STOICHIOMETRY

Stoichiometry deals with the mass relationship among the reactants and products in a chemical reaction. Stoichiometry allows us to predict how much reactant we need to ɔduce a certain amount of product or how much product we can expect to get out of a certain reaction based upon how much reactant we use.

Mole Ratios

The central step in any stoichiometric calculation is the mole ratio. A mole ratio is a conversion factor that allows us to convert from a given amount of reactant or product to the desired quantity. Mole ratios are derived from balanced chemical equations.

Consider, for example, the chemical equation for the electrolysis of aluminum oxide to produce aluminum and oxygen.

$$2Al_2O_{3(l)} \longrightarrow 4Al_{(s)} + 3O_{2(g)}$$

The following mole ratios can be written for the reaction above:

$$\frac{2 \text{ mol } Al_2O_3}{4 \text{ mol Al}} \quad \text{or} \quad \frac{4 \text{ mol Al}}{2 \text{ mol } Al_2O_3} \quad \text{or} \quad \frac{4 \text{ mol Al}}{3 \text{ mol } O_2}$$

$$\frac{3 \text{ mol } O_2}{4 \text{ mol Al}} \quad \text{or} \quad \frac{2 \text{ mol } Al_2O_3}{3 \text{ mol } O_2} \quad \text{or} \quad \frac{3 \text{ mol } O_2}{2 \text{ mol } Al_2O_3}$$

Example 1: Write all of the possible mole ratios for the reaction below:

$$\underline{1}\, N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$$

$$\frac{2 \text{ mol } N_2}{3 \text{ mol } H_2} \quad \text{or} \quad \frac{3 \text{ mol } H_2}{2 \text{ mol } N_2}$$

$$\frac{2 \text{ mol } N_2}{2 \text{ mol } NH_3} \quad \text{or} \quad \frac{2 \text{ mol } NH_3}{2 \text{ mol } N_2}$$

$$\frac{3 \text{ mol } H_2}{2 \text{ mol } NH_3} \quad \text{or} \quad \frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2}$$

N-1

Problem type #1: $2KClO_3 \rightarrow 2KCl + 3O_2$

Mole-Mole Calculations (*Given* and *unknown* quantities are amounts in moles)

In a mole-mole calculation, you will calculate the number of moles of one substance that will react with or be produced from a given number of moles of another substance. This requires one step: a conversion using a mole-mole ratio.

Example 2: How many moles of oxygen are produced when 3.2 moles of potassium chlorate decomposes? The other product is potassium chloride.

$$\frac{3.2 \text{ mol } KClO_3}{} \Big| \frac{3 \text{ mol } O_2}{2 \text{ mol } KClO_3} = 4.8 \text{ mole } O_2$$

coefficients!

Example 3: Given: $2NH_4NO_3 \longrightarrow 2N_2 + 4H_2O + 1O_2$

How many moles of water are produced if 15.0 moles of oxygen are produced in the above reaction?

$$\frac{15 \text{ moles } O_2}{} \Big| \frac{4 \text{ mole } H_2O}{1 \text{ mole } O_2} = 60 \text{ mole } H_2O$$

coefficients

Problem type #2:

Mole-Mass Calculations (*Given* is an amount in moles and the *unknown* is a mass that is often expressed in grams.)

In a mole-mass calculation, you will calculate the mass of one substance that will react with or be produced from a given number of moles of another substance. This requires two steps: 1) Convert from what you are given to what you are looking for using a mole - mole ratio. 2) Convert from moles of what you are looking for to grams using molar mass. These two steps are done using one factor-label set-up.

$N2(14)=28$

Example 4: Given: $1 N_2 + 3H_2 \longrightarrow 2NH_3$

mass of N_2

What mass of nitrogen is required to produce 35.0 moles of ammonia?

$$\frac{35 \text{ moles } NH_3}{} \times \frac{1 \text{ mole } N_2}{2 \text{ mole } NH_3} \times \frac{28 \text{ g } N_2}{1 \text{ mole } N_2} = 490 \text{ g } N_2$$

coefficients!

N-2

$1\,N_2H_4 + 1\,O_2 \rightarrow 1\,N_2 + 2\,H_2O$

Example 5: Hydrazine, N_2H_4, is used as a rocket fuel. Reacting with oxygen, it produces nitrogen and water. What mass of oxygen is needed to burn 100.0 moles of hydrazine?

$$100 \text{ moles } N_2H_4 \times \frac{1 \text{ mole } O_2}{1 \text{ mole } N_2H_4} \times \frac{32 \text{ g } O_2}{1 \text{ mole } O_2} = \boxed{3200 \text{ g } O_2}$$

coefficients!

Problem type #3:

Mass-Mole Calculations (*Given* is a mass in grams and the *unknown* is an amount in moles.)

In a mass-mole calculation, you will calculate the number of moles of one substance that will react with or be produced from a given mass of another substance. This requires two steps: 1) Convert from grams of what you are given to moles using molar mass. 2) Convert from moles of what you are given to moles of what you are looking for using a mole-mole ratio. This calculation is also done in one factor-label set-up.

Example 6: Given: $2Na + 2H_20 \longrightarrow 2NaOH + 1\,H_2$

How many moles of hydrogen gas are produced if 56.7 g of Na are added to water?

$$56.7 \text{ g Na} \times \frac{1 \text{ mol Na}}{23 \text{ g Na}} \times \frac{1 \text{ mol } H_2}{2 \text{ mol Na}} = 1.23$$

coefficients!

Example 7: Given: $4Fe + 3O_2 \longrightarrow 2Fe_2O_3$

How many moles of iron are needed to produce 250.0 g of iron (III) oxide?

$$250 \text{ g} \times \frac{1 \text{ mole } Fe_2O_3}{159.6 \text{ g } Fe_2O_3} \times \frac{4 \text{ moles Fe}}{2 \text{ mol } Fe_2O_3} = \boxed{3.13 \text{ moles Fe}}$$

coefficients

$Fe^{+3}\,O^{-2}$

Fe_2O_3

Fe $2(55.8) = 111.6$
O $3(16) = 48$
159.6

Problem type #4:

Mass-Mass Calculations (*Given* is a mass in grams and the *unknown* is a mass in grams.)

In a mass-mass calculation, you will calculate the mass of one substance that will react with or be produced from a given mass of another substance. This will require three steps: 1) Convert grams of what you are given to moles using molar mass. 2) Convert from moles of what you are given to moles of what you are looking for using a mole-mole ratio. 3) Convert from moles of what you are looking for to grams using molar mass. These steps can be carried out using one factor-label set-up.

Example 8: What mass of oxygen is needed to completely burn 137.9 g of butane (C_4H_{10})?

$$2C_4H_{10} + 13O_2 \longrightarrow 8CO_2 + 10H_2O$$

$$\frac{137.9\text{g } C_4H_{10}}{} \left| \frac{1\text{mole } C_4H_{10}}{58\text{g } C_4H_{10}} \right| \frac{13 \text{ moles } O_2}{2 \text{ mol } C_4H_{10}} \left| \frac{32\text{g } O_2}{1\text{mole } O_2} \right. = \boxed{494.5\text{g } O_2}$$

C4 (12) = 48
H 10(1) = 10
58

coefficients

Example 9: Water decomposes with electricity to produce hydrogen and oxygen gas. How many grams of hydrogen can be produced if 75.0 g of water decomposes?

$$2H_2O \rightarrow 2H_2 + 1O_2$$

$$\frac{75\text{g } H_2O}{} \left| \frac{1\text{mole } H_2O}{18\text{g } H_2O} \right| \frac{2 \text{ moles } H_2}{2 \text{ mol } H_2O} \left| \frac{2\text{g } H_2}{1\text{ mol } H_2} \right. = 8.33\text{g } H_2$$

coefficients

H 2(1) = 2

H 2(1) = 2
O 1(16) = 16
18

N-4

Limiting Reactants

Important vocab!!

Limiting Reactant Vocabulary

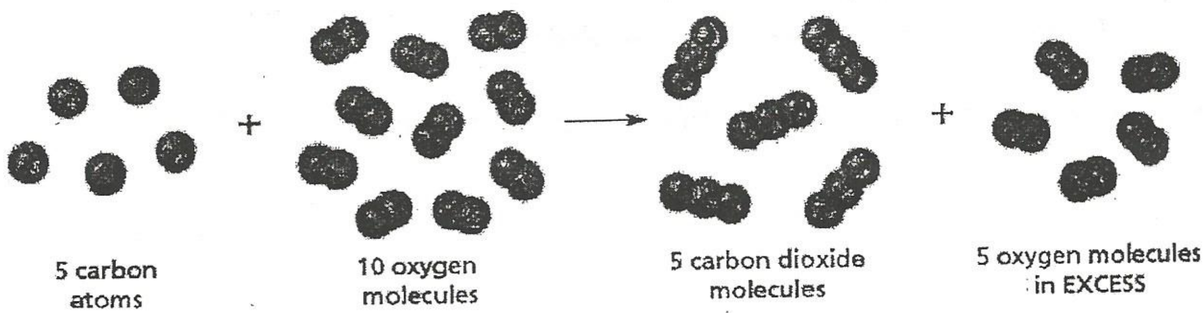
1) limiting reactant - the reactant that limits the amount of product produced by a chemical reaction. This is the reactant that will be completely consumed during the reaction. When the limiting reactant "runs out" the reaction will stop and no more product will be formed.

2) excess reactant - this is the "other" reactant that does not get completely consumed during a reaction.

Consider the reaction between carbon and oxygen to form carbon dioxide.

$$C(s) + O_2(g) \longrightarrow CO_2(g)$$

According to the equation, one mole of carbon reacts with one mole of oxygen to form one mole of carbon dioxide. Suppose you could mix 5 mol of C with 10 mol of O_2 and allow the reaction to take place. The diagram below shows that there is more oxygen than is needed to react with the carbon. Carbon is the limiting reactant in this situation, and it limits the amount of CO_2 that is formed. Oxygen is the excess reactant, and 5 l of O_2 will be left over at the end of the reaction.



NOTE: When two amounts of reactants are given in a stoichiometry problem, it must be treated as a limiting reactant problem. You must determine which reactant is "limiting" and use this quantity for all subsequent calculations! At the end of a reaction involving a limiting reactant, the reaction mixture will be composed of the product(s) and the excess reagent. The limiting reagent is completely consumed. The following examples will show you how to handle a stoichiometry problem involving a limiting reactant.

N-5

does not matter which product! Easier mass

Example 10: Given: $2Al + 6HCl \longrightarrow 2AlCl_3 + 3H_2$

a) If 3.2 moles of Al react with 10.1 moles of HCl, which reactant is limiting? Find a product amount for each given

Given 1

$$\frac{3.2\,mol\ Al}{} \times \frac{3\,mol\ H_2}{2\,mol\ Al} = 4.8\,mol\ H_2\ \text{produced}$$

*Aluminum Limits production

b) How many moles H_2 are formed?

4.8 mol H_2

Pick smaller Answer!

Given 2

$$\frac{10.1\,mol\ HCl}{} \times \frac{3\,mol\ H_2}{6\,mol\ HCl} = 5.05\,mol\ H_2\ \text{produced}$$

Not enough Aluminum to make this amount

~~c) How many mole of excess reagent remain after the reaction?~~ PREAP ONLY!

product

Example 11: Given: $1N_2 + 3H_2 \longrightarrow 2NH_3$

a) If 64.3 g of N_2 react with 16.5 g of H_2, identify the limiting reactant. Nitrogen limits product

Given #1

$$\frac{64.3\,g\ N_2}{} \times \frac{1\,mol\ N_2}{28\,g\ N_2} \times \frac{2\,mol\ NH_3}{1\,mol\ N_2} \times \frac{17\,g\ NH_3}{1\,mol\ NH_3} = 78.07\,g\ NH_3$$

b) How many grams of NH_3 will be produced?

$$\frac{16.5\,g\ H_2}{} \times \frac{1\,mol\ H_2}{2\,g\ H_2} \times \frac{2\,mol\ NH_3}{3\,mol\ H_2} \times \frac{17\,g\ NH_3}{1\,mol\ NH_3}$$

~~93.5 g NH_3~~

Not enough Nitrogen

78.07 g NH_3

*Always pick smaller Answer!

~~c) How many grams of excess reagent remain after the reaction?~~ PREAP ONLY!!

N-6

?   product

Example 12: Given: $Ca(OH)_2 + 2HCl \longrightarrow CaCl_2 + 2H_2O$

If 110.0 g of $Ca(OH)_2$ react with 96.6 g of HCl, how many grams of each product will form and ~~how many grams of excess reagent remain after the reaction stops?~~

$$\frac{110g\ Ca(OH)_2}{} \left| \frac{1mol\ Ca(OH)_2}{74g\ Ca(OH)_2} \right| \frac{2mol\ H_2O}{1mol\ Ca(OH)_2} \left| \frac{18g\ H_2O}{1mol\ H_2O} \right. = \cancel{53.5g\ H_2O}$$

$$\frac{96.6g\ HCl}{} \left| \frac{1mol\ HCl}{36.5g\ HCl} \right| \frac{2mol\ H_2O}{2mol\ HCl} \left| \frac{18g\ H_2O}{1mol\ H_2O} \right. \quad 47.63g\ H_2O$$

HCl limits production!

If produced $\frac{47.63g\ H_2O}{} \left| \frac{1mol\ H_2O}{18g\ H_2O} \right| \frac{1mol\ CaCl_2}{2mol\ H_2O} \left| \frac{113g}{1mol\ CaCl_2} \right.$

149.5g $CaCl_2$

PERCENT YIELD VOCABULARY

Ca 40
O 32
H 2

1) theoretical yield -the maximum amount of product that can be produced from a iven amount of reactant.

2) actual yield -the measured amount of product obtained from a reaction.

3) Percent yield -

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

Example 13: Methanol can be produced through the reaction of CO and H2 in the presence of a catalyst:

$$CO(g) + 2H_2(g) \xrightarrow{\text{catalyst}} CH_3OH(l)$$

If 75.0g of CO reacts to $\boxed{\text{produce 68.4g of } CH_3OH,}$ what is the percent yield of CH_3OH?

Experiment (Actual yield)

(Math aka Theoretical)

$$\frac{75g\ CO}{} \left| \frac{1mol\ CO}{28g\ CO} \right| \frac{1mol\ CH_3OH}{1mol\ CO} \left| \frac{32g\ CH_3OH}{1mol\ CH_3OH} \right. = 85.7g\ CH_3OH$$

C 12
O 16
28

THUS... $\frac{A}{T} \times 100$

$$\frac{68.4g}{85.7g} \times 100 = \boxed{79.8\%}$$

N-7

PRACTICE PROBLEMS

5.5

1) Given: $N_2 + 3H_2 \longrightarrow 2NH_3$

If 5.50g of hydrogen react with excess nitrogen to produce 20.4g of ammonia, what is the percent yield?

$$\frac{5.50\text{g } H_2}{} \left| \frac{1 \text{ mol } H_2}{2\text{g } H_2} \right| \frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2} \left| \frac{17\text{g } NH_3}{1 \text{ mol } NH_3} \right. = 31.16\text{g}$$

Theoretical or math

Actual → 20.4

Theory → 31.16

$$\frac{20.4}{31.16} \times 100 = 65.4\%$$

2) Given: $3NaBH_4 + 4BF_3 \longrightarrow 3NaBF_4 + 2B_2H_6$

If you begin with 18.9g of $NaBH_4$ and excess BF_3, and you isolate 7.50g of B_2H_6, what is the percent yield of B_2H_6?

$$\frac{18.9\text{g } NaBH_4}{} \left| \frac{1 \text{ mol } NaBH_4}{37.8\text{g } NaBH_4} \right| \frac{2 \text{ mol } B_2H_6}{3 \text{ mol } NaBH_4} \left| \frac{27.6\text{g } B_2H_6}{1 \text{ mol } B_2H_6} \right. =$$

23
10.8
4

10.8(2)
+6

21.6
+6
27.6

Theoreti
9.2g
math

Actual → 7.5

Theory → 9.2

$$\frac{7.5}{9.2} \times 100 = 81.5\%$$

3) Given: $SiO_2 + 4HF \longrightarrow SiF_4 + 2H_2O$

Calculate the percent yield if 1.28g of SiO_2 gives 1.08g of SiF_4 as product.

Theory Math

$$\frac{1.28\text{g } SiO_2}{} \left| \frac{1 \text{ mol } SiO_2}{60.1\text{g } SiO_2} \right| \frac{1 \text{ mol } SiF_4}{1 \text{ mol } SiO_2} \left| \frac{104.1\text{g } SiF_4}{1 \text{ mol } SiF_4} \right. = 2.2\text{g } SiF_4$$

28.1
32
60.1

28.1
+4(19)
104.1

Actual 1.08

Theory 2.2

$$\frac{1.08}{2.2} \times 100 = 49.1\%$$

N-8