



Chapter 2: Mass Relations in Formulas, Chemical Reactions, and Stoichiometry

Section 2.1: The Atomic Mass

The atomic mass is the mass of 1 atom. Atoms are very small and their mass is a very small number. A more practical unit is used to describe the mass of an atom. This unit is called the **atomic mass unit** (expressed as **a.m.u.** or just **u**).

The value of 1 atomic mass unit is chosen as 1/12 of the mass of one carbon-12 isotope. Remember that for carbon-12 the mass number A is equal to 12 (that is carbon-12 has 12 nucleons). The mass of a carbon-12 atom is equal to 1.9926×10^{-23} g. Hence, we conclude $1 \text{ u} = 1.6605 \times 10^{-24}$ g. The mass of carbon-12 atom is measured with an instrument called the mass spectrometer.

The atomic masses of elements are generally given in the Periodic Table and are located below the element symbol. For the element carbon, we note that the atomic mass is not 12 but 12.011 u. This is because the element carbon has several isotopes and the number 12.011 u is the average atomic mass of all the isotopes of the element carbon present in a typical sample on earth.

Note: Atomic masses are also called atomic weights.

Section 2.2: The Atomic Mass in the Periodic Table

Check the Periodic Table and look up the atomic mass of different elements by rolling your mouse over the element's symbol.

Section 2.3: Avogadro's Number and the Mole

Typical samples of matter contain huge numbers of atoms, often numbers as large as 10^{24} or more. The **mole** was established as a unit that is very useful when counting the numbers of atoms, ions or molecules. One mole is equal to the number of carbon atoms in 12 g of carbon-12.

1 atom of carbon-12 has a mass of 1.9927×10^{-23} g. Hence, in 12 g of carbon-12, there are:

$$\frac{12 \text{ g}}{1.9927 \times 10^{-23} \text{ g / C atom}} = 6.022 \times 10^{23} \text{ C atoms}$$

The number of 6.022×10^{23} is called **Avogadro's number**. Avogadro's number is expressed by the symbol **N_A**. Hence, one mole of atoms of carbon-12 (i.e. 12 g of carbon-12) contains Avogadro's number or 6.022×10^{23} atoms of carbon-12.

Note: The term "mole" is analogous to the term "dozen". While a dozen eggs refer to twelve eggs, a mole of particles (atoms, ions or molecules) refers to 6.022×10^{23} particles. It follows that while 2 dozen eggs consists of 24 eggs, 2 moles of particles consists of $2 \times (6.022 \times 10^{23})$ particles (i.e. 1.2044×10^{24} particles).

Sections 2.4 - 2.5: Concept of Formula Mass or Molar Mass

The formula mass, or molar mass, is the sum of atomic masses in a chemical formula.

Examples:

Chemical Formula	Formula Mass (a.m.u.)	Molar Mass (g/mol)
H	1.0	1.0
H ₂	2.0	2.0
O	16.0	16.0
O ₂	32.0	32.0

Many texts require you to use atomic masses to the nearest hundredth. However, in this DVD you need to always round off to the nearest tenth, unless otherwise specified.

Formula mass is the sum of atomic masses of all atoms in a formula of any molecular or ionic compound. The formula mass is expressed in a.m.u.

Molar mass is the sum of atomic masses of all atoms in a mole of pure substance. The molar mass is expressed in g/mol.

It is important that you know how to write chemical formulas from chemical names and vice versa.

Example: What is the formula mass of tungsten? Click on the Periodic Table. Roll the mouse and find tungsten, W. You will see the mass in red under W. Hence, the formula mass of tungsten, W, is 183.9 a.m.u.

Example: Calculate the molar mass of vitamin A, C₂₀H₃₀O.

The vitamin A molecule consists of:

Element	Number of Atoms		Atomic Mass
C	20	x	12.0 = 240.0
H	30	x	1.0 = 30.0
O	1	x	16.0 = 16.0
			286.0

Hence, the molar mass of vitamin A is 286.0 g/mol.

Example: Calculate the molar mass of sodium chloride. Note: Here the chemical name is given but not the chemical formula. Hence, it is important to know the names and formulas of chemicals.

Sodium chloride has the chemical formula NaCl.

Element	Number of Atoms		Atomic Mass
Na	1	x	23.0 = 23.0
Cl	1	x	35.5 = 35.5
			58.5 g/mol

Hence, the molar mass of sodium chloride is 58.5 g/mol.

In Section 2.5, practice Interactive Problems to master these concepts.

Sections 2.6 - 2.8: Conversion between Number of Moles and Mass

When discussing the amount of a substance, it is common practice to use the word “moles” instead of the more rigorous wording “number of moles”. The symbol “n” is often used to describe the number of moles or “moles” of a substance.

How many moles of a substance are present in a given sample of that substance can be calculated from the mass and the molar mass of that substance, according to the equation.

$$\text{Number of moles of a substance is equal to : } n = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

$$\text{Hence, mass} = n \times \text{molar mass}$$

Thus, if the moles and the chemical formula are given, one can calculate the mass in grams of that chemical.

Example: Calculate the number of moles of NH₃ in 1.0 g of NH₃.

$$n = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

The mass of NH₃ is 1.0 g. However, we need to calculate the molar mass of NH₃.

Element	Number of Atoms		Atomic Mass
N	1	x	14.0 = 14.0
H	3	x	1.0 = 3.0
			17.0 g/mol

$$n = \frac{1.0 \text{ g}}{17.0 \text{ g/mol}} = 0.0588235 \text{ mol}$$

In the correct number of significant figures, the answer is 0.059 mol.

Let us continue further with this problem (not covered in the DVD).

Example: How many moles of H and N atoms are present in 1.0 g of NH₃?

From the previous problem, we know that 1.0 g of NH₃ contains 0.059 mol NH₃. From the chemical formula of ammonia, we know that for every molecule of NH₃, there is 1 atom of nitrogen and 3 atoms of hydrogen. So, for every mole of NH₃, there is 1 mole of nitrogen and 3 moles of hydrogen.

Hence, moles of N = moles NH₃ = 0.059 mol N and Moles of H = 3 x moles of NH₃ = 3 x 0.059 mol = 0.18 mol H

Example: Calculate the mass in grams of 12.00 moles of C₂H₃Cl.

$$n = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}} \quad \text{Hence, mass} = n \times \text{molar mass}$$

Calculate the molar mass of C₂H₃Cl.

Element	Number of Atoms		Atomic Mass
C	2	x	12.0 = 24.0
H	3	x	1.0 = 3.0
Cl	1	x	35.5 = 35.5
			62.5 g/mol

$$\begin{aligned} \text{mass} &= 12.00 \text{ mol} \times 62.5 \text{ g/mol} \\ \text{mass} &= 750. \text{g} \end{aligned}$$

Example: Calculate the number of moles in 150.0 g of iron(III) oxide.

The chemical formula of iron(III) oxide is Fe_2O_3 . In the problem, the mass Fe_2O_3 is given. Hence, the molar mass must be calculated.

Element	Number of Atoms		Atomic Mass
Fe	2	x	55.8 = 111.6
O	3	x	16.0 = 48.0
			159.6 g/mol

$$n = \frac{150.0 \text{ g}}{159.6 \text{ g/mol}} = 0.93984 \text{ mol}$$

In the correct number of significant figures, the answer is 0.9398 mol.

In Section 2.7, visualize one mole for a variety of elements and compounds.

In Section 2.8, practice Interactive Problems.

Sections 2.9 - 2.10: Problems on Avogadro's Number

In some instances, we want to know how many atoms, ions or molecules are involved in a chemical or physical process. We can calculate this number from the mass of the substance, the molar mass of the substance and Avogadro's number.

Avogadro's number is given by the symbol $N_A = 6.022 \times 10^{23}$ or 6.022E23

$$\text{Remember mole, } n = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}} \text{ and one mole} = 6.022 \times 10^{23}$$

Example 1: The molecular formula of ethylene glycol is $\text{C}_2\text{H}_6\text{O}_2$. In 13.68 g of ethylene glycol, (a) calculate the number of moles

Element	Number of Atoms		Atomic Mass
C	2	x	12.0 = 24.0
H	6	x	1.0 = 6.0
O	2	x	16.0 = 32.0
			62.0

$$n = \frac{13.68 \text{ g}}{62.0 \text{ g/mol}}$$

$$n = 0.221 \text{ mol}$$

Example 1: The molecular formula of ethylene glycol is $\text{C}_2\text{H}_6\text{O}_2$. In 13.68 g of ethylene glycol, **(b) calculate the number of molecules**

$$\frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} \times 0.221 \text{ mol} = 1.33 \times 10^{23} \text{ molecules or } 1.33\text{E}23 \text{ molecules}$$

Example 1: The molecular formula of ethylene glycol is $\text{C}_2\text{H}_6\text{O}_2$. In 13.68 g of ethylene glycol, **(c) Calculate the number of oxygen atoms**

In each molecule of $\text{C}_2\text{H}_6\text{O}_2$, there are 2 atoms of O. Therefore, in 1.33×10^{23} molecules of $\text{C}_2\text{H}_6\text{O}_2$ there are:

$$2 \times 1.33 \times 10^{23} = 2.66 \times 10^{23} \text{ or } 2.66\text{E}23 \text{ atoms of oxygen}$$

Example 2: A sample of the compound, $\text{C}_3\text{H}_6\text{O}$, contains 14.0×10^{14} carbon atoms. (a) Calculate the number of $\text{C}_3\text{H}_6\text{O}$ molecules

Each molecule of $\text{C}_3\text{H}_6\text{O}$ contains 3 carbon atoms. Therefore, the number of $\text{C}_3\text{H}_6\text{O}$ molecules containing 14.0×10^{14} C atoms is:

$$\frac{14.0 \times 10^{14}}{3} = 4.67 \times 10^{14} \text{ molecules of } \text{C}_3\text{H}_6\text{O}$$

Example 2: A sample of the compound, $\text{C}_3\text{H}_6\text{O}$, contains 14.0×10^{14} carbon atoms. **(b) Calculate the number of moles of $\text{C}_3\text{H}_6\text{O}$.**

$$\frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} \times 4.67 \times 10^{14} \text{ molecules} = 7.75 \times 10^{-10} \text{ or } 7.75\text{E}-10 \text{ mol}$$

Example2: A sample of the compound, $\text{C}_3\text{H}_6\text{O}$, contains 14.0×10^{14} carbon atoms. **(c) Calculate the number of grams of $\text{C}_3\text{H}_6\text{O}$.**

Element	Number of Atoms		Atomic Mass
C	3	x	12.0 = 36.0
H	6	x	1.0 = 6.0
O	1	x	16.0 = 16.0

		58.0 g/mol
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The molar mass of C₃H₆O is 58.0 g/mol

$$n = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

Therefore, mass (g) = n x molar mass (g/mol)

$$\begin{aligned} \text{Thus, mass} &= 7.75 \times 10^{-10} \text{ mol} \times 58.0 \text{ g/mol} \\ \text{mass} &= 4.50 \times 10^{-8} \text{ or } 4.50\text{E-}8 \text{ g} \end{aligned}$$

In Section 2.10, practice the Interactive Problems.

Sections 2.11 - 2.12: Percent Composition

The percent composition of a compound is the mass percent of the elements present.

$$\text{Mass percent of element} = \frac{\text{mass of element in sample}}{\text{total mass of sample}} \times 100$$

Knowing the chemical formula of a compound, the mass percent of its constituent elements can be calculated.

Note: the subscripts in a chemical formula allow us to define the atom ratio as well as the mole ratio in which the different elements are combined.

Example:

In H₂O: the atom ratio is 2 atoms H: 1 atom O
 the mole ratio is 2 moles H: 1 mole O

In Na₂SO₄: the atom ratio is 2 atoms Na: 1 atom S: 4 atoms O
 the mole ratio is 2 moles Na: 1 mole S: 4 moles O

Example 1: Calculate the mass percent of H and O in H₂O.

Element	Number of Moles	Molar Mass (g/mol)
H	2 x	1.0 = 2.0 g
O	1 x	16.0 = 16.0 g
		18.0 g

$$\begin{aligned}\text{Mass \% of H} &= \frac{2.0 \text{ g}}{18.0 \text{ g}} \times 100 \\ &= 11\%\end{aligned}$$

$$\begin{aligned}\text{Mass \% O} &= \frac{16.0 \text{ g}}{18.0 \text{ g}} \times 100 \\ &= 88.9\%\end{aligned}$$

Example 2: Calculate the mass percent of N in C₃H₃N

Element	Number of Moles	Molar Mass (g/mol)
C	3 x	12.0 = 36.0 g
H	3 x	1.0 = 3.0 g
N	1 x	14.0 = 14.0 g
		53.0 g

$$\begin{aligned}\text{Mass \% N} &= \frac{14.0 \text{ g}}{53.0 \text{ g}} \times 100 \\ &= 26.4\%\end{aligned}$$

In Section 2.12, practice the Interactive Problems.

Sections 2.13 - 2.14: Empirical Formula

When a new compound is formed or discovered, it is important to determine the chemical formula. Most often, this is done by taking a known amount of sample and decomposing, or breaking down this compound into its component elements.

Or

Taking a known amount of sample and reacting it with oxygen to produce CO₂ and H₂O. The component elements or CO₂ and H₂O, are then collected and weighed. The results of such analyses give the mass of each element in the compound.

This is used to determine the mass percent of each element in the compound. Knowing the mass percent of each element in the compound makes it possible to determine its chemical formula.

Empirical formula is the simplest chemical formula. The simplest formula gives only the ratios of atoms in a compound.

Example: A 50.00 g sample contains 13.28 g of potassium, 17.68 g of chromium, and 19.04 g of oxygen. Find the simplest formula.

Analyze the problem. The sample mass is given, the masses of elements are also given. Make sure the sum of masses of all elements in the sample is equal to the mass of the sample.

Step1: Calculate the number of moles of K, Cr and O in the given masses.

$$\text{Remember that moles} = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

$$13.28 \text{ g (K)} \Rightarrow \frac{13.28 \text{ g}}{39.1 \text{ g/mol}} = 0.340 \text{ mol (K)}$$

$$17.68 \text{ g (Cr)} \Rightarrow \frac{17.68 \text{ g}}{52.0 \text{ g/mol}} = 0.340 \text{ mol (Cr)}$$

$$19.04 \text{ g (O)} \Rightarrow \frac{19.04 \text{ g}}{16.0 \text{ g/mol}} = 1.19 \text{ mol (O)}$$

Step 2: Divide each of the numbers of moles by the smallest number of moles to obtain the relative amounts in moles of each element in the substance.

$$\text{K : } \frac{0.340 \text{ mol}}{0.340 \text{ mol}} = 1.00$$

$$\text{Cr : } \frac{0.340 \text{ mol}}{0.340 \text{ mol}} = 1.00$$

$$\text{O : } \frac{1.19 \text{ mol}}{0.340 \text{ mol}} = 3.50$$

Step 3: Write the formula using these relative numbers of moles of each element. Remember that the subscripts in a formula give the relative numbers of atoms or moles of atoms in that substance.

The results in Step 2 suggest that the simplest formula is: $\text{K}_1\text{Cr}_1\text{O}_{3.5}$

Step 4: Write the final formula, ensuring all subscripts are whole numbers.

We multiply each subscript by 2 to get the empirical formula: **$\text{K}_2\text{Cr}_2\text{O}_7$** . This is potassium dichromate. This formula makes sense because the dichromate ion has a -2 charge. The potassium ion has a +1 charge. Hence, this substance has a neutral formula, as it should.

In Section 2.14, practice the Interactive Problem.

Sections 2.15 - 2.16: More Problems on Empirical Formula

To find the composition of a substance, it is often useful to react that substance with oxygen gas. This is an example of **chemical analysis**. Combustion reactions are useful for the chemical analysis of substances containing carbon and hydrogen as they produce CO_2 and H_2O . Measuring the amount of CO_2 and H_2O produced by a given amount of substance allows the determination of how much carbon and hydrogen are present in that amount of substance.

Example: When 5.000 g of ibuprofen is burnt with oxygen gas ($\text{O}_{2(g)}$), 13.86 g of $\text{CO}_{2(g)}$ and 3.926 g of $\text{H}_2\text{O}_{(l)}$ are formed. Use the following information to determine the empirical formula of ibuprofen. Ibuprofen is known to contain only carbon, oxygen and hydrogen elements.

Step 1: How much carbon is there in 5.000 g of ibuprofen?

All the carbon in ibuprofen ends up in the 13.86 g of $\text{CO}_{2(g)}$. So, the question is how many moles of carbon are present in 13.86 g of $\text{CO}_{2(g)}$.

$$13.86 \text{ g } (\text{CO}_2) \Rightarrow \frac{13.86 \text{ g } (\text{CO}_2)}{44.01 \text{ g/mol } (\text{CO}_2)} = 0.3149 \text{ mol } (\text{CO}_2) \Rightarrow 0.3149 \text{ mol } (\text{C})$$

$$\Rightarrow \text{mass (C) in 5 g ibuprofen} = 0.3149 \text{ mol } (\text{C}) \times 12.01 \text{ g/mol } (\text{C}) = 3.782 \text{ g } (\text{C})$$

Step 2: How much hydrogen is there in 5.000 g of ibuprofen?

All the hydrogen in ibuprofen ends up in the 3.926 g of $\text{H}_2\text{O}_{(l)}$. So, the question is how many moles of hydrogen are present in 3.926 g of $\text{H}_2\text{O}_{(l)}$.

$$3.926 \text{ g } (\text{H}_2\text{O}) \Rightarrow \frac{3.926 \text{ g } (\text{H}_2\text{O})}{18.00 \text{ g/mol } (\text{H}_2\text{O})} = 0.2181 \text{ mol } (\text{H}_2\text{O}) \Rightarrow 0.4362 \text{ mol } (\text{H})$$

$$\Rightarrow \text{mass (H) in 5 g ibuprofen} = 0.4362 \text{ mol } (\text{H}) \times 1.00 \text{ g/mol } (\text{H}) = 0.4362 \text{ g } (\text{H})$$

Step 3: How much oxygen is there in 5.000 g of ibuprofen?

Since ibuprofen only contains oxygen, carbon and hydrogen and 5.000 g of ibuprofen contain 3.782 g (C) and 0.4362 g (H), then, the mass of oxygen is:

$$\text{Mass (O)} = 5.000 \text{ g} - 3.782 \text{ g} - 0.4362 \text{ g} = 0.7818 \text{ g } (\text{O}).$$

Step 4: Now, we can use the strategy shown in Section 2.13 to determine the formula of ibuprofen.

$$\text{C} : 3.782 \text{ g} \Rightarrow \frac{3.782 \text{ g (C)}}{12.0 \text{ g/mol (C)}} = 0.3149 \text{ mol (C)}$$

$$\text{H} : 0.4362 \text{ g} \Rightarrow \frac{0.4362 \text{ g (H)}}{1.0 \text{ g/mol (H)}} = 0.4362 \text{ mol (H)}$$

$$\text{O} : 0.7818 \text{ g} \Rightarrow \frac{0.7818 \text{ g (O)}}{16.0 \text{ g/mol (O)}} = 0.04886 \text{ mol (O)}$$

$$\text{For Carbon} \Rightarrow \frac{0.3149 \text{ mol (C)}}{0.04886 \text{ mol (O)}} = 6.4 \text{ (C) per (O) approximately 6.5}$$

$$\text{For Hydrogen} \Rightarrow \frac{0.4362 \text{ mol (H)}}{0.04886 \text{ mol (O)}} = 8.9 \text{ (H) per (O) approximately 9}$$

Hence, the formula for ibuprofen is: $\text{C}_{6.5}\text{H}_9\text{O}_1$ or more appropriately, $\text{C}_{13}\text{H}_{18}\text{O}_2$.

In Section 2.16, practice the Interactive Problem.

Sections 2.17 - 2.18: Molecular Formula

The empirical formula of a substance is always written using the smallest possible whole number subscripts to give the relative number of atoms of each element in the substance.

Hence, the empirical formula for sodium chloride is written as NaCl and not Na_2Cl_2 . NaCl is an ionic compound, not a molecule. Hence, the entity NaCl is called a **formula unit**.

Remember that for ionic compounds, the chemical formula and the empirical formula are always one and the same formula. For molecular (covalent) compounds, however, molecular and empirical formula may be different.

A molecular formula is a whole number multiple of the simplest chemical formula.

Or

A molecular formula is a whole multiple of the empirical formula.

To find the multiple, the molar mass is needed. The empirical formula mass can be calculated from the empirical formula.

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

Example: The mass composition of lindane is 24.78% C, 2.08% H and 73.14% Cl. The molar mass of lindane is 290.85 g/mol. Determine the molecular formula.

Element	Mass (g)	Molar Mass (g/mol)	Moles	Mole Ratio
C	24.78	12.0	2.06	1
H	2.08	1.0	2.08	1
Cl	73.14	35.5	2.06	1

Hence, the simplest formula or the empirical formula is CHCl.

The empirical formula mass is $12 + 1 + 35.5 = 48.5 \text{ g/mol}$

$$\text{Multiple} = \frac{290.85 \text{ g/mol}}{48.5 \text{ g/mol}} = 6$$

Hence, the molecular formula is: $\text{C}_6\text{H}_6\text{Cl}_6$

In Section 2.18, practice the Interactive Problem.

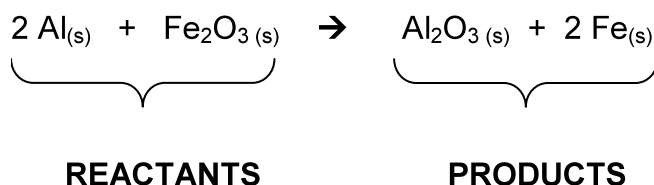
Section 2.19: Balancing Chemical Reactions

Sections 2.19.1 - 2.19.3: Introduction

A **chemical reaction** is an actual transformation of substances called reactants into substances called products.

To represent a chemical reaction we use a **chemical equation**, a sort of recipe which shows in a symbolic form 1) who the participating substances are (reactants and products), 2) the state or phase these substances are in (solid, liquid, gas, aqueous solution) and 3) the amount in which they must be present (number of atoms, molecules (for covalent compounds) or formula units (for ionic compounds)).

Example: Consider the reaction of aluminum metal ($\text{Al}_{(s)}$) with solid iron oxide ($\text{Fe}_2\text{O}_{3(s)}$) forming solid aluminum oxide ($\text{Al}_2\text{O}_{3(s)}$) and solid iron ($\text{Fe}_{(s)}$). This reaction is represented by the following chemical equation:



The arrow (\rightarrow) shows the direction in which the chemical transformation takes place. The **reactants** (shown on the left-hand side of the arrow) are the substances with which the reaction is started. The **products** (shown on the right-hand side of the arrow) are the substances resulting from the reaction. In the above reaction all substances are in the solid state, as indicated by the subscript “s” in parentheses.

The **state** or **phase** of a substance is always indicated by a subscript in parentheses after the chemical formula of that substance. The following notations are used for the various phases encountered in chemical reactions:

- (s) for solids
- (l) for liquids
- (g) for gases
- (aq) for aqueous solutions

Finally, one of the most important pieces of information conveyed by a chemical equation is the number of atoms, ions, formula units or molecules associated with each substance. The number in front of each substance is called the **stoichiometric coefficients** or more simply the coefficient. The bulk of this information is often referred to as the **stoichiometry** of the chemical reaction.

For the above reaction, the stoichiometric coefficients are 2, 1, 1 and 2, respectively. Note that when a stoichiometric coefficient is 1, it is not shown (as is the case for Al_2O_3 and Fe_2O_3).

Stoichiometric coefficients play a very important role in chemical equations. Their presence insures that the number of atoms of each type is the same on the reactants and products sides. For instance, in the above reaction there are two aluminum atoms, two iron atoms and three oxygen atoms on each side of the chemical equation. This observation reflects Dalton’s hypothesis that, in a chemical reaction, atoms are neither destroyed nor created.

To keep a chemical equation looking as simple as possible, we will generally ensure that stoichiometric coefficients are written using the smallest possible whole numbers (integers). **Balancing** a chemical equation consists in determining each of the stoichiometric coefficients.

Section 2.19.4: Balancing the Chemical Reaction: $\text{CO}_{2(g)} + \text{H}_2\text{O}_{(l)} \rightarrow \text{C}_6\text{H}_{12}\text{O}_{6(s)} + \text{O}_{2(g)}$

To balance a chemical reaction, we always start balancing the elements that are present in the least number of compounds. In the above equation, we can start with either carbon or hydrogen.

Balancing the element carbon.

There is one carbon atom on the left-hand side (in $\text{CO}_{2(g)}$) and six carbon atoms on the right-hand side (in $\text{C}_6\text{H}_{12}\text{O}_{6(s)}$). Hence, we place a coefficient of 6 in front of $\text{CO}_{2(g)}$ and 1 in front of $\text{C}_6\text{H}_{12}\text{O}_{6(s)}$. Doing so leads to the balancing of carbon.

When one assigns stoichiometric coefficients to some substances (here, $\text{CO}_{2(g)}$ and $\text{C}_6\text{H}_{12}\text{O}_{6(s)}$), one cannot change these coefficients when balancing another element.

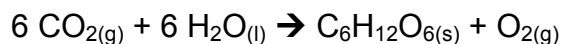
Balancing the element hydrogen.



There are two hydrogen atoms on the left-hand side (in $\text{H}_2\text{O}_{(l)}$) and 12 hydrogen atoms on the right-hand side (in $\text{C}_6\text{H}_{12}\text{O}_{6(s)}$). Hence, we place a coefficient of 6 in front of $\text{H}_2\text{O}_{(l)}$. Doing so leads to the balancing of hydrogen.

Assigning a stoichiometric coefficient to $\text{H}_2\text{O}_{(l)}$, implies that this coefficient can no longer be changed when balancing the last element, oxygen.

Balancing the element oxygen.



There are $12 + 6 = 18$ oxygen atoms on the left-hand side. On the right-hand side here are 6 oxygen atoms in $\text{C}_6\text{H}_{12}\text{O}_{6(s)}$ and 2 oxygen atoms in $\text{O}_{2(g)}$. At this stage, we can no longer change the stoichiometric coefficients of $\text{CO}_{2(g)}$, $\text{H}_2\text{O}_{(l)}$, or $\text{C}_6\text{H}_{12}\text{O}_{6(s)}$ since these coefficients have already been assigned when balancing C and H. Hence, to obtain 18 oxygen atoms on the right-hand side, we must assign a stoichiometric coefficient of 6 to $\text{O}_{2(g)}$.

The balanced chemical equation is:

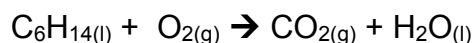


The sum of the stoichiometric coefficients for reactants and products is **19**.

Section 2.19.5: Balancing the Equation of a Combustion Reaction

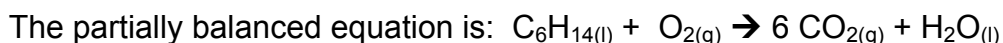
A combustion reaction is a reaction in which a substance (element or compound) is burnt with oxygen gas (O₂). The combustion reactions of organic molecules (molecules based on carbon, hydrogen, oxygen, etc...) lead to the formation of carbon dioxide (CO₂) gas and water (H₂O).

Consider the combustion reaction of hexane, C₆H_{14(l)}. For the combustion of hexane, oxygen gas must be a reactant and CO₂ and H₂O must be products.



First, we **balance carbon**:

There are 6 carbon atoms in C₆H₁₄ and 1 carbon atom in CO₂. Hence, we will use stoichiometric coefficients of 1 (not shown) for C₆H₁₄ and 6 for CO₂.



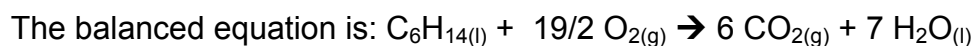
Next, we **balance hydrogen**:

There are 14 hydrogen atoms in one molecule of C₆H₁₄ and 2 hydrogen atoms in one molecule of H₂O. Hence, we will use a coefficient of 7 for H₂O.

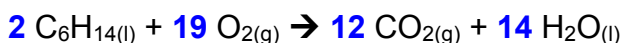


Finally, we **balance oxygen**:

On the left-hand side there are 2 oxygen atoms. On the right-hand side there are 6x2 + 7x1 = 19 oxygen atoms. However, oxygen is present as O₂ on the reactant side. Hence, we should use 19/2 O₂ molecules on the right-hand side.



If we decide to use only whole numbers as stoichiometric coefficients in a balanced equation, then, we must multiply all coefficients by 2. Hence, the final form for the balanced combustion reaction of hexane is:



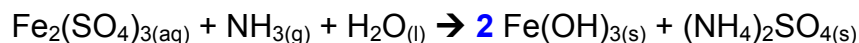
The sum of the stoichiometric coefficients for reactants and products is 47.

Section 2.19.6: Balancing the Chemical Reaction: $\text{Fe}_2(\text{SO}_4)_{3(\text{aq})} + \text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Fe}(\text{OH})_{3(\text{s})} + (\text{NH}_4)_2\text{SO}_{4(\text{s})}$

First, we **balance iron**:

There are 2 iron atoms in $\text{Fe}_2(\text{SO}_4)_3$ and 1 iron atom in $\text{Fe}(\text{OH})_3$. Hence, we will use stoichiometric coefficients of 1 (not shown) for $\text{Fe}_2(\text{SO}_4)_3$ and 2 for $\text{Fe}(\text{OH})_3$.

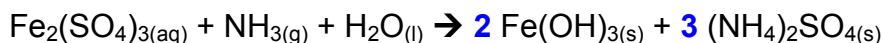
The partially balanced equation is:



Next, we **balance sulfur**:

There are 3 sulfur atoms in one formula unit of $\text{Fe}_2(\text{SO}_4)_3$ and 1 sulfur atom in one formula unit of $(\text{NH}_4)_2\text{SO}_4$. Hence, we will use a stoichiometric coefficient of 3 for $(\text{NH}_4)_2\text{SO}_4$.

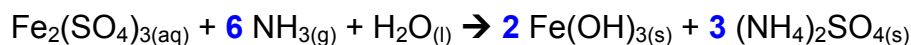
The partially balanced equation becomes:



Next, we **balance nitrogen**:

On the left-hand side there is 1 nitrogen atom in NH_3 . On the right-hand side there are $3 \times 2 = 6$ nitrogen atoms in 3 formula units of $(\text{NH}_4)_2\text{SO}_4$. Hence, we will use a stoichiometric coefficient of 6 for NH_3 .

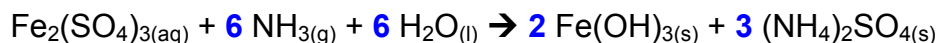
The partially balanced equation becomes:



Next, we **balance oxygen**:

On the left-hand side there are $3 \times 4 = 12$ oxygen atoms in one formula unit of $\text{Fe}_2(\text{SO}_4)_3$ and 1 oxygen atom in one molecule of $\text{H}_2\text{O}(\text{l})$. On the right-hand side there are 18 oxygen atoms [$2 \times 3 = 6$ oxygen atoms in 2 formula units of $\text{Fe}(\text{OH})_3$ and $3 \times 4 = 12$ oxygen atoms in 3 formula units of $(\text{NH}_4)_2\text{SO}_4$]. The only coefficient that can be modified is the coefficient for H_2O since all other coefficients have already been assigned. Hence, we will use a stoichiometric coefficient of 6 for H_2O .

The partially balanced equation becomes:

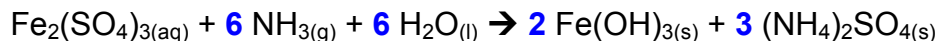


Finally, we **balance hydrogen**:

Note that we expect the chemical equation to be balanced for hydrogen as all stoichiometric coefficients have already been assigned.

On the left-hand side there are 30 hydrogen atoms [$6 \times 3 = 18$ hydrogen atoms in three formula units of NH_3 and $6 \times 2 = 12$ atoms in two molecules of H_2O]. On the right-hand side there are also 30 hydrogen atoms [$2 \times 3 = 6$ hydrogen atoms in 2 formula units of $\text{Fe}(\text{OH})_3$ and $3 \times 4 \times 2 = 24$ hydrogen atoms in 3 formula units of $(\text{NH}_4)_2\text{SO}_4$].

Hence, the fully balanced chemical equation is:



The sum of the stoichiometric coefficients for products and reactants is **18**.

Sections 2.19.7 - 2.19.8: Balancing Chemical Reactions

These are interactive sessions where you are guided to practice what you have learned in the first 6 sections of part 2.19.

Section 2.19.9: Practice Problems on Balancing Reactions

Practice, practice, practice.

If you have problems, go back and study Sections 2.19.1 through 2.19.8!

Sections 2.19.10 - 2.19.12: Balancing Chemical Equations in Word Problems

In these sections, you are taught to analyze statements in a word problem and come up with a chemical reaction. Demonstration of the reaction is shown in the video. Following this, a step by step explanation of balancing the chemical reaction is offered.

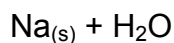
Chemical reactions are represented by chemical equations. Chemical equations consist of reactants and products.

Consider the reaction: Solid sodium metal reacts with water giving a solution of sodium hydroxide and releasing hydrogen gas.

Write a chemical reaction using complete formulas with phases.

Sodium is Na. It exists as a solid. Hence, put "(s)" next to $\text{Na}_{(\text{s})}$

The formula for water is H_2O . Add H_2O



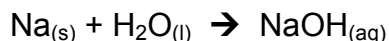
H₂O is a liquid. Hence, put "(l)" next to it. $\text{Na}_{(s)} + \text{H}_2\text{O}_{(l)}$

Giving sodium hydroxide means, produces sodium hydroxide.

Draw the line with arrow. $\text{Na}_{(s)} + \text{H}_2\text{O}_{(l)} \rightarrow$

Chemical formula for sodium hydroxide is NaOH.

Solution of NaOH, put "(aq)" next to it.



Remember hydrogen exists as H₂.



Since H₂ exists as gas, put "(g)" next to it.



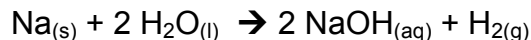
Reactants are on the left-hand side (LHS) of the equation. Products are on the right-hand side (RHS) of the equation.

The coefficients of a balanced equation are positive integers. Note: Integers will be used here only for the sake of convenience. In reality one can use rational numbers (fractions) to balance chemical reactions. To balance a chemical reaction only the coefficients can be changed. The balanced chemical reaction contains the smallest possible coefficients.



On the left-hand side of the equation there is one Na. On the right-hand side of the equation there is one Na. Hence, Na is balanced.

On the left-hand side of the equation there are two H. On the right-hand side of the equation there are three H. In this case there should be an even number of H on both sides. The smallest even number that could work would be four. Hence, a coefficient of 2 is placed for H₂O and a coefficient of 2 is placed for NaOH.



On the left-hand side of the equation there are two O. On the right-hand side of the equation there are two O. Hence, O is balanced.

Now, on the left-hand side of the equation there is one Na. On the right-hand side of the equation there are two Na. Hence, a coefficient of 2 should be placed for Na on the left-hand side.



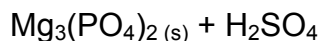
Note: When no coefficient appears, it is understood that the coefficient is 1. Hence, the sum of coefficients of reactants and products for the balanced chemical reaction is 7.

Solid magnesium phosphate reacts with an aqueous solution of sulfuric acid giving magnesium sulfate as a solid and phosphoric acid as a solution. Write a chemical reaction using complete formulas with phases.

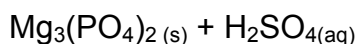
Magnesium phosphate has a chemical formula $\text{Mg}_3(\text{PO}_4)_2$. It exists as a solid. Hence, put "(s)" next to it.



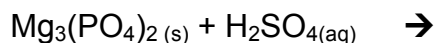
The formula for sulfuric acid is H_2SO_4



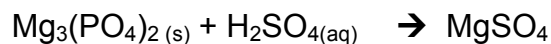
H_2SO_4 is an aqueous solution. Hence, put "(aq)" next to it.



Giving magnesium sulfate means produces magnesium sulfate. Draw the line with the arrow.



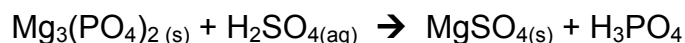
Chemical formula for magnesium sulfate is MgSO_4



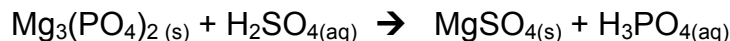
MgSO_4 is produced as a solid. Hence, put "(s)" next to it.



Chemical formula for phosphoric acid is H_3PO_4



H₃PO₄ as a solution, put "(aq)" next to it.



Reactants are on the left-hand side (LHS) of the equation. Products are on the right-hand side (RHS) of the equation.

On the left-hand side of the equation there are three Mg. On the right-hand side of the equation there is one Mg. Hence, a coefficient of 3 is placed for Mg on the RHS.

On the left-hand side of the equation there are two P. On the right-hand side of the equation there is one P. Hence, a coefficient of 2 is placed for P on the RHS.

On the left-hand side of the equation there are two H. On the right-hand side there are six H. Hence, a coefficient of 3 is placed for H on the LHS.

On the left-hand side of the equation there are three S. On the right-hand side of the equation there are three S. Hence, S is balanced.

On the left-hand side of the equation there are eight + twelve O = 20 O.
On the right-hand side of the equation there are twelve + eight O = 20 O.
Hence, O is balanced.

Note: When no coefficient appears, it is understood that the coefficient is 1. Hence, the sum of coefficients of reactants and products for the balanced chemical reaction is 9.

Sections 2.20 – 2.21: Stoichiometry

We are interested in **reaction stoichiometry** whenever we ask questions such as:

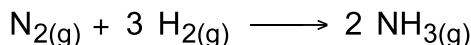
- 1) What is the amount of each reactant required to produce a known amount of product, or,
- 2) What is the amount of product formed from a known amount of reactants?

Practically speaking, "amounts" of reactants or products are the masses of these compounds measured in the laboratory in grams.

To relate masses of products to masses of reactants requires relating:

- 1) Masses to moles using the molar masses, and,
- 2) Moles or reactants to moles of products, using the stoichiometry of the balanced chemical reaction (that is, using the values of the stoichiometric coefficients).

Example: In the Haber process, nitrogen reacts with hydrogen to produce ammonia gas.



The coefficients (1, 3 and 2) represent the number of moles.

Hence, this equation can be represented as: one mole of N_2 reacts with three moles of H_2 to produce two moles of NH_3 .

Whenever we carry out stoichiometric calculations (relating masses of products to masses of reactants), we will always follow the 4 steps given below.

- 1) Write the balanced chemical equation.
- 2) Convert masses given for reactants and products to moles, using the molar masses.
- 3) Write down the **mole ratio** using the stoichiometric coefficients of the balanced chemical equation.

$$\text{Mole Ratio} = \frac{\text{Moles Desired}}{\text{Moles Given}}$$

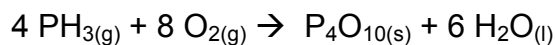
Moles Desired: Stoichiometric coefficient of the substance for which we want to calculate the amount reacted or produced.

Moles Given: Stoichiometric coefficient of the substance for which we know the amount reacted or produced.

- 4) Multiply the mole ratio by the number of moles given in the problem or calculated in step 2. Then, convert the calculated number of moles to the mass if necessary.

Example 1: Consider the following reaction: $\text{PH}_{3(\text{g})} + \text{O}_{2(\text{g})} \rightarrow \text{P}_4\text{O}_{10(\text{s})} + \text{H}_2\text{O}_{(\text{l})}$. How many moles of $\text{PH}_{3(\text{g})}$ are required for the production of 3.48 mol (P_4O_{10})?

- 1) Balance the chemical equation:



2) Convert masses to moles (already done).

3) Write the Mole Ratio:

“How many moles of $\text{PH}_3(\text{g})$ ” implies what is desired is $\text{PH}_3(\text{g})$.

”Production of 3.48 mol (P_4O_{10})” implies what is given is $\text{P}_4\text{O}_{10}(\text{s})$.

$$\text{Mole Ratio} = \frac{4 \text{ mol (PH}_3\text{)}}{1 \text{ mol (P}_4\text{O}_{10}\text{)}}$$

4) Multiply the mole ratio by the number of moles given in the problem:

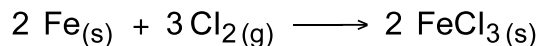
$$\frac{4 \text{ mol (PH}_3\text{)}}{1 \text{ mol (P}_4\text{O}_{10}\text{)}} \times 3.48 \text{ mol (P}_4\text{O}_{10}\text{)} = 13.92 \text{ mol PH}_3$$

Example 2: Considering the same chemical reaction, how many moles of $\text{O}_2(\text{g})$ react with 16.3 moles of PH_3 ?

To solve this problem, let us use the conversion factor (mole ratio).

$$\text{mol O}_2 = \frac{8 \text{ mol (O}_2\text{)}}{4 \text{ mol (PH}_3\text{)}} \times 16.3 \text{ mol (PH}_3\text{)} = 32.6 \text{ mol O}_2$$

Example 3: Iron reacts with chlorine gas to form iron(III) chloride. Calculate the mass of iron(III) chloride produced from 12.24 moles of iron.



The reactant Fe is given in moles:

$$\begin{aligned} \text{In the chemical equation there are : } & \frac{2 \text{ mol FeCl}_3}{2 \text{ mol Fe}} \times 12.24 \text{ mol Fe} \\ & = 12.24 \text{ mol FeCl}_3 \end{aligned}$$

However, the problem asks for the mass of FeCl_3 .

mass = moles x molar mass

mass = 12.24 mol x 162.4 g/mol

mass = 1988 g

Example 4: Silicon dioxide heated with excess of carbon (coke) produces pure silicon and carbon monoxide. How many grams of carbon monoxide are formed when 32.55 g of silicon are produced?



In the chemical equation there are: $\frac{2 \text{ mol CO}}{1 \text{ mol Si}} \times 1.158 \text{ mol Si}$
 $= 2.316 \text{ mol CO}$

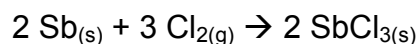
However, the problem asks for the mass of CO.

mass = moles x molar mass
 $= 2.316 \text{ mol} \times 28.0 \text{ g/mol}$
 $= 64.90 \text{ g CO}$

In Section 2.21, practice the Interactive Problems.

Sections 2.22 - 2.23: Theoretical Yield, Limiting Reactant and Percent Yield

Consider the chemical reaction:



The reaction equation provides a recipe for the preparation of antimony chloride. It indicates that one needs 3 moles of chlorine gas for every two moles of antimony solid to form 2 moles of antimony chloride.

Whenever reactants are present in the relative amount of 3 moles of chlorine for 2 moles of antimony, we say that the reactants are present in the **stoichiometric amount**.

For example, if we are given 0.4 mol Sb with 0.6 mol Cl₂, we have reactants present in the stoichiometric amount (ratio of 0.4 to 0.6 = ratio of 2 to 3).

Whenever we have reactants present in the stoichiometric amount, we can use any of the reactants to calculate how many moles of product are formed from a given amount of reactants.

The maximum mass of products that can be formed from a given amount of reactants, assuming the reaction is complete, is called the **Theoretical Yield**.

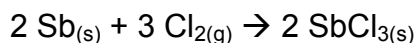
For example: If we react 2 moles of Sb with 3 moles of Cl₂, we obtain 2 moles of SbCl₃. The mass of 2 moles of SbCl₃ is equal to 2 × (121.8 + 3 × 35.5) = 456.6 g. We say that when the reaction is carried out with 2 moles of Sb and 3 moles of Cl₂, the theoretical yield is equal to 456.6 g.

When the reaction is carried out with 4 moles Sb and 6 moles Cl₂, the theoretical yield is equal to 913.2 g. When the reaction is carried out with 0.2 moles Sb and 0.3 moles Cl₂, the theoretical yield is 45.7 g

When the reactants are present in the stoichiometric amount and the reaction is complete, there is no reactant left at the end of the reaction.

What happens if reactants are not present in the stoichiometric amount? When reactants are not present in the stoichiometric amount, then, some of the reactants are completely consumed in the reaction (**limiting reactant**) and some reactants remain at the end of the reaction (**excess reactant**).

For example: Consider the reaction between antimony and chlorine gas, again.



When the reaction is started with 4 moles Sb and 6 moles Cl₂, we said above that the reactants are present in the stoichiometric amount.

When the reaction is started with 4 moles Sb and 7 moles Cl₂, then, at the end of the reaction, there will be one mole of Cl₂ left, since to completely react 4 moles Sb, only requires 6 moles Cl₂. We say that Cl₂ is in excess and Sb is the limiting reactant (it is the one limiting the amount of product that can be formed).

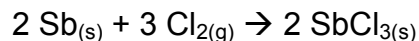
When the reaction is started with 5 moles Sb and 6 moles Cl₂, then, at the end of the reaction, there will be one mole of Sb left, since to completely react 6 moles Cl₂, only requires 4 moles Sb. We say that Sb is in excess and Cl₂ is the limiting reactant (it is the one limiting the amount of product that can be formed).

Remember that the theoretical yield is always determined by the amount of limiting reactant.

Example: Assume that the reaction between antimony and chlorine is carried out with 1.5 mol Sb and 1.5 mol Cl₂. Determine the limiting reactant, the excess reactant and the theoretical yield.

To determine which reactant is the limiting reactant, we calculate the amount of product formed from each of the given amounts of reactants. The limiting reactant is the reactant that produces the least amount of products.

1) Write the balanced chemical reaction.



2) Determine the amounts of product formed from each reactant:

How much SbCl₃ is formed with 1.5 mol Sb?

Using the mole ratio derived from the balanced chemical reaction, we write:

$$\text{moles (SbCl}_3) = \frac{2 \text{ mol SbCl}_3}{2 \text{ mol Sb}} \times 1.5 \text{ mol Sb} = 1.5 \text{ mol SbCl}_3$$

How much SbCl₃ is formed with 1.5 mol Cl₂?

Using the mole ratio derived from the balanced chemical reaction, we write:

$$\text{moles (SbCl}_3) = \frac{2 \text{ mol SbCl}_3}{3 \text{ mol Cl}_2} \times 1.5 \text{ mol Cl}_2 = 1.0 \text{ mol SbCl}_3$$

3) The limiting reactant is the reactant producing the smallest amount of product.

The limiting reactant is Cl₂. Hence, the excess reactant is Sb.

The theoretical yield is the amount of products formed from the limiting reactant. 1.5 mol Cl₂ produce 1.0 mol SbCl₃. The mass of 1.0 mol SbCl₃ is equal 228.3 g.

Hence, the theoretical yield is 228.3 g.

We found that the limiting reactant is chlorine gas and the excess reactant is antimony. So we know that all chlorine (1.5 mol) is consumed in the reaction. Now, let us calculate how much antimony is left after the reaction.

First, we calculate how much antimony is used to fully react with 1.5 mol Cl₂. We do this using the mole ratio.

$$\text{moles (Sb)} = \frac{2 \text{ mol Sb}}{3 \text{ mol Cl}_2} \times 1.5 \text{ mol Cl}_2 = 1.0 \text{ mol Sb}$$

Hence, 1.0 mol Sb is used out of an initial amount of 1.5 mol Sb. We conclude that 0.5 mol Sb remains after the reaction.

Experimental Yield, Theoretical Yield and Percent Yield

For a variety of reasons, the vast majority of chemical reactions carried out in laboratories or in chemical plants do not produce the maximum amount of product possible (theoretical yield).

Remember the **Theoretical Yield** is calculated from the amounts of reactants present and from the balanced chemical reaction.

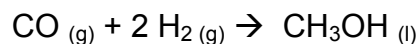
The amount of products actually obtained from a given amount of reactants is called the **Experimental Yield**. The experimental yield is thus always smaller than the theoretical yield.

The **Percent Yield** is defined as the ratio of experimental yield to theoretical yield multiplied by 100.

$$\% \text{ Yield} = 100 \times \frac{\text{Experimental Yield}}{\text{Theoretical Yield}}$$

Example: 68.5 g of CO_(g) is reacted with 8.60 g of H_{2(g)} to form 35.7 g of CH₃OH (methanol). Determine if any reactant is in excess. Calculate the theoretical yield and the percent yield.

- 1) Write the balanced chemical reaction.



- 2) Calculate the amount of methanol formed from 68.5 g CO.

Use the method of conversion factors.

$$\text{g}(\text{CH}_3\text{OH}) = 68.5 \text{ g CO} \times \frac{1 \text{ mol (CO)}}{28.0 \text{ g (CO)}} \times \frac{1 \text{ mol (CH}_3\text{OH)}}{1 \text{ mol (CO)}} \times \frac{32.0 \text{ g (CH}_3\text{OH)}}{1 \text{ mol (CH}_3\text{OH)}}$$

$$\text{g}(\text{CH}_3\text{OH}) = 78.2 \text{ g (CH}_3\text{OH)}$$

3) Calculate the amount of methanol formed from 8.60 g H₂.

Use the method of conversion factors.

$$\text{g}(\text{CH}_3\text{OH}) = 8.60 \text{ g H}_2 \times \frac{1 \text{ mol (H}_2\text{)}}{2.0 \text{ g(H}_2\text{)}} \times \frac{1 \text{ mol (CH}_3\text{OH)}}{2 \text{ mol (H}_2\text{)}} \times \frac{32.0 \text{ g (CH}_3\text{OH)}}{1 \text{ mol (CH}_3\text{OH)}}$$

$$\text{g}(\text{CH}_3\text{OH}) = 68.8 \text{ g (CH}_3\text{OH)}$$

4) The mass of methanol formed with 8.60 g hydrogen is less than that formed with 68.5 g carbon monoxide. Hence, hydrogen is the limiting reactant and carbon monoxide is the excess reactant. The theoretical yield is the amount of methanol formed with 8.6 mol H₂ (limiting reactant). The theoretical yield is equal to 68.8 g.

5) The percent yield is then expressed as:

$$\% \text{ Yield} = 100 \times \frac{35.7 \text{ g}}{68.8 \text{ g}} = 51.9 \%$$

In Section 2.23, practice Interactive Problems.

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