of moles = mass(g) / molar mass

(Q)In a preparation rxn., 45.6 g of lead(II) chromate is obtained as a precipitate. How many moles of $PbCrO_4$ is this?
molar mass of $PbCrO_4$ = 323 g/mol

(Q) How many hydrogen atoms are present in 25.6 g of urea $[(NH_2)_2CO]$? molar mass of urea = 60.06 g/mol.

grams of urea $\longrightarrow$ moles of urea $\longrightarrow$ moles of H $\longrightarrow$ atoms of H

$$25.6\ \cancel{\text{g }(NH_2)_2CO} \times \frac{1\ \cancel{\text{mol }(NH_2)_2CO}}{60.06\ \cancel{\text{g }(NH_2)_2CO}} \times \frac{4\ \cancel{\text{mol H}}}{1\ \cancel{\text{mol }(NH_2)_2CO}} \times \frac{6.022\times10^{23}\ \text{H atoms}}{1\ \cancel{\text{mol H}}}$$

$= 1.03 \times 10^{24}$ H atoms

(Q) Calculate the number of moles of calcium in 2.53 moles of $Ca_3(PO_4)_2$

A. 2.53 mol Ca

B. 0.432 mol Ca

C. 3.00 mol Ca

D. 7.59 mol Ca

E. 0.843 mol Ca

2.53 moles of $Ca_3(PO_4)_2$ = ? mol Ca

3 mol Ca $\Leftrightarrow$ 1 mol $Ca_3(PO_4)_2$

$$2.53\,\text{mol}\,\cancel{Ca_3(PO_4)_2}\left(\frac{3\,\text{mol Ca}}{1\,\text{mol}\,\cancel{Ca_3(PO_4)_2}}\right)$$

= 7.59 mol Ca

(Q) A sample of sodium carbonate, Na_2CO_3, is found to contain 10.8 moles of sodium. How many moles of oxygen atoms (O) are present in the sample?

A. 10.8 mol O

B. 7.20 mol O

C. 5.40 mol O

D. 32.4 mol O

E. 16.2 mol O

2 mol Na ⇔ 3 mol O

10.8 moles of Na = ? mol O

$$10.8\,\cancel{\text{mol Na}}\left(\frac{3\,\text{mol O}}{2\,\cancel{\text{mol Na}}}\right)$$

= 16.2 mol O

(Q) How many g of iron are required to use up all of 25.6 g of oxygen atoms (O) to form Fe_2O_3?

A. 59.6 g

B. 29.8 g

C. 89.4 g

D. 134 g

E. 52.4 g

mass O → mol O → mol Fe → mass Fe

3 mol O ⇔ 2 mol Fe

25.6 g O → ? g Fe

$$25.6\,\cancel{\text{g O}} \times \left(\frac{1\,\cancel{\text{mol O}}}{16.0\,\cancel{\text{g O}}}\right) \times \left(\frac{2\,\cancel{\text{mol Fe}}}{3\,\cancel{\text{mol O}}}\right) \times \left(\frac{55.845\,\text{g Fe}}{1\,\cancel{\text{mol Fe}}}\right)$$

= 59.6 g Fe

(Q) Silver is often found in nature as the ore, argentite (Ag_2S). How many grams of pure silver can be obtained from a 836 g rock of argentite?

mass Ag_2S → mol Ag_2S → mol Ag → mass Ag

A. 7.75 g

B. 728 g

C. 364 g

D. 775 g

E. 418 g

1 mol Ag_2S ⇔ 2 mol Ag

836 g Ag_2S → ? g Ag

$$836\,\text{g}\,\cancel{\text{Ag}_2\text{S}} \times \left(\frac{1\,\text{mol}\,\cancel{\text{Ag}_2\text{S}}}{247.8\,\text{g}\,\cancel{\text{Ag}_2\text{S}}}\right) \times \left(\frac{2\,\cancel{\text{mol}\,\text{Ag}}}{1\,\text{mol}\,\cancel{\text{Ag}_2\text{S}}}\right) \times \left(\frac{107.9\,\text{g}\,\text{Ag}}{1\,\cancel{\text{mol}\,\text{Ag}}}\right)$$

= 728 g Ag

CONCEPT CHECK!

Which of the following 100.0 g samples contains the greatest number of atoms?

$$100\,g \times \frac{1\,mol}{1\,g} \times \frac{6.022 \times 10^{23}}{1\,mol}$$

a) Magnesium
b) Zinc
c) Silver

EXERCISE!

Rank the following according to number of atoms (greatest to least):

a) 107.9 g of silver
b) 70.0 g of zinc
c) 21.0 g of magnesium

b) a) c)

EXERCISE!

Consider separate 100.0 gram samples of each of the following:

H_2O, N_2O, $C_3H_6O_2$, CO_2

- Rank them from greatest to least number of oxygen atoms.

H_2O, CO_2, $C_3H_6O_2$, N_2O

Conceptual Problem Solving

- Where are we going?
  - Read the problem and decide on the final goal.
- How do we get there?
  - Work backwards from the final goal to decide where to start.
- Reality check.
  - Does my answer make sense? Is it reasonable?

- Mass percent of an element:

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

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

$$\text{mass \% Fe} = \frac{2(55.85\text{ g})}{2(55.85\text{ g})+3(16.00\text{ g})} = \frac{111.70\text{ g}}{159.70\text{ g}} \text{ x } 100\% = 69.94\%$$

Percentage Composition

Example:

A sample of a liquid with a mass of 8.657 g was decomposed into its elements and gave 5.217 g of carbon, 0.9620 g of hydrogen, and 2.478 g of oxygen. What is the percentage composition of this compound?

5.217 + 0.9620 + 2.478 = 8.657g

$$\text{Mass C \%} = \frac{5.217\,\text{g C}}{8.657\,\text{g}} \times 100\% = 60.26\%\ \text{C}$$

$$\text{Mass H \%} = \frac{0.9620\,\text{g H}}{8.657\,\text{g}} \times 100\% = 11.11\%\ \text{H}$$

$$\text{Mass O \%} = \frac{2.478\,\text{g O}}{8.657\,\text{g}} \times 100\% = 28.62\%\ \text{O}$$

Sum of percentages: 100 %

(Q) A sample was analyzed and found to contain 0.1417 g nitrogen and 0.4045 g oxygen.

What is the percentage composition of this compound?

Total sample mass = 0.1417 g + 0.4045 g = 0.5462 g

% Composition of N

$$= \left(\frac{0.1417\,\text{g N}}{0.5462\,\text{g}} \right) \times 100\% \quad \mathbf{= 25.94\%\ N}$$

% Composition of O

$$= \left(\frac{0.4045\,\text{g O}}{0.5462\,\text{g}} \right) \times 100\% \quad \mathbf{= 74.06\%\ O}$$

(Q) a. Calculate the mass percentages of the elements in formaldehyde (CH_2O) molar mass = 30g/mol

$$\% \text{ C} = \frac{12.0 \text{ g}}{30.0 \text{ g}} \times 100\% = \mathbf{40.0\%}$$

$$\% \text{ H} = \frac{2 \times 1.01 \text{ g}}{30.0 \text{ g}} \times 100\% = \mathbf{6.73\%}$$

$$\% \text{ O} = 100\% - (40.0\% + 6.73\%) = \mathbf{53.3\%}$$

% O = 16/30 x100% = 53.3 %

b. How many grams of carbon are there in 83.5 g of CH_2O?

CH_2O is 40.0% C, so the mass of carbon in 83.5 g CH_2O is:
83.5 g X 0.400 = **33.4 g**

(Q) Calculate the mass percentages of the elements in H_3PO_4

molar mass = 97.99 g/mol

$$\%H = \frac{3(1.008\ g)\ H}{97.99\ g\ H_3PO_4} \times 100\% = 3.086\%$$

$$\%P = \frac{30.97\ g\ P}{97.99\ g\ H_3PO_4} \times 100\% = 31.61\%$$

$$\%O = \frac{4(16.00\ g)\ O}{97.99\ g\ H_3PO_4} \times 100\% = 65.31\%$$

EXERCISE!

Consider separate 100.0 gram samples of each of the following:

H_2O, N_2O, $C_3H_6O_2$, CO_2

- Rank them from highest to lowest percent oxygen by mass.

H_2O, CO_2, $C_3H_6O_2$, N_2O

- Phenol, commonly known as carbolic acid, was used by Joseph Lister as an antiseptic for surgery in 1865. Its principal use today is in the manufacture of phenolic resins and plastics. Combustion of 5.23 mg of phenol yields 14.67 mg CO_2 and 3.01 mg H_2O. Phenol contains only C, H, and O. What is the percentage of each element in this substance?

$$14.67\text{ mg } CO_2 \times \frac{1\text{ mol } CO_2}{44.01\text{ g } CO_2} \times \frac{12.01\text{ g C}}{1\text{ mol C}} = 4.00\underline{3}3\text{ mg C}$$

$$3.01\text{ mg } H_2O \times \frac{1\text{ mol } H_2O}{18.02\text{ g } H_2O} \times \frac{2\text{ H}}{1\text{ } H_2O} \times \frac{1.008\text{ g H}}{1\text{ mol H}} = 0.33\underline{6}8\text{ mg H}$$

$$\text{Mass O} = 5.23\text{ mg} - (4.0033 + 0.3368) = 0.8\underline{8}99\text{ mg O}$$

$$\text{Percent C} = (4.0033\text{ mg}/5.23\text{ mg}) \times 100\% = 76.\underline{5}4 = 76.5\%$$

$$\text{Percent H} = (0.3368\text{ mg}/5.23\text{ mg}) \times 100\% = 6.4\underline{3}9 = 6.44\%$$

$$\text{Percent O} = (0.8899\text{ mg}/5.23\text{ mg}) \times 100\% = 1\underline{7}.0 = 17\%$$

- **Determining Empirical and Molecular Formulas** **Empirical Formula**
  - Simplest ratio of atoms of each element in compound
  - Obtained from experimental analysis of compound

Molecular Formula
  - Exact composition of one molecule
  - Exact whole number ratio of atoms of each element in molecule

Formulas

- Empirical formula = CH
  - Simplest whole-number ratio
- Molecular formula = (empirical formula)$_n$
  [n = integer]
- Molecular formula = C_6H_6 = $(CH)_6$
  - Actual formula of the compound

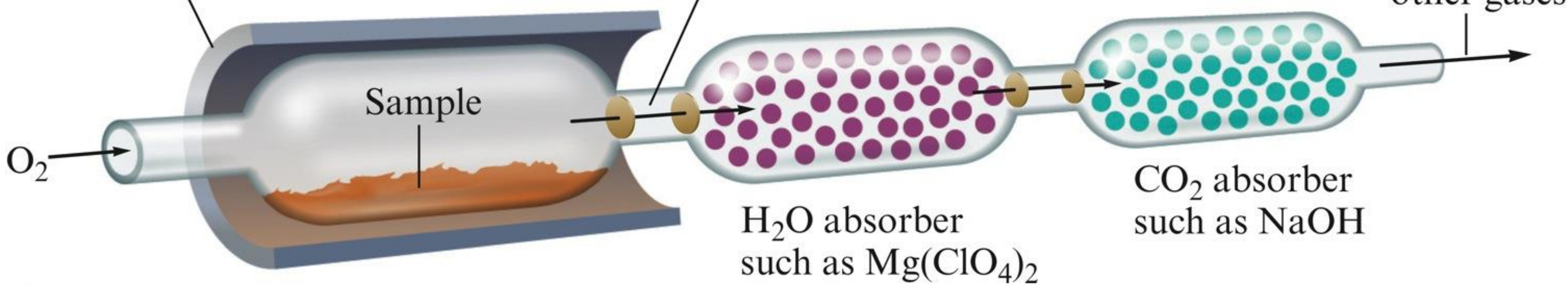
glucose

Molecular formula $C_6H_{12}O_6$

Empirical formula CH_2O

Analyzing for Carbon and Hydrogen

- Device used to determine the mass percent of each element in a compound.



➢ Three Ways to Calculate Empirical Formulas

1. From Masses of Elements

e.g., 2.448 g sample of which 1.771 g is Fe and 0.677 g is O.

2. **From Percentage Composition** **e.g.,** 43.64% P and 56.36% O

احتراق

3. From Combustion Data

- Given masses of combustion products

e.g., The combustion of a 5.217 g sample of a compound of C, H, and O in pure oxygen gave

7.406 g CO_2 and 4.512 g of H_2O

1. Empirical Formula from Mass Data

When a 0.1156 g sample of a compound was analyzed, it was found to contain 0.04470 g of C, 0.01875 g of H, and 0.05215 g of N. Calculate the empirical formula of this compound.

Step 1: Calculate moles of each substance

$$0.04470\,\cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.011\,\cancel{\text{g C}}} = 3.722 \times 10^{-3}\ \text{mol C}$$

$$0.01875\,\cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.008\,\cancel{\text{g H}}} = 1.860 \times 10^{-2}\ \text{mol H}$$

$$0.05215\,\cancel{\text{g N}} \times \frac{1\ \text{mol N}}{14.0067\,\cancel{\text{g N}}} = 3.723 \times 10^{-3}\ \text{mol N}$$

Step 2: Select the smallest number of moles

- Smallest is 3.722×10^{-3} mole

	Mole ratio	Integer ratio
$C = \dfrac{3.722 \times 10^{-3}\ \text{mol C}}{3.722 \times 10^{-3}\ \text{mol C}} =$	1.000	= 1
$H = \dfrac{1.860 \times 10^{-2}\ \text{mol H}}{3.722 \times 10^{-3}\ \text{mol C}} =$	4.997	= 5
$N = \dfrac{3.723 \times 10^{-3}\ \text{mol N}}{3.722 \times 10^{-3}\ \text{mol C}} =$	1.000	= 1

Step 3: Divide all number of moles by the smallest one

Empirical formula = CH_5N

2. Empirical Formula from Percentage Composition

Calculate the empirical formula of a compound whose percentage composition data is 43.64% P and 56.36% O. If the molar mass is determined to be 283.9 g/mol, what is the empirical formula and molecular formula?

Step 1: <u>Assume 100 g of compound</u>

- 43.64 g P — 1 mol P = 30.97 g
- 56.36 g O — 1 mol O = 16.00 g

$$43.64\,\cancel{\text{g P}} \times \frac{1\,\text{mol P}}{30.97\,\cancel{\text{g P}}} = 1.409\ \text{mol P}$$

$$56.36\,\cancel{\text{g O}} \times \frac{1\,\text{mol O}}{16.00\,\cancel{\text{g O}}} = 3.523\ \text{mol P}$$

Step 2: Divide by smallest number of moles

$$\frac{1.409\,\text{mol P}}{1.409\,\text{mol P}} = 1.000$$

$$\frac{3.523\,\text{mol O}}{1.409\,\text{mol P}} = 2.500$$

Step 3: Multiple to get integers

$$1.000 \times 2 = 2$$
$$2.500 \times 2 = 5$$

Empirical formula = P_2O_5

$$\text{Molecular formula} = \frac{\text{Molar mass of sample}}{\text{Molar mass of empirical formula}}$$

$$n = \frac{283.88}{141.94} = 2$$

(Empirical formula)n = molecular formula

$$(P_2O_5)_2 \longrightarrow P_4O_{10}$$

(Q) Ascorbic acid (vitamin C) is composed of 40.92 percent carbon (C), 4.58 percent hydrogen (H), and 54.50 percent oxygen (O) by mass. Determine its empirical formula.

Assume you have 100 g.

$$n_C = 40.92\ \cancel{g\ C} \times \frac{1\ \text{mol C}}{12.01\ \cancel{g\ C}} = 3.407\ \text{mol C}$$

$$n_H = 4.58\ \cancel{g\ H} \times \frac{1\ \text{mol H}}{1.008\ \cancel{g\ H}} = 4.54\ \text{mol H}$$

→formula $C_{3.407}H_{4.54}O_{3.406}$

$$n_O = 54.50\ \cancel{g\ O} \times \frac{1\ \text{mol O}}{16.00\ \cancel{g\ O}} = 3.406\ \text{mol O}$$

$$C: \frac{3.407}{3.406} \approx 1 \qquad H: \frac{4.54}{3.406} = 1.33 \qquad O: \frac{3.406}{3.406} = 1$$

→formula $C_1H_{1.33}O_1$ X 3 → formula $C_3H_4O_3$

Hydroquinone, used as aphotographic developer, is 65.4% C, 5.5% H, and 29.1% O, by mass. What is the empirical formula of hydroquinone?

Assume a sample of 100.0 g of hydroquinone

$$\text{Mol C} = 65.4 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 5.445 \text{ mol}$$

$$\text{Mol H} = 5.5 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 5.46 \text{ mol}$$

$$\text{Mol O} = 29.1 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.819 \text{ mol}$$

Integer for C = 5.445 ÷ 1.819 = 2.99, or 3

Integer for H = 5.46 ÷ 1.819 = 3.0, or 3

Integer for O = 1.819 ÷ 1.819 = 1.00, or 1

The empirical formula is thus C_3H_3O.

Compounds of boron with hydrogen are called boranes. One of these boranes has the empirical formula BH_3 and a molecular weight of 28 amu. What is its molecular formula?

Empirical Formula mass = 10.81 amu + (3 x 1.008 amu) = 13.83 amu

$$\text{Molecular formula (n)} = \frac{\text{Molar mass of sample}}{\text{Molar mass of empirical formula}}$$

n =28 amu $\div$ 13.83 amu = 2.02

(Empirical formula)n = molecular formula

$(BH_3)_2$, or B_2H_6.

- Carbon dioxide and water are separated and weighed separately
  - All C ends up as CO_2
  - All H ends up as H_2O
  - Mass of C can be derived from amount of CO_2
    - mass $CO_2 \rightarrow$ mol $CO_2 \rightarrow$ mol C $\rightarrow$ mass C
  - Mass of H can be derived from amount of H_2O
    - mass $H_2O \rightarrow$ mol $H_2O \rightarrow$ mol H $\rightarrow$ mass H
  - **Mass of oxygen is obtained by difference**
- **mass O = mass sample – (mass C + mass H)**

(Q) The combustion of a 5.217 g sample of a compound of C, H, and O in pure oxygen gave 7.406 g CO_2 and 4.512 g of H_2O. Calculate the empirical formula of the compound.

1. Calculate mass of C from mass of CO_2.

mass $CO_2 \rightarrow$ mole $CO_2 \rightarrow$ mole C $\rightarrow$ mass C

$$7.406\,\mathrm{g\,CO_2}\left(\frac{1\ \mathrm{mol\,CO_2}}{44.01\,\mathrm{g\,CO_2}}\right)\left(\frac{1\,\mathrm{mol\,C}}{1\,\mathrm{mol\,CO_2}}\right)\left(\frac{12.011\,\mathrm{g\,C}}{1\ \mathrm{mol\,C}}\right) \quad \mathbf{= 2.021\ g\ C}$$

2. Calculate mass of H from mass of H_2O.

mass $H_2O \rightarrow$ mol $H_2O \rightarrow$ mol H $\rightarrow$ mass H

$$4.512\,\mathrm{g\,H_2O}\left(\frac{1\,\mathrm{mol\,H_2O}}{18.015\,\mathrm{g\,H_2O}}\right)\left(\frac{2\,\mathrm{mol\,H}}{1\,\mathrm{mol\,H_2O}}\right)\left(\frac{1.008\,\mathrm{g\,H}}{1\,\mathrm{mol\,H}}\right) \quad \mathbf{= 0.5049\ g\ H}$$

3. Calculate mass of O from difference.

5.217 g sample – 2.021 g C – 0.5049 g H **= 2.691 g O**

4. Calculate mol of each element

$$\text{mol C} = \frac{\text{g C}}{\text{MM C}} = \frac{2.021\text{g}}{12.011\text{g/mol}} \qquad = 0.1683 \text{ mol C}$$

$$\text{mol H} = \frac{\text{g H}}{\text{MM H}} = \frac{0.5049\text{g}}{1.008\text{g/mol}} \qquad = 0.5009 \text{ mol H}$$

$$\text{mol O} = \frac{\text{g O}}{\text{MM O}} = \frac{2.691\text{g}}{15.999\text{g/mol}} \qquad = 0.1682 \text{ mol O}$$

- Preliminary empirical formula

  $C_{0.1683}H_{0.5009}O_{0.1682}$

5. Calculate mol ratio of each element

divide by the smallest mole.

$$C_{\frac{0.1683}{0.1682}}H_{\frac{0.5009}{0.1682}}O_{\frac{0.1682}{0.1682}} = C_{1.00}H_{2.97}O_{1.00}$$

Empirical Formula = CH_3O

(Q) The combustion of a 13.660 g sample of a compound of C, H, and S in pure oxygen gave 19.352 g CO_2 and 11.882 g of H_2O. Calculate the empirical formula of the compound.

A. $C_4H_{12}S$
B. CH_3S
C. C_2H_6S
D. $C_2H_6S_3$
E. CH_3S_2

(1) mass CO_2 → mole CO_2 → mole C → mass C
(2) mass H_2O → mole H_2O → mole H → mass H
(3) Calculate mass of S from difference

$$19.352\,\text{g CO}_2\left(\frac{1\ \text{mol CO}_2}{44.01\,\text{g CO}_2}\right)\left(\frac{1\,\text{mol C}}{1\,\text{mol CO}_2}\right)\left(\frac{12.011\,\text{g C}}{1\ \text{mol C}}\right) = 5.281\text{ g C}$$

$$11.882\,\text{g H}_2\text{O}\left(\frac{1\,\text{mol H}_2\text{O}}{18.015\,\text{g H}_2\text{O}}\right)\left(\frac{2\,\text{mol H}}{1\,\text{mol H}_2\text{O}}\right)\left(\frac{1.008\,\text{g H}}{1\,\text{mol H}}\right) = 1.330\text{ g H}$$

13.66 g sample – 5.281 g C – 1.330 g H = 7.049 g S

$$\text{mol C} = \frac{\text{g C}}{\text{MM C}} = \frac{5.281\,\cancel{\text{g}}}{12.011\,\cancel{\text{g}}/\text{mol}} = 0.4497 \text{ mol C}$$

$$\text{mol H} = \frac{\text{g H}}{\text{MM H}} = \frac{1.330\,\cancel{\text{g}}}{1.008\,\cancel{\text{g}}/\text{mol}} = 1.319 \text{ mol H}$$

$$\text{mol S} = \frac{\text{g S}}{\text{MM S}} = \frac{7.049\,\cancel{\text{g}}}{32.065\,\cancel{\text{g}}/\text{mol}} = 0.2198 \text{ mol S}$$

- Preliminary empirical formula
  - $C_{0.4497}H_{1.319}S_{0.2198}$

$$C_{\frac{0.4497}{0.2198}}H_{\frac{1.319}{0.2198}}S_{\frac{0.2198}{0.2198}} = C_{2.03}H_{6.00}S_{1.00}$$

Empirical Formula = C_2H_6S

- Determining Molecular Formula from emirical formula

-In some cases molecular and empirical formulas are the same

-When they are different, the subscripts of molecular formula are integer multiples of those in empirical formula

- If empirical formula is A_xB_y
- Molecular formula will be $A_{(n\times x)}\ B_{(n\times y)}$

(Q)The empirical formula of hydrazine is NH_2, and its molecular mass is 32.0. What is its molecular formula?

Atomic masses: N = 14.007; H = 1.008; O = 15.999

Molar mass of NH_2 = (1 × 14.01) g + (2 × 1.008) g = 16.017 g

n = (32.0/16.02) = 2 (NH_2) X 2 = N_2H_4

EXERCISE!

The composition of adipic acid is 49.3% C, 6.9% H, and 43.8% O (by mass). The molar mass of the compound is about 146 g/mol.

- What is the empirical formula?

  $C_3H_5O_2$

- What is the molecular formula?

  $C_6H_{10}O_4$

- A representation of a chemical reaction:

$$C_2H_5OH + 3O_2 \longrightarrow 2CO_2 + 3H_2O$$

reactants products

- Reactants are only placed on the left side of the arrow, products are only placed on the right side of the arrow.

1 + 6 = 7 = 4 + 3

$$C_2H_5OH + 3O_2 \longrightarrow 2CO_2 + 3H_2O$$

- The equation is balanced.
- All atoms present in the reactants are accounted for in the products.
- 1 mole of ethanol reacts with 3 moles of oxygen to produce 2 moles of carbon dioxide and 3 moles of water.

- The balanced equation represents an overall ratio of reactants and products, not what actually "happens" during a reaction.
- Use the coefficients in the balanced equation to decide the amount of each reactant that is used, and the amount of each product that is formed.

Writing and Balancing the Equation for a Chemical Reaction

1. Determine what reaction is occurring. What are the reactants, the products, and the physical states involved?
2. Write the *unbalanced* equation that summarizes the reaction described in step 1.
3. Balance the equation by inspection, starting with the most complicated molecule(s). The same number of each type of atom needs to appear on both reactant and product sides. Do NOT change the formulas of any of the reactants or products.

(a) $4 H_3PO_3 \rightarrow 3 H_3PO_4 + PH_3$

(b) $Ca + 2 H_2O \rightarrow Ca(OH)_2 + H_2$

(c) $Fe_2(SO_4)_3 + 6 NH_3 + 6 H_2O \rightarrow 2Fe(OH)_3 + 3 (NH_4)_2SO_4$

- Find the coefficients that balance the following equations.

a. $1/2 O_2 + PCl_3 \rightarrow POCl_3$

b. $P_4 + 6 N_2O \rightarrow P_4O_6 + 6 N_2$

c. $As_2S_3 + 9/2 O_2 \rightarrow As_2O_3 + 3 SO_2$

d. $Ca_3(PO_4)_2 + 4 H_3PO_4 \rightarrow 3 Ca(H_2PO_4)_2$

EXERCISE!

Which of the following correctly balances the chemical equation given below? There may be more than one correct balanced equation. If a balanced equation is incorrect, explain what is incorrect about it.

$$CaO + C \longrightarrow CaC_2 + CO_2$$

I. $CaO_2 + 3C \longrightarrow CaC_2 + CO_2$

II. $\boxed{2CaO + 5C \longrightarrow 2CaC_2 + CO_2}$

III. $\boxed{CaO + (2.5)C \longrightarrow CaC_2 + (0.5)CO_2}$

IV. $\boxed{4CaO + 10C \longrightarrow 4CaC_2 + 2CO_2}$

CONCEPT CHECK!

Which of the following are true concerning balanced chemical equations? There may be more than one true statement.

I. The number of molecules is conserved.
II. The coefficients tell you how much of each substance you have.
III. Atoms are neither created nor destroyed.
IV. The coefficients indicate the mass ratios of the substances used.
V. The sum of the coefficients on the reactant side equals the sum of the coefficients on the product side.

Notice

- The number of atoms of each type of element must be the same on both sides of a balanced equation.
- Subscripts must not be changed to balance an equation.
- A balanced equation tells us the ratio of the number of molecules which react and are produced in a chemical reaction.
- Coefficients can be fractions, although they are usually given as lowest integer multiples.

Stoichiometric Calculations

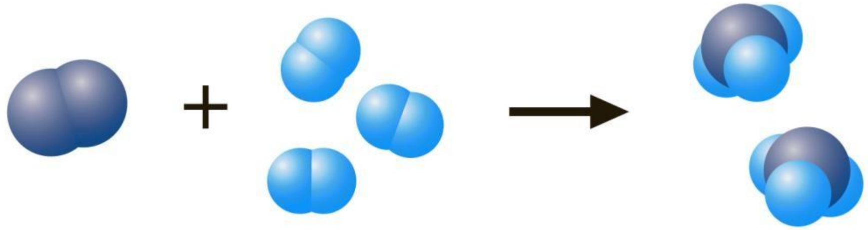
- Chemical equations can be used to relate the masses of reacting chemicals.

Calculating Masses of Reactants and Products in Reactions

1. Balance the equation for the reaction.
2. Convert the known mass of the reactant or product to moles of that substance.
3. Use the balanced equation to set up the appropriate mole ratios.
4. Use the appropriate mole ratios to calculate the number of moles of the desired reactant or product.
5. Convert from moles back to grams if required by the problem.

➢ Stoichiometry:Quantitative Relations in Chemical Reactions

Molar Interpretation of a Chemical Equation



$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

N_2	$+$	$3H_2$	$\longrightarrow$	$2NH_3$	
1 molecule N_2	$+$	3 molecules H_2	$\longrightarrow$	2 molecules NH_3	(molecular interpretation)
1 mol N_2	$+$	3 mol H_2	$\longrightarrow$	2 mol NH_3	(molar interpretation)
28.0 g N_2	$+$	3×2.02 g H_2	$\longrightarrow$	2×17.0 g NH_3	(mass interpretation)

Amounts of Substances in a Chemical Reaction

Relating the Quantity of Reactant to Quantity of Product In the following reaction:

$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

How many grams of Fe(s) can be produced from 1.00 kg Fe_2O_3? Molar masses are: Fe = 55.8 g/mol and Fe_2O_3 = 160 g/mol

Solution The calculation is as follows:

$$1.00 \times 10^3\ \cancel{\text{g Fe}_2\text{O}_3} \times \frac{1\ \cancel{\text{mol Fe}_2\text{O}_3}}{160\ \cancel{\text{g Fe}_2\text{O}_3}} \times \frac{2\ \cancel{\text{mol Fe}}}{1\ \cancel{\text{mol Fe}_2\text{O}_3}} \times \frac{55.8\ \text{g Fe}}{1\ \cancel{\text{mol Fe}}} = \mathbf{698\ g\ Fe}$$

Example

Relating the Quantities of Two Reactants (or Two Products)

In the following reaction:

$4HCl(\mathit{aq}) + MnO_2(\mathit{s}) \rightarrow 2H_2O(\mathit{l}) + MnCl_2(\mathit{aq}) + Cl_2(\mathit{g})$

How many grams of HCl react with 5.00 g of manganese dioxide, according to this equation?

$$5.00\ \cancel{\text{g MnO}_2} \times \frac{1\ \cancel{\text{mol MnO}_2}}{86.9\ \cancel{\text{g MnO}_2}} \times \frac{4\ \cancel{\text{mol HCl}}}{1\ \cancel{\text{mol MnO}_2}} \times \frac{36.5\ \text{g HCl}}{1\ \cancel{\text{mol HCl}}} = \mathbf{8.40\ g\ HCl}$$

Exercise

oxygen can be prepared by heating mercury(II) oxide, HgO. Mercury metal is the other product. If 6.47 g of oxygen is collected, how many grams of mercury metal are also produced? $2HgO \rightarrow 2Hg + O_2$

$$6.47\ \text{g}\ O_2 \times \frac{1\ \text{mol}\ O_2}{32.00\ \text{g}\ O_2} \times \frac{2\ \text{mol Hg}}{1\ \text{mol}\ O_2} \times \frac{200.59\ \text{g Hg}}{1\ \text{mol Hg}} = 81.\underline{1}1 = 81.1\ \text{g Hg}$$

How many grams of Al_2O_3 are produced when 41.5 g Al react?

$$2Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2Fe(l)$$

A. 78.4 g

B. 157 g

C. 314 g

D. 22.0 g

E. 11.0 g

n g $Al_2O_3 \rightarrow Al$

$\frac{41.5}{27} = 1.5$ mol Al

$$41.5\ \cancel{g\ Al}\left(\frac{1\ \cancel{mol\ Al}}{26.98\ \cancel{g\ Al}}\right)\left(\frac{1\ \cancel{mol\ Al_2O_3}}{2\ \cancel{mol\ Al}}\right)\left(\frac{101.96\ g\ Al_2O_3}{1\ \cancel{mol\ Al_2O_3}}\right)$$

$\frac{1.5}{2} = .75$

$= 78.4$ g Al_2O_3

$2Al \rightarrow 1 Al_2O_3$

$1.5 \rightarrow n$

$.75 \times 102 = 77$

How many grams of sodium dichromate are required to produce 24.7 g iron(III) chloride from the following reaction?

$$14HCl + Na_2Cr_2O_7 + 6FeCl_2 \rightarrow$$

$$2CrCl_3 + 7H_2O + 6FeCl_3 + 2NaCl$$

A. 6.64 g $Na_2Cr_2O_7$

B. 0.152 g $Na_2Cr_2O_7$

C. 8.51 g $Na_2Cr_2O_7$

D. 39.9 g $Na_2Cr_2O_7$

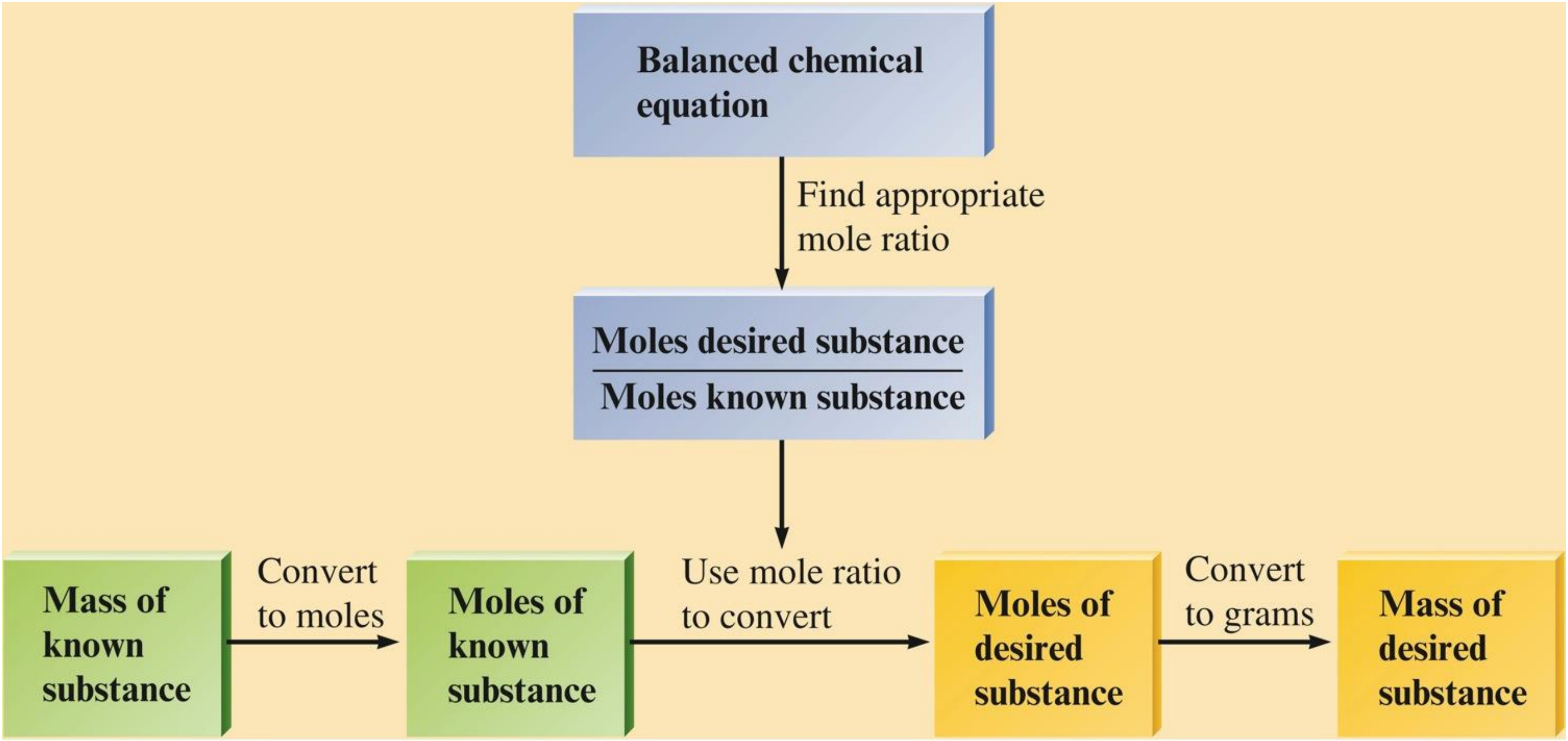
E. 8.04 g $Na_2Cr_2O_7$

$$24.7\ \cancel{\text{g FeCl}_3} \times \left(\frac{1\ \cancel{\text{mol FeCl}_3}}{162.2\ \cancel{\text{g FeCl}_3}}\right) \times$$

$$\left(\frac{1\ \cancel{\text{mol Na}_2\text{Cr}_2\text{O}_7}}{6\ \cancel{\text{mol FeCl}_3}}\right) \times \left(\frac{262.0\ \text{g Na}_2\text{Cr}_2\text{O}_7}{1\ \cancel{\text{mol Na}_2\text{Cr}_2\text{O}_7}}\right)$$

$$= 6.64\ \text{g Na}_2\text{Cr}_2\text{O}_7$$

Calculating Masses of Reactants and Products in Reactions



EXERCISE!

Consider the following reaction:

$$P_4(s) + 5\ O_2(g) \rightarrow 2\ P_2O_5(s)$$

If 6.25 g of phosphorus is burned, what mass of oxygen does it combine with?

8.07 g O_2

EXERCISE! (Part I)

Methane (CH_4) reacts with the oxygen in the air to produce carbon dioxide and water.

Ammonia (NH_3) reacts with the oxygen in the air to produce nitrogen monoxide and water.

- Write balanced equations for each of these reactions.

EXERCISE! (Part II)

Methane (CH_4) reacts with the oxygen in the air to produce carbon dioxide and water.

Ammonia (NH_3) reacts with the oxygen in the air to produce nitrogen monoxide and water.

- What mass of ammonia would produce the same amount of water as 1.00 g of methane reacting with excess oxygen?

Let's Think About It

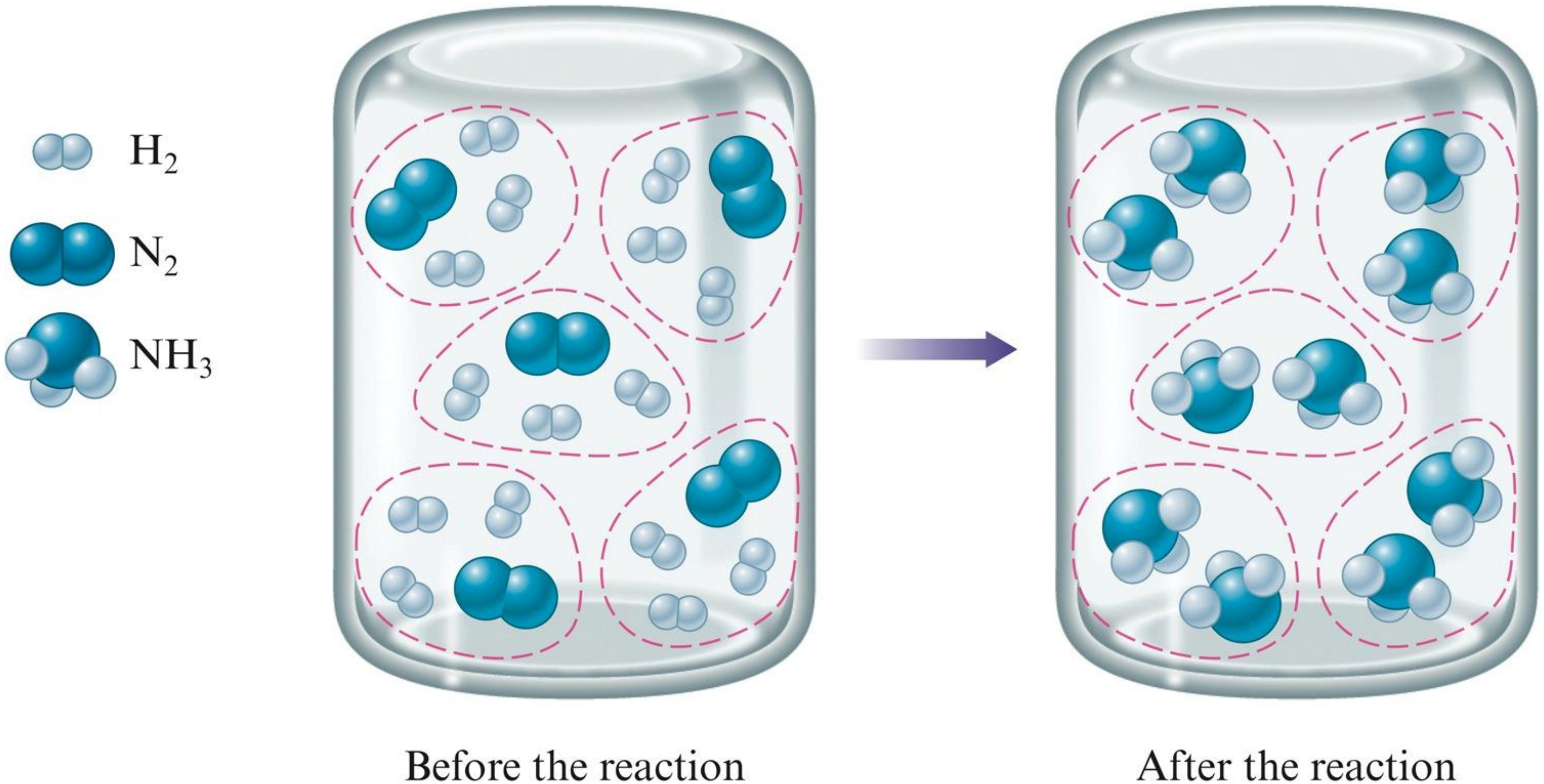
- Where are we going?
  - To find the mass of ammonia that would produce the same amount of water as 1.00 g of methane reacting with excess oxygen.
- How do we get there?
  - We need to know:
    - How much water is produced from 1.00 g of methane and excess oxygen.
    - How much ammonia is needed to produce the amount of water calculated above.

Limiting Reactants

- Limiting reactant – the reactant that runs out first and thus limits the amounts of products that can be formed.
- Determine which reactant is limiting to calculate correctly the amounts of products that will be formed.

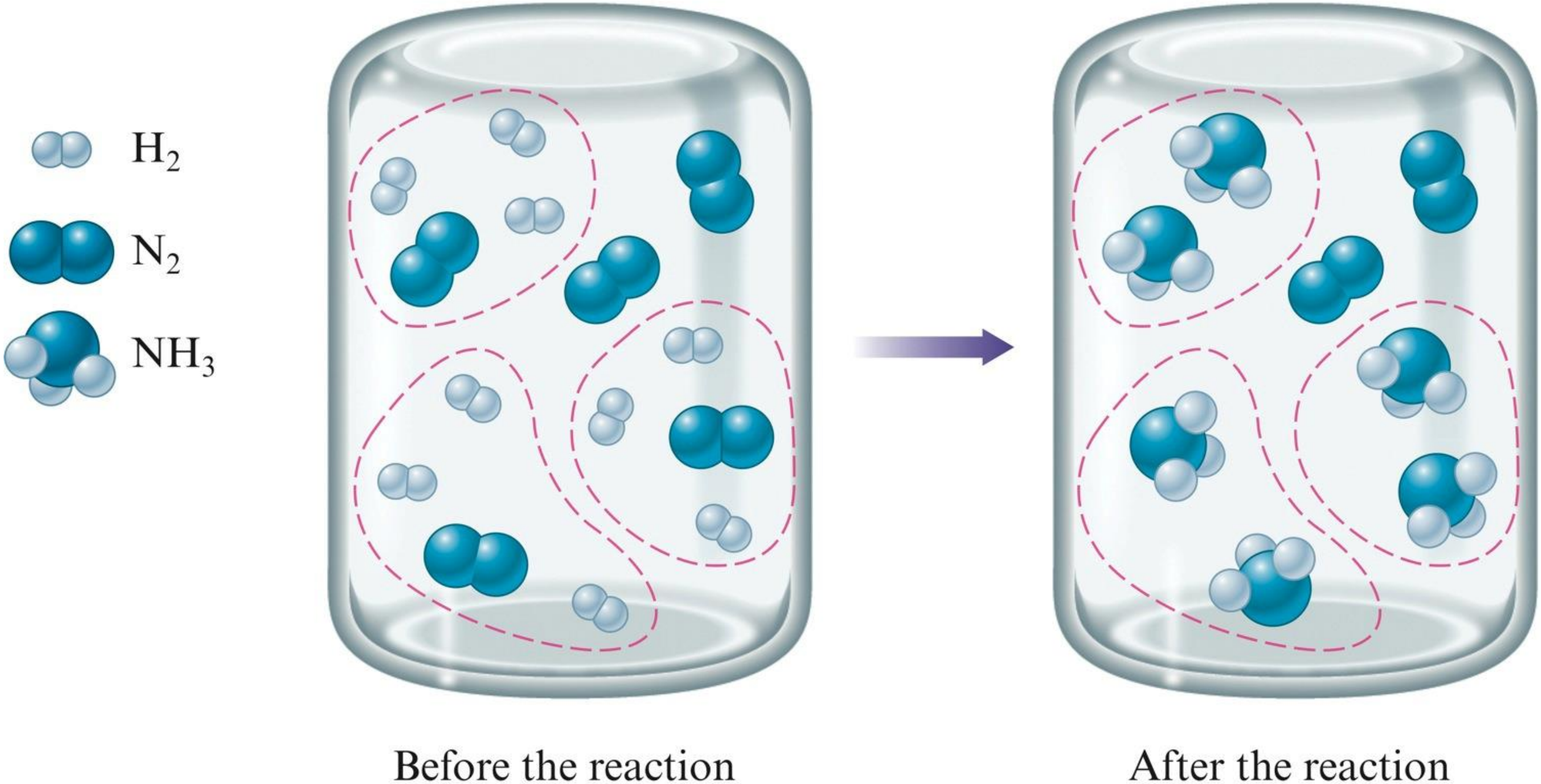
A. The Concept of Limiting Reactants

- Stoichiometric mixture
  - $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$



A. The Concept of Limiting Reactants

- Limiting reactant mixture



A. The Concept of Limiting Reactants

- Limiting reactant mixture
  - $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$
  - Limiting reactant is the reactant that runs out first.
    - H_2

Limiting Reactants

- The amount of products that can form is limited by the methane.
- Methane is the limiting reactant.
- Water is in excess.

CONCEPT CHECK!

Which of the following reaction mixtures could produce the greatest amount of product? Each involves the reaction symbolized by the equation:

$$2H_2 + O_2 \rightarrow 2H_2O$$

a) 2 moles of H_2 and 2 moles of O_2
b) 2 moles of H_2 and 3 moles of O_2
c) 2 moles of H_2 and 1 mole of O_2
d) 3 moles of H_2 and 1 mole of O_2
e) Each produce the same amount of product.

Notice

- We cannot simply add the total moles of all the reactants to decide which reactant mixture makes the most product. We must always think about how much product can be formed by using what we are given, and the ratio in the balanced equation.

CONCEPT CHECK!

- You know that chemical A reacts with chemical B. You react 10.0 g of A with 10.0 g of B.
  - What information do you need to know in order to determine the mass of product that will be produced?

Let's Think About It

- Where are we going?
  - To determine the mass of product that will be produced when you react 10.0 g of A with 10.0 g of B.
- How do we get there?
  - We need to know:
    - The mole ratio between A, B, and the product they form. In other words, we need to know the balanced reaction equation.
    - The molar masses of A, B, and the product they form.

EXERCISE!

You react 10.0 g of A with 10.0 g of B. What mass of product will be produced given that the molar mass of A is 10.0 g/mol, B is 20.0 g/mol, and C is 25.0 g/mol? They react according to the equation:

$$A + 3B \rightarrow 2C$$

Zinc metal reacts with hydrochloric acid by the following reaction:

$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

If 0.30 mol Zn is added to a solution containing 0.52 mol HCl, how many moles of H_2 are produced?

$$0.30 \cancel{\text{mol Zn}} \times \frac{1 \text{ mol } H_2}{1 \cancel{\text{mol Zn}}} = 0.30 \text{ mol } H_2$$

$$0.52 \cancel{\text{mol HCl}} \times \frac{1 \text{ mol } H_2}{2 \cancel{\text{mol HCl}}} = 0.26 \text{ mol } H_2$$

You see that hydrochloric acid must be the limiting reactant and that some zinc must be left unconsumed. (Zinc is the excess reactant.)

Step 2: Since HCl is the limiting reactant, the amount of H_2 produced must be **0.26 mol.**

Potassium superoxide, KO_2, is used in rebreathing gas masks to generate oxygen.

$4KO_2(s) + 2H_2O(l) \rightarrow 4KOH(s) + 3O_2(g)$

If a reaction vessel contains 0.25 mol KO_2 and 0.15 mol H_2O, what is the limiting reactant? How many moles of oxygen can

$$0.15 \text{ mol } H_2O \times \frac{3 \text{ mol } O_2}{2 \text{ mol } H_2O} = 0.2\underline{2}5 \text{ mol } O_2$$

$$0.25 \text{ mol } KO_2 \times \frac{3 \text{ mol } O_2}{4 \text{ mol } KO_2} = 0.1\underline{8}7 \text{ mol } O_2 \text{ (}KO_2\text{ is the limiting reactant.)}$$

Moles of O_2 produced $= 0.19$ mol

Hydrogen cyanide, HCN, is prepared from ammonia, air, and natural gas (CH_4) by the following process: $2NH_3(g) + 3O_2(g) + 2CH_4(g) \rightarrow 2HCN(g) + 6H_2O(g)$

If a reaction vessel contains 11.5 g NH_3, 12.0 g O_2, and 10.5 g CH_4, what is the maximum mass in grams of hydrogen cyanide that could be made, assuming the reaction goes to completion as written?

$$2NH_3 + 3O_2 + 2CH_4 \rightarrow 2HCN + 6H_2O$$

$$11.5 \text{ g } NH_3 \times \frac{1 \text{ mol } NH_3}{17.03 \text{ g } NH_3} \times \frac{2 \text{ mol HCN}}{2 \text{ mol } NH_3} = 0.67\underline{5} \text{ mol HCN}$$

$$10.5 \text{ g } CH_4 \times \frac{1 \text{ mol } CH_4}{16.04 \text{ g } CH_4} \times \frac{2 \text{ mol HCN}}{2 \text{ mol } CH_4} = 0.65\underline{4} \text{ mol HCN}$$

$$12.0 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} \times \frac{2 \text{ mol HCN}}{3 \text{ mol } O_2} = 0.25\underline{0}0 \text{ mol HCN}$$

O_2 is the limiting reactant.

$$\text{Mass HCN formed} = 0.2500 \text{ mol HCN} \times \frac{27.03 \text{ g HCN}}{1 \text{ mol HCN}} = 6.7\underline{5}7 = 6.76 \text{ g HCN}$$

Percent Yield

- An important indicator of the efficiency of a particular laboratory or industrial reaction.

$$\frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\% = \text{percent yield}$$

(Q) When 6.40 g of CH_3OH was mixed with 10.2 g of O_2 and ignited, 6.12 g of CO_2 was obtained. What was the percentage yield of CO_2?

$$2CH_3OH + 3O_2 \longrightarrow 2CO_2 + 4H_2O$$

MM(g/mol) (32.04) (32.00) (44.01) (18.02)

A. 6.12%

B. 8.79%

C. 100%

D. 142%

E. 69.6%

$$6.40\ \text{g}\ CH_3OH \times \frac{1\ \text{mol}\ CH_3OH}{32.04\ \text{g}\ CH_3OH} \times \frac{3\ \text{mol}\ CO_2}{2\ \text{mol}\ CH_3OH} \times \frac{32.00\ \text{g}\ O_2}{1\ \text{mol}\ O_2}$$

= 9.59 g O_2 needed; CH_3OH limiting

$$6.40\ \text{g}\ CH_3OH \times \frac{1\ \text{mol}\ CH_3OH}{32.04\ \text{g}\ CH_3OH} \times \frac{2\ \text{mol}\ CO_2}{2\ \text{mol}\ CH_3OH} \times \frac{44.01\ \text{g}\ CO_2}{1\ \text{mol}\ CO_2}$$

= 8.79 g CO_2 in theory

$$\frac{6.12\ \text{g}\ CO_2\ \text{actual}}{8.79\ \text{g}\ CO_2\ \text{theory}} \times 100\ \% = 69.6\%$$

Aspirin (acetylsalicylic acid) is prepared by heating salicylic acid, $C_7H_6O_3$, with acetic anhydride, $C_4H_6O_3$. The other product is acetic acid, $C_2H_4O_2$.

$$C_7H_6O_3 + C_4H_6O_3 \longrightarrow C_9H_8O_4 + C_2H_4O_2$$

What is the theoretical yield (in grams) of aspirin, $C_9H_8O_4$, when 2.00 g of salicylic acid is heated with 4.00 g of acetic anhydride? If the actual yield of aspirin is 1.86 g, what is the percentage yield?

$$C_7H_6O_3 + C_4H_6O_3 \rightarrow C_9H_8O_4 + C_2H_4O_2$$

$$4.00\text{ g } C_4H_6O_3 \times \frac{1\text{ mol } C_4H_6O_3}{102.09\text{ g } C_4H_6O_3} \times \frac{1\text{ mol } C_9H_8O_4}{1\text{ mol } C_4H_6O_3} = 0.039\underline{1}8\text{ mol } C_9H_8O_4$$

$$2.00\text{ g } C_7H_6O_3 \times \frac{1\text{ mol } C_7H_6O_3}{138.12\text{ g } C_7H_6O_3} \times \frac{1\text{ mol } C_9H_8O_4}{1\text{ mol } C_7H_6O_3} = 0.014\underline{4}8\text{ mol } C_9H_8O_4$$

Thus, $C_7H_6O_3$ is the limiting reactant. The theoretical yield of $C_9H_8O_4$ is

$$0.014\underline{4}8\text{ mol } C_9H_8O_4 \times \frac{180.15\text{ g } C_9H_8O_4}{1\text{ mol } C_9H_8O_4} = 2.6\underline{0}9\text{ g } C_9H_8O_4$$

The percentage yield is

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{1.86\text{ g}}{2.609\text{ g}} \times 100\% = 71.\underline{2}9 = 71.3\%$$

EXERCISE!

Consider the following reaction:

$$P_4(s) + 6F_2(g) \rightarrow 4PF_3(g)$$

- What mass of P_4 is needed to produce 85.0 g of PF_3 if the reaction has a 64.9% yield?

46.1 g P_4

Methyl salicylate (oil of wintergreen) is prepared by heating salicylic acid, $C_7H_6O_3$, with methanol, CH_3OH.

$$C_7H_6O_3 + CH_3OH \longrightarrow C_8H_8O_3 + H_2O$$

In an experiment, 2.50 g of salicylic acid is reacted with 10.31 g of methanol. The yield of methyl salicylate, C8H8O3, is 1.27 g. What is the percentage yield?

$$C_7H_6O_3 + CH_3OH \rightarrow C_8H_8O_3 + H_2O$$

$$11.20 \text{ g } CH_3OH \text{ x } \frac{1 \text{ mol } CH_3OH}{32.04 \text{ g } CH_3OH} \text{ x } \frac{1 \text{ mol } C_8H_8O_3}{1 \text{ mol } CH_3OH} = 0.34\underline{9}6 \text{ mol } C_8H_8O_3$$

$$1.50 \text{ g } C_7H_6O_3 \text{ x } \frac{1 \text{ mol } C_7H_6O_3}{138.12 \text{ g } C_7H_6O_3} \text{ x } \frac{1 \text{ mol } C_8H_8O_3}{1 \text{ mol } C_7H_6O_3} = 0.010\underline{8}6 \text{ mol } C_8H_8O_3$$

Thus, $C_7H_6O_3$ is the limiting reactant. The theoretical yield of $C_8H_8O_3$ is

$$0.010\underline{8}6 \text{ mol } C_8H_8O_3 \text{ x } \frac{152.14 \text{ g } C_8H_8O_3}{1 \text{ mol } C_8H_8O_3} = 1.6\underline{5}2 \text{ g } C_8H_8O_3$$

The percentage yield is

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \text{ x } 100\% = \frac{1.27 \text{ g}}{1.652 \text{ g}} \text{ x } 100\% = 76.\underline{8}7 = 76.9\%$$