Mass Relationships

Stoichiometry

mass relationships

how much reactant is needed to yield a certain amount of product

Amounts of Reactants and Products

The mole method

1. Write and balance the equation.
2. Convert the given quantities into moles.
3. Use the coefficients in the balanced equation to relate the number of moles of known substances to the desired unknown one.
4. Convert to desired units.
5. Check your answer.

Stoichiometry

Molar ratio of Y to X

$$\boxed{\frac{n_y}{n_x}}$$

Moles of X

Moles of Y

Mass of X

$$X \longrightarrow Y$$

Mass of Y

Example

How many grams of nitrogen dioxide can be formed by reaction of 1.44 g of nitrogen monoxide with oxygen?

$$2NO + O_2 \longrightarrow 2NO_2$$

Stoichiometry

Molar ratio of Y to X

$$\boxed{\frac{n_y}{n_x}}$$

Moles of X

Moles of Y

1.44 g of NO

Mass of NO_2

$$2NO + O_2 \longrightarrow 2NO_2$$

Stoichiometry

Molar ratio of Y to X

$$\boxed{\frac{n_y}{n_x}}$$

0.048 Moles NO

Moles of Y

1.44 g of NO

Mass of NO_2

$$2NO + O_2 \longrightarrow 2NO_2$$

Stoichiometry

Molar ratio NO_2 to NO

$\boxed{\dfrac{2}{2}}$

0.048 Moles NO

Moles of Y

1.44 g of NO

Mass of NO_2

$$2NO + O_2 \longrightarrow 2NO_2$$

Stoichiometry

Molar ratio NO_2 to NO

$$\boxed{\frac{2}{2}}$$

0.048 Moles NO

0.048 Moles NO_2

1.44 g of NO

Mass of NO_2

$$2NO + O_2 \longrightarrow 2NO_2$$

Stoichiometry

Molar ratio NO_2 to NO

$$\boxed{\frac{2}{2}}$$

0.048 Moles NO

0.048 Moles NO_2

1.44 g of NO

2.21 g of NO_2

$$2NO + O_2 \longrightarrow 2NO_2$$

Example

$$2NO + O_2 \longrightarrow 2NO_2$$

$$1.44\text{g } \cancel{NO} \times \frac{1\text{mol } \cancel{NO}}{30\text{g } \cancel{NO}} \times \frac{2\text{mol } \cancel{NO_2}}{2\text{mol } \cancel{NO}}$$

$$\times \frac{46\text{g } NO_2}{1\text{mol } \cancel{NO_2}} = 2.21\text{g } NO_2$$

Limiting Reagents

Limiting Reagent

Reactants are not always present (or available) in "stoichiometric" quantities.

One reactant may be present in quantities such that it is completely consumed while excess amounts of other reactants remain.

- called "limiting reactant" or "limiting reagent"

The limiting reagent will limit the amount of product produced.

Example

How many moles of $MgCl_2$ will be produced?

$$Mg + Cl_2 \longrightarrow MgCl_2$$

	Mg	Cl_2	$MgCl_2$
Start	1 mol	1 mol	0
Finish	0	0	1 mol

Example

How many moles of $MgCl_2$ will be produced?

$$Mg + Cl_2 \longrightarrow MgCl_2$$

	Mg	Cl_2	$MgCl_2$
Start	1 mol	2 mol	0
Finish	0	1 mol	1 mol

magnesium is the limiting reagent

1 mol of chlorine will be left unchanged

Limiting Reagent

Molar ratio Y to X

$$\boxed{\frac{n_y}{n_x}}$$

Moles of X

Moles of W

Moles of Y

Mass of X

Mass of W

W + X $\longrightarrow$ Y

Mass of Y

Compare molar ratio W to X to their coefficients in balancd equation; identify LR

Molar ratio Y to LR

Moles of X

Moles of W

Moles of Y

Mass of X

$W + X \longrightarrow Y$

Mass of W

Mass of Y

Example

Determine the limiting reagent and the amount of PI_3 produced when 6.00g P_4 reacts with 25.0g of I_2.

$$P_4 + 6I_2 \longrightarrow 4PI_3$$

$$6.00\text{g}\ \cancel{P_4} \times \frac{1\text{mol}\ P_4}{124\text{g}\ \cancel{P_4}} = .0484\text{mol}\ P_4$$

$$25.0\text{g}\ \cancel{I_2} \times \frac{1\text{mol}\ I_2}{254\text{g}\ \cancel{I_2}} = .0984\text{mol}\ I_2$$

Example cont...

Determine how much I_2 would be needed to react completely with the available amount of P_4.

$$P_4 + 6I_2 \longrightarrow 4PI_3$$

$$.0484\text{mol } \cancel{P_4} \times \frac{6\text{mol } I_2}{1\text{mol } \cancel{P_4}} = 0.290\text{mol } I_2$$

amount of I_2 needed

but only .0984mol I_2 is available

I_2 is the limiting reagent

Example cont...

…the amount of PI_3 produced from the limiting reagent...

$$.0984\text{mol } \cancel{I_2} \times \frac{4\text{mol } \cancel{PI_3}}{6\text{mol } \cancel{I_2}} \times \frac{412\text{g } PI_3}{1\text{mol } \cancel{PI_3}} =$$

$27.0g\ PI_3$

another method

Determine the limiting reagent and the amount of PI_3 produced when 6.00g P_4 reacts with 25.0g of I_2.

$$P_4 + 6I_2 \longrightarrow 4PI_3$$

$$6.00\text{g } \cancel{P_4} \times \frac{1\text{mol } P_4}{124\text{g } \cancel{P_4}} = .0484\text{mol } P_4$$

$$25.0\text{g } \cancel{I_2} \times \frac{1\text{mol } I_2}{254\text{g } \cancel{I_2}} = .0984\text{mol } I_2$$

another method

$$P_4 + 6I_2 \longrightarrow 4PI_3$$

$$\frac{.0984\text{mol } I_2}{.0484\text{mol } P_4} = 2.033$$

The actual I_2/P_4 ratio is less than the stiochiometic ratio

$$\frac{6 \text{ mol } I_2}{1 \text{ mol } P_4} = 6$$

So there is not enough I_2 to react with all the P_4

I_2 is the limiting reagent

Example

How much aluminum oxide is formed from 124 g of Al and 601 g Fe_2O_3?

$$2Al + Fe_2O_3 \longrightarrow Al_2O_3 + 2Fe$$

124 g   601 g

$\boxed{4.6 \text{ mol}}$ 3.8 mol

excess of 1.5 mol

limiting reactant

Example

How much aluminum oxide is formed from 124 g of Al and 601 g Fe_2O_3?

$$2Al + Fe_2O_3 \longrightarrow Al_2O_3 + 2Fe$$

124 g

2.3 mol

= 234 g

$\boxed{4.6 \text{ mol}}$

limiting reactant

Example

From the reaction between of 10.0g of Hg and 9.0g of Br_2 . What mass of which reagent is left unreacted?

$$Hg + Br_2 \longrightarrow HgBr_2$$

$$10.0\text{g Hg} \times \frac{1\ \text{molHg}}{200.6\text{gHg}} = 4.99 \times 10^{-2}\ \text{molHg}$$

$$9.0\text{g } Br_2 \times \frac{1\ \text{mol}Br_2}{159.8\text{g}Br_2} = 5.63 \times 10^{-2}\ \text{mol}Br_2$$

$$Hg + Br_2 \longrightarrow HgBr_2$$

Hg is limiting

$$4.99 \times 10^{-2}\ \text{molHg} \times \frac{1\ \text{molBr}_2}{1\ \text{molHg}} = 4.99 \times 10^{-2}\ \text{molBr}_2$$

moles of Br_2 needed to use up Hg available

$$4.99 \times 10^{-2}\ \text{molBr}_2 \times \frac{159.8\text{gBr}_2}{1\ \text{molBr}_2} = 7.97\ \text{g Br}_2$$

grams of Br_2 used

$$9.0\text{g Br}_2 - 7.97\ \text{g Br}_2 = \boxed{1.03\ \text{g Br}_2\ \text{excess}}$$

Reaction Yield

Theoretical yield

the amount of product that would result if all the limiting reagent reacted

Actual yield

the amount of product actually obtained from the reaction

Almost always less than the theoretical yield

Percent Yield

$$\%\text{yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

Determines how efficient a reaction is

Example

In a certain industrial operation 3.54×10^7g of $TiCl_4$ is reacted with 1.13×10^7g of Mg. (a) Calculate the theoretical yield of Ti in grams. (b) Calculate the percent yield if 7.91×10^6g are actually obtained.

$$TiCl_4 + 2Mg \longrightarrow Ti + 2MgCl_2$$

Calculate theoretical yield

$$3.54 \times 10^7 \cancel{g\ TiCl_4} \times \frac{1\text{mol } TiCl_4}{187.7\cancel{g\ TiCl_4}} = 1.87 \times 10^5 \text{mol } TiCl_4$$

$$25.0\cancel{g\ Mg} \times \frac{1\text{mol Mg}}{24.31\cancel{g\ Mg}} = 4.65 \times 10^5 \text{ mol Mg}$$

$$1.87 \times 10^5 \cancel{\text{mol } TiCl_4} \times \frac{2\text{mol Mg}}{1\cancel{\text{mol } TiCl_4}} = 3.74 \times 10^5 \text{ mol Mg}$$

there is more than enough Mg

$TiCl_4$ is limiting

$$3.54 \times 10^7 \text{g } \cancel{\text{TiCl}_4} \times \frac{1\cancel{\text{mol TiCl}_4}}{\cancel{187.7\text{g TiCl}_4}} \times \frac{1\cancel{\text{mol}}\text{ Ti}}{\cancel{1\text{mol TiCl}_4}}$$

$$\times \frac{47.88\text{g Ti}}{\cancel{1\text{mol Ti}}} = 8.93 \times 10^6\text{g Ti}$$

$$\%\text{yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

$$100\% \times \frac{7.91 \times 10^6\text{g Ti}}{8.93 \times 10^6\text{g Ti}} =$$

$$= 88.6\%$$