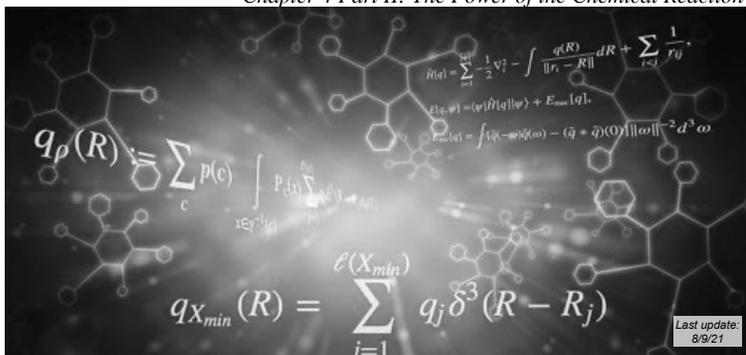


Chemistry 151: Basic Chemistry

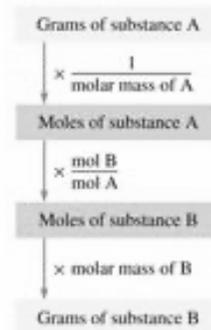
Chapter 4 Part II: The Power of the Chemical Reaction



The Power of Chemical Reactions

A balanced chemical reaction will show the relative amounts of reactants and products. In this section we will apply the balanced reaction to "real world" situations whereby quantities of products created or reactants needed can be predicted... and more!

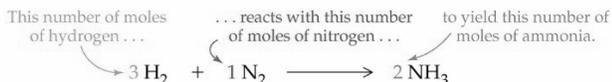
This section exemplifies why chemists get paid: bosses want to know 'how much' plastic will be made for cell phones, they do not care about moles (lol).... this is an important chapter!



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Coefficients in Chemical Equations

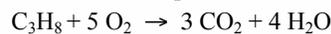
Coefficients in a *balanced* chemical equation tell how many molecules (and thus how many moles) of each reactant are needed *and* how many molecules (and thus moles) of each product are formed.



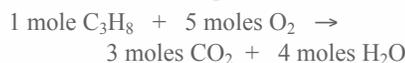
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Coefficients in Chemical Equations

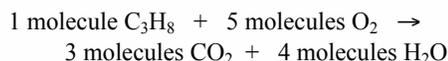
In the equation:



this equation means



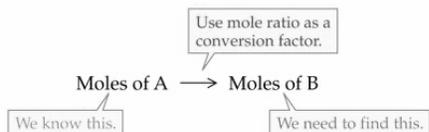
OR



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Converting Moles in Equations

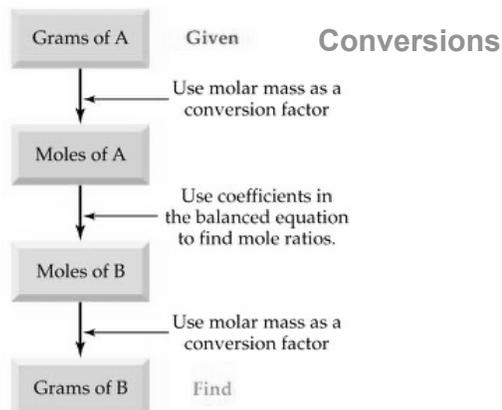
We can use a mole ratio from a chemical equation to convert mol (*or* g) of A into mol (*or* g) of B



This is useful in determining how much product is created from so much reactant

Also used for determining how much reactant necessary to create so much product

MAR



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Example

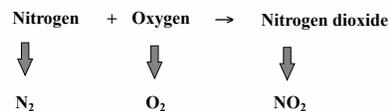
Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

MAR

Example

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

Translate Word Equation

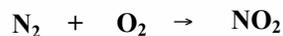


MAR

Example

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

Translate Word Equation



...but note that there are 2 N and 2 O reactants with only 1 N and 2 O products; need to *balance* equation

MAR

Example

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

Balance the equation



Now there are 2 N and 4 O reactants with 2 N and 4 O products; equation is *balanced*

MAR

Example:

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

Convert mass of compound available (nitrogen) to moles.

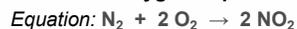
$$133 \text{ g N}_2 \cdot \frac{1 \text{ mol N}_2}{28.0 \text{ g N}_2} = 4.75 \text{ mol N}_2$$

MAR

Example:

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

Now relate the moles of nitrogen available to moles of oxygen required.



$$4.75 \text{ mol N}_2 \cdot \frac{2 \text{ mol O}_2 \text{ required}}{1 \text{ mol N}_2 \text{ available}} =$$

9.50 mol O₂ required

MAR

Example:

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

Convert moles of oxygen consumed to grams.

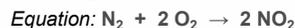
$$9.50 \cancel{\text{ mol O}_2} \text{ required} * \frac{32.0 \text{ g O}_2}{1 \cancel{\text{ mol O}_2}} = 304 \text{ g O}_2 \text{ required to combust 133 g N}_2$$

MAR

Example:

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

Relate the moles of nitrogen available to moles of NO₂ produced.



$$4.75 \cancel{\text{ mol N}_2} * \frac{2 \text{ mol NO}_2 \text{ produced}}{1 \cancel{\text{ mol N}_2} \text{ available}} = 9.50 \text{ mol NO}_2 \text{ produced}$$

MAR

Example:

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

Convert moles of NO₂ produced to grams.

$$9.50 \cancel{\text{ mol NO}_2} * \frac{46.0 \text{ g NO}_2}{1 \cancel{\text{ mol NO}_2}} = 437 \text{ g NO}_2 \text{ created when 133 g N}_2 \text{ burned in O}_2$$

MAR

Test Yourself

How many grams of water will be produced if 6.000*10⁹ molecules of sulfur dioxide are created? The equation:



MAR

Percent Yield

Actual yield - The quantity of product (g) actually obtained in a reaction.

Theoretical yield - The quantity of product (g) that is expected from a chemical reaction.

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

MAR

Example - Percent Yield

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

The theoretical yield of NO₂ was 437 g (see previous example)

In the actual experiment, only 247 g of NO₂ was recovered; this is the actual yield

$$\begin{aligned} \text{Percent yield} &= (\text{actual} / \text{theoretical}) * 100\% \\ &= (247 / 437) * 100\% \\ &= 56.5\% \end{aligned}$$

MAR

Practice, practice, practice!

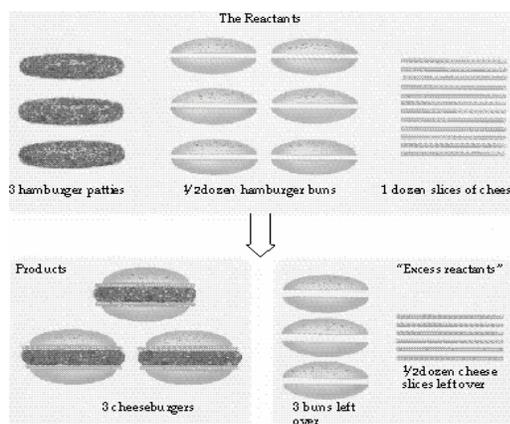
Limiting Reactant

Most of the time, one reactant limits how much product can be produced.

This reactant is called the **limiting reactant**, and the other reactant(s) is called the **reactant in excess** or **excess reactant**.

The limiting reactant *limits* the amount of product that can be made and hence controls the reaction.

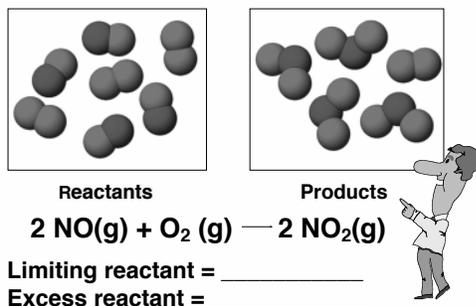
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Limiting Reactant



Limiting Reactant

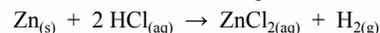


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Example - Limiting Reactant

Zinc metal (7.00 g) reacts with 0.100 mol HCl to make ZnCl_2 and H_2 gas. Find the limiting reactant, the reactant in excess, and find the theoretical yield of ZnCl_2 .

balanced chemical equation:



MAR

Example - Limiting Reactant

Zinc metal (7.00 g) reacts with 0.100 mol HCl to make ZnCl_2 and H_2 gas. Find the limiting reactant, the reactant in excess, and find the theoretical yield of ZnCl_2 .

$$\text{Zn}_{(\text{s})} + 2 \text{HCl}_{(\text{aq})} \rightarrow \text{ZnCl}_{2(\text{aq})} + \text{H}_{2(\text{g})}$$

First, convert Zn to mol to use mol ratio:

$$7.00 \text{ g Zn} * (\text{mol Zn} / 65.4 \text{ g Zn}) = 0.107 \text{ mol Zn}$$

Second, convert mol Zn and mol HCl to mol ZnCl_2 .

$$0.107 \text{ mol Zn} * (\text{mol ZnCl}_2 / \text{mol Zn}) = 0.107 \text{ mol ZnCl}_2$$

$$0.100 \text{ mol HCl} * (\text{mol ZnCl}_2 / 2 \text{ mol HCl}) = 0.0500 \text{ mol ZnCl}_2$$

Notice unequal quantities of ZnCl_2 created!

MAR

Example - Limiting Reactant

Zinc metal (7.00 g) reacts with 0.100 mol HCl to make ZnCl_2 and H_2 gas. Find the limiting reactant, the reactant in excess, and find the theoretical yield of ZnCl_2 .

$$\text{Zn}_{(\text{s})} + 2 \text{HCl}_{(\text{aq})} \rightarrow \text{ZnCl}_{2(\text{aq})} + \text{H}_{2(\text{g})}$$

Third, compare ZnCl_2 quantities:

$$7.00 \text{ g Zn} = 0.107 \text{ mol Zn} \text{ which gives } 0.107 \text{ mol ZnCl}_2$$

$$0.100 \text{ mol HCl} \text{ gives } 0.0500 \text{ mol ZnCl}_2$$

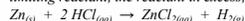
HCl gives *less* ZnCl_2 than Zn! Hence, HCl is *limiting* how much product can be made, and HCl is the **limiting reactant**.

Zn could produce a lot more ZnCl_2 than HCl, but HCl cannot keep up. Hence, Zn is the reactant in excess.

MAR

Example - Limiting Reactant

Zinc metal (7.00 g) reacts with 0.100 mol HCl to make ZnCl₂ and H₂ gas. Find the limiting reactant, the reactant in excess, and find the theoretical yield of ZnCl₂.



Fourth, find theoretical yield of ZnCl₂.

Need to use limiting reactant to calculate theoretical yield of ZnCl₂.

$$0.0500 \text{ mol ZnCl}_2 * (136.3 \text{ g ZnCl}_2 / \text{mol ZnCl}_2) = 6.82 \text{ g ZnCl}_2, \text{ the theoretical yield}$$

If we used the 0.107 mol ZnCl₂ from Zn:

$$0.107 \text{ mol ZnCl}_2 * (136.3 \text{ g ZnCl}_2 / \text{mol ZnCl}_2) = 14.6 \text{ g ZnCl}_2, \text{ which is not possible!}$$

Practice, practice, practice!

MAR

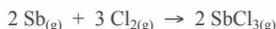
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Test Yourself Part 1

Write a balanced reaction for the formation of gaseous antimony(III) chloride from gaseous antimony and chlorine gas.

Test Yourself Part 2

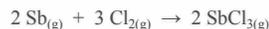
Determine the limiting reactant and theoretical yield of SbCl₃ if 129 g of Sb and 106 g of Cl₂ are mixed.



MAR

Test Yourself Part 3

Only 113.5 g of SbCl₃ were collected. Calculate the percent yield for the reaction. (Theoretical yield = 227 g)



MAR

Test Yourself Part 4

How much excess reactant is left at the end of the reaction if 129 g of Sb and 106 g of Cl₂ are mixed? $2 \text{Sb}_{(g)} + 3 \text{Cl}_{2(g)} \rightarrow 2 \text{SbCl}_{3(g)}$

MAR

Practice, practice, practice!

End of Chapter 4 Part II

