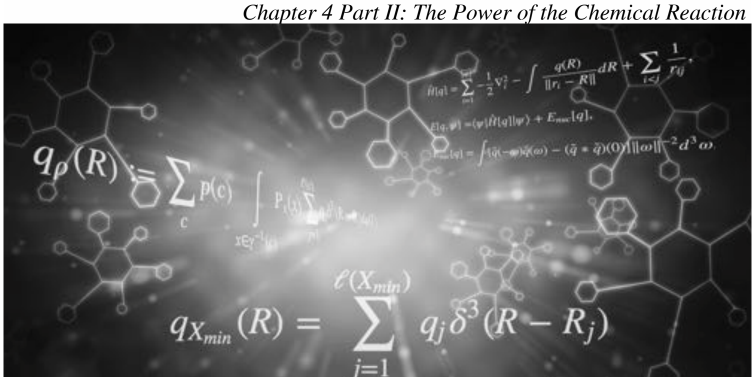
Chemistry 151: Basic Chemistry

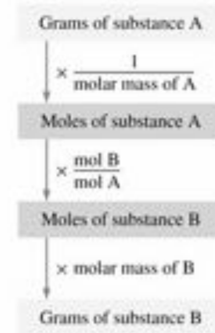
Chapter 4 Part II: The Power of the Chemical Reaction



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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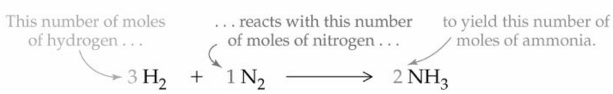
Coefficients in Chemical Equations

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:

$C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$

this equation means

1 mole C_3H_8 + 5 moles O_2 →
3 moles CO_2 + 4 moles H_2O

OR

1 molecule C_3H_8 + 5 molecules O_2 →
3 molecules CO_2 + 4 molecules H_2O

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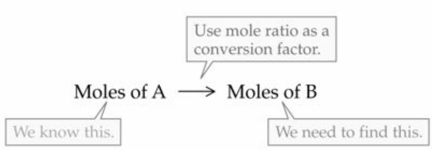
Converting Moles in Equations

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

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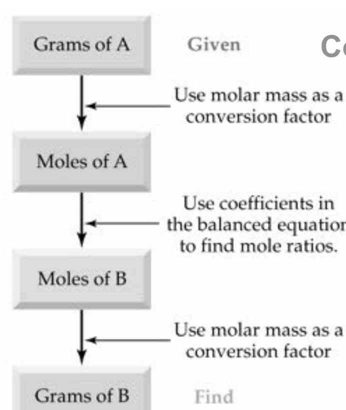
Conversions



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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?

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Translate Word Equation

Nitrogen + Oxygen → Nitrogen dioxide

$$N_2 \qquad O_2 \qquad NO_2$$

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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?

Translate Word Equation

$$N_2 + O_2 \rightarrow NO_2$$

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

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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?

Balance the equation

$$N_2 + O_2 \rightarrow NO_2$$

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

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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?

Convert mass of compound available (nitrogen) to moles.

$$133 \cancel{\text{ g } N_2} * \frac{1 \text{ mol } N_2}{28.0 \cancel{\text{ g } N_2}} = 4.75 \text{ mol } N_2$$

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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?

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

Equation: $N_2 + 2\ O_2 \rightarrow 2\ NO_2$

$$4.75 \cancel{\text{ mol } N_2} * \frac{2 \text{ mol } O_2 \text{ required}}{1 \cancel{\text{ mol } N_2} \text{ available}} =$$

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9.50 mol O_2 required

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 g O_2 required to combust 133 g N_2

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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?

Relate the moles of nitrogen available to moles of NO_2 produced.

Equation: $N_2 + 2\ O_2 \rightarrow 2\ NO_2$

$$4.75 \cancel{\text{mol } N_2} * \frac{2 \text{ mol } NO_2 \text{ produced}}{1 \cancel{\text{mol } N_2} \text{ available}} =$$

9.50 mol NO_2 produced

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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?

Convert moles of NO_2 produced to grams.

$$9.50 \cancel{\text{mol } NO_2} * \frac{46.0 \text{ g } NO_2}{1 \cancel{\text{mol } NO_2}} =$$

437g NO_2 created when 133 g N_2 burned in O_2

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Test Yourself

How many grams of water will be produced if $6.000*10^9$ molecules of sulfur dioxide are created? The equation:

$$2\ H_2SO_4 + Cu \rightarrow SO_2 + 2\ H_2O + CuSO_4$$

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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.

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

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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 NO_2 was 437 g (see previous example)

In the actual experiment, only 247 g of NO_2 was recovered; this is the actual yield

Percent yield = (actual / theoretical) * 100%

= (247 / 437) * 100%

= 56.5%

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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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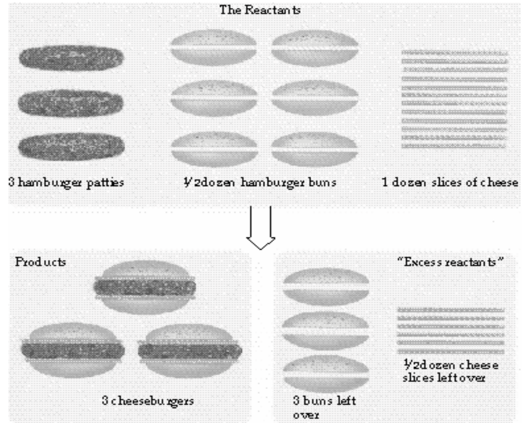
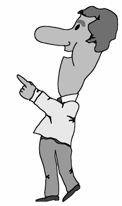
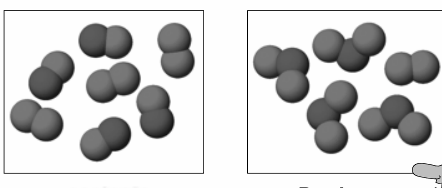
Limiting Reactant

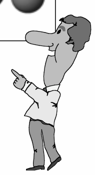




Limiting Reactant



Reactants **Products**



$2\ NO(g) + O_2\ (g) \rightarrow 2\ NO_2(g)$

Limiting reactant = ___________

Excess 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:

$$Zn_{(s)} + 2\ HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

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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 $ZnCl_2$.

$Zn_{(s)} + 2\ HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$

First, convert Zn to mol to use mol ratio:

7.00 g Zn * (mol Zn / 65.4 g Zn) = 0.107 mol Zn

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

0.107 mol Zn * (mol $ZnCl_2$ / mol Zn) = 0.107 mol $ZnCl_2$

0.100 mol HCl * (mol $ZnCl_2$ / 2 mol HCl) = 0.0500 mol $ZnCl_2$

Notice unequal quantities of $ZnCl_2$ created!

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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 $ZnCl_2$.

$Zn_{(s)} + 2\ HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$

Third, compare $ZnCl_2$ quantities:

7.00 g Zn = 0.107 mol Zn which gives 0.107 mol $ZnCl_2$

0.100 mol HCl gives 0.0500 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**.

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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 $ZnCl_2$.

$Zn_{(s)} + 2\ HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$

Fourth, find theoretical yield of $ZnCl_2$.

Need to use limiting reactant to calculate theoretical yield of $ZnCl_2$.

0.0500 mol $ZnCl_2$ * (136.3 g $ZnCl_2$ / mol $ZnCl_2$) =

6.82 g $ZnCl_2$, the **theoretical yield**

If we used the 0.107 mol $ZnCl_2$ from Zn:

0.107 mol $ZnCl_2$ * (136.3 g $ZnCl_2$ / mol $ZnCl_2$) =

14.6 g $ZnCl_2$, *which is not possible!*

Practice, practice, practice!

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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.

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Test Yourself Part 2

Determine the limiting reactant and theoretical yield of $SbCl_3$ if 129 g of Sb and 106 g of Cl_2 are mixed.

$2\ Sb_{(g)} + 3\ Cl_{2(g)} \rightarrow 2\ SbCl_{3(g)}$

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Test Yourself Part 3

Only 113.5 g of $SbCl_3$ were collected. Calculate the percent yield for the reaction. (Theoretical yield = 227 g)

$2\ Sb_{(g)} + 3\ Cl_{2(g)} \rightarrow 2\ SbCl_{3(g)}$

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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 Cl_2 are mixed? $2\ Sb_{(g)} + 3\ Cl_{2(g)} \rightarrow 2\ SbCl_{3(g)}$

Practice, practice, practice!

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End of Chapter 4 Part II