

Introductory Chemistry - UNIT 7 SUPPLEMENT

CHEMICAL STOICHIOMETRY

STOICHIOMETRY - THE MEANING OF THE WORD

The word stoichiometry derives from two Greek words: *stoicheion* (meaning "element") and *metron* (meaning "measure"). Stoichiometry deals with calculations about the masses (sometimes volumes) of reactants and products involved in a chemical reaction. It is a very mathematical part of chemistry, so be prepared for lots of calculator use.

Objectives for Unit 7:

- ✓ To understand a chemical equation and the relationships between reactants and products.
- ✓ To convert mole and mass amounts in a chemical equation (Stoichiometry).
- ✓ To perform percent yield calculations.
- ✓ To perform limiting reactant problems.
- ✓ To understand the unit of energy and how it works in a chemical reaction.

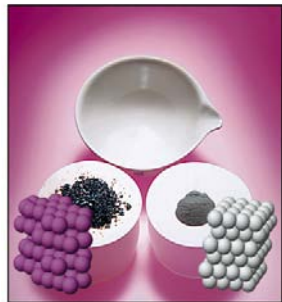
CONVERSION FACTORS FROM A CHEMICAL EQUATION

STOICHIOMETRY - HISTORY



Jeremias Benjamin Richter, a chemist in 1774 became a mining official in Germany. He performed some of the earliest determinations of the quantities by weight in which acids saturate bases. He proposed the Law of Definite Proportions which he stated that the ratio by weight of the compounds consumed in a chemical reaction is always the same. In his published works he introduced the term **STOICHIOMETRY** which he defined as *the art of chemical measurements, which has to deal with the laws according to which substances unite to form chemical compounds.*

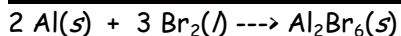
REVIEW: CHEMICAL REACTIONS/EQUATIONS



Reactants: $\text{Zn} + \text{I}_2 \rightarrow$ Product: ZnI_2

Chemical Equations

stoichiometric coefficients



This equation means

2 Al atoms + 3 Br₂ molecules \rightarrow 1 molecule of Al₂Br₆

2 moles of Al + 3 moles of Br₂ \rightarrow 1 mole of Al₂Br₆

The letters (s), (g), and (l) are the physical states of compounds.

Why is this important?

As fascinating as it was to find out we could measure out 27.0 grams of aluminum and know it was exactly 6.02×10^{23} atoms of aluminum, chemical stoichiometry is just as amazing.

Before we let elements and compounds meet, we can predict how they are going to react, and even further, how much stuff they are going to produce. We could also be on the other end of things and find a waste product and determine exactly how much of a particular compound or element was needed to produce the waste. How WONDERFUL!! This is STOICHIOMETRY!!

REACTION STOICHIOMETRY involves the mass and mole relationships among reactants and products in a chemical reaction. A balanced chemical equation is used to determine these relationships.

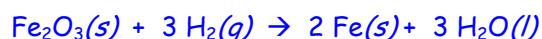
Note: The stoichiometric coefficients in a balanced equation give the relative numbers of atoms, or molecules, or the number of moles of each substance reacting.

MOLE - MOLE RELATIONSHIPS involves mole conversions between substances in a chemical equation. To solve:

Write and Balance the chemical equation.

Convert from moles of the given material to moles of the unknown material (or moles of the given material to grams of the unknown material), using the balanced equation. *bridge step*

A reaction with iron oxide:



The chemical equation indicates:

1 mole Fe_2O_3 and 3 moles H_2 react, 2 moles Fe and 3 moles H_2O are produced

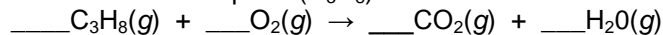
How many moles of hydrogen are needed to react to produce 15.65 moles of iron?

Perform Step 2 (bridge step): mol Fe to mol H_2

$$15.65 \text{ mol Fe} \times \frac{3 \text{ mol H}_2}{2 \text{ mol Fe}} = 23.48 \text{ mol H}_2$$

Another Example:

Combustion of Propane (C_3H_8).



What does the Chemical Equation tell us?



When the equation is balanced: 3 moles C_3H_8 react with 5 moles O_2 , 3 moles CO_2 and 4 moles H_2O are produced.

Calculate the number of moles of oxygen reacting if 3.00 moles of CO_2 is produced.

$$3.00 \text{ mol CO}_2 \times \frac{5 \text{ mol O}_2}{3 \text{ mol CO}_2} = 5.00 \text{ mol O}_2$$

Calculate the number of moles of water produced if 2.00 moles of propane (C_3H_8) is burned.

$$2.00 \text{ mol C}_3\text{H}_8 \times \frac{4 \text{ mol H}_2\text{O}}{1 \text{ mol C}_3\text{H}_8} = 8.00 \text{ mol H}_2\text{O}$$

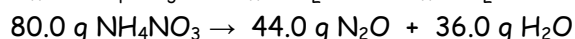
Calculate the number of moles of carbon dioxide produced if 4.60 moles of oxygen is needed to react

$$\text{with propane. } 4.60 \text{ mol O}_2 \times \frac{3 \text{ mol CO}_2}{5 \text{ mol O}_2} = 2.76 \text{ mol CO}_2$$



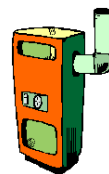
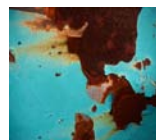
MASS - MASS RELATIONSHIPS involve gram conversions between substances in a chemical equation. The same steps provided above may be followed. Example:

If 454 g of NH_4NO_3 decomposes into N_2O and H_2O . What is the mass of H_2O produced? Balanced Chemical Equation: $\text{NH}_4\text{NO}_3 \rightarrow \text{N}_2\text{O} + 2 \text{H}_2\text{O}$

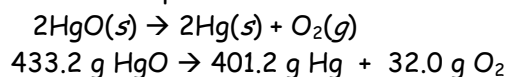


Convert g NH_4NO_3 to g H_2O (bridge step) (finding the theoretical yield)

$$454 \text{ g NH}_4\text{NO}_3 \times \frac{36.0 \text{ g H}_2\text{O}}{80.0 \text{ g NH}_4\text{NO}_3} = 205 \text{ g H}_2\text{O}$$



Other Examples:



Let's Put It In One Step

5.00 g mercury (II) oxide decomposes to mercury and oxygen. Determine the mass of mercury produced.

$$5.00\text{g HgO} \times \frac{401.2 \text{ g Hg}}{433.2 \text{ g HgO}} = 4.63\text{g Hg}$$



PERCENT YIELD

PERCENT YIELD: Experimentally, quantities of products are determined that may not always match those quantities of products determined through stoichiometric calculations. The percentage of product obtained experimentally can be determined by percent yield.

$$\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}}$$

PERCENT YIELD tells the efficiency of the reaction.

ACTUAL YIELD is the amount of product really obtained (by experimentation).

THEORETICAL YIELD is the amount of product that should be obtained (by calculation).

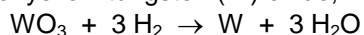
Example:

A little history, to illustrate a reaction of incandescent light bulbs.



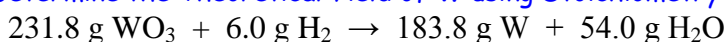
It's a little known fact that if Lewis Latimer hadn't been on Thomas Edison's team (the only African American) that the light bulb might have been too impractical for most people to use. You see, Edison's light bulb used bamboo as the filament and the light bulb burned out after only 30 hours. Carbon lasted much longer and made the light bulb practical. Latimer, with Joseph V. Nichols, came up with both idea to use carbon filaments and the process for manufacturing the carbon filaments.

At present, tungsten metal, W is used to make incandescent bulb filaments. The metal is separated out of yellow tungsten (VI) oxide, WO_3 by reacting with hydrogen according to the following reaction:



If 2.76 grams of W is experimentally obtained from the reaction between 15.8 grams of WO_3 and excess of H_2 , calculate the percent yield obtained from the reaction.

First, Determine the Theoretical Yield of W using Stoichiometry



$$15.8\text{g WO}_3 \times \frac{183.8 \text{ g W}}{231.8\text{g WO}_3} = 12.5\text{g W}$$

Second, Determine the Percent Yield of W

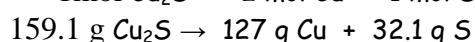
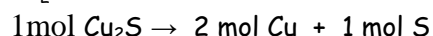
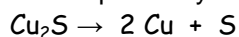
$$\text{Percent Yield} = \frac{2.76 \text{ g W}}{12.5 \text{ g W}} \times 100 = 22.0\%$$

Thus, only 22.0% of the tungsten was recovered.



Another Example: Copper can be recovered from a mineral in our environment: Cu_2S (called chalcocite). When I performed this experiment, I started with 15.0 grams of Cu_2S and recovered 9.95 grams of copper. What was the percent yield of copper from my experiment?

Write the chemical equation:



First, Determine the Theoretical Yield of copper using stoichiometry.

$$15.0 \text{ g Cu}_2\text{S} \times \frac{127 \text{ g Cu}}{159.1 \text{ g Cu}_2\text{S}} = 12.0 \text{ g Cu}$$

Second, Determine the Percent Yield of W

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

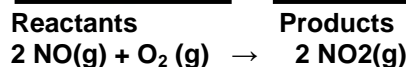
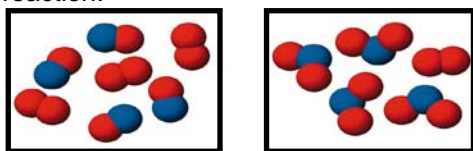
$$\text{Percent Yield} = \frac{9.95 \text{ g}}{12.0 \text{ g}} \times 100 = 82.9\% \text{ Cu}$$

Thus, only 82.9% copper was recovered.

LIMITING REACTANT PROBLEMS

LIMITING REACTANT is the reactant that is completely used up in a chemical reaction. It limits the amount of other reactants that can combine and the amount of products formed in a reaction.

EXCESS REACTANT is the reactant in excess in a chemical reaction. It is not completely used up in a reaction.



Limiting reactant = NO

Excess reactant = O₂

Why is this important?

In our age of environmental awareness, we try to minimize our waste production. The same is true in chemical processes. We learned with stoichiometry that we can determine the exact amount of product that will be produced when using a known amount of reactant. Well, we've seen in many chemical reactions, that often more than one reactant is used. How can we minimize using too much of our reactants to produce a certain amount of product?

Remember the following things about Limiting Reactants:

- The reaction will stop when the reactants are used up.
- If one reactant is used up before the other, the reaction will stop.
- The first reactant used up is the **LIMITING REACTANT**, use it for the calculation.
- The other reactant is the **EXCESS REACTANT**.

Example:

Mix 5.40 g of Al with 8.10 g of Cl₂. What mass of Al₂Cl₆ can form?



Convert g Al to g Al₂Cl₆ and convert g Cl₂ to g Al₂Cl₆.

$$5.40 \text{ g Al} \times \frac{1 \text{ mol Al}}{27.0 \text{ g Al}} = 0.200 \text{ mol Al}$$

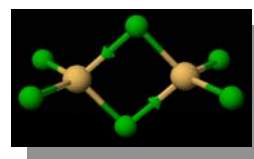
$$8.10 \text{ g Cl}_2 \times \frac{1 \text{ mol Cl}_2}{71.0 \text{ g Cl}_2} = 0.114 \text{ mol Cl}_2$$

Step 2 (bridge step) convert mol Al to mol Al₂Cl₆ and convert mol Cl₂ mol Al₂Cl₆.

Step 3 The reactant that produced the smaller mol Al₂Cl₆ is the L.R..

$$0.200 \text{ mol Al} \times \frac{1 \text{ mol Al}_2\text{Cl}_6}{2 \text{ mol Al}} = 0.100 \text{ mol Al}_2\text{Cl}_6$$

$$0.114 \text{ mol Cl}_2 \times \frac{1 \text{ mol Al}_2\text{Cl}_6}{3 \text{ mol Cl}_2} = 0.0380 \text{ mol Al}_2\text{Cl}_6$$



Cl₂ is the limiting reactant and Al is the excess reactant.

Step 4: Calculate mass of Al₂Cl₆ expected, using the limiting reactant.

$$0.0380 \text{ mol Al}_2\text{Cl}_6 \times \frac{267.0 \text{ g Al}_2\text{Cl}_6}{1 \text{ mol Al}_2\text{Cl}_6} = 10.1 \text{ g Al}_2\text{Cl}_6$$

How much of excess reactant will remain when reaction is complete?

Step 1: Determine the amount of excess reactant (Al) is actually needed in the reaction.

$$0.114 \text{ mol Cl}_2 \times \frac{2 \text{ mol Al}}{3 \text{ mol Cl}_2} = 0.0760 \text{ mol Al}$$

$$0.0760 \text{ mol Al} \times \frac{27.0 \text{ g Al}}{1 \text{ mol Al}} = 2.05 \text{ g Al}$$

Step 2: Subtract this amount from the amount of Al the reaction started with.

5.40 g Al - 2.05 g Al = 3.35 g Al is in excess when the reaction is complete.

Other Example:



To See a movie of the Shuttle Launch: [Click Here](#)

The fuel in the main cells of the Space Shuttle are liquid oxygen and liquid hydrogen. The explosive reaction results in the production of H₂O. Obviously, those working on the space shuttle had to have an understanding of limiting reactant problems.

The reaction: $2 \text{ H}_2 + \text{ O}_2 \rightarrow 2 \text{ H}_2\text{O}$

If 45,450 grams of oxygen and 45,620 grams of hydrogen were put into the fuel cells of the space shuttle.

Which reactant is the limiting reactant?

$$45,450 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.0 \text{ g O}_2} \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} = 2,841 \text{ mol H}_2\text{O}$$

$$45,620 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.0 \text{ g H}_2} \times \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2} = 22,810 \text{ mol H}_2\text{O}$$

Oxygen is the limiting reactant.

How much H₂O will be produced?

$$2,841 \text{ mol H}_2\text{O} \times \frac{18.0 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 51,140 \text{ g H}_2\text{O}$$

How much excess reactant was not needed to be put into the space shuttle?

$$45,450 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.0 \text{ g O}_2} \times \frac{2 \text{ mol H}_2}{1 \text{ mol O}_2} \times \frac{2.0 \text{ g H}_2}{1 \text{ mol H}_2} = 5,681 \text{ g H}_2$$

45,620 g H₂ - 5,681 g H₂ = 39,940 g H₂ in excess.