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Chapter 3

Mass Relationships in Chemical Reactions

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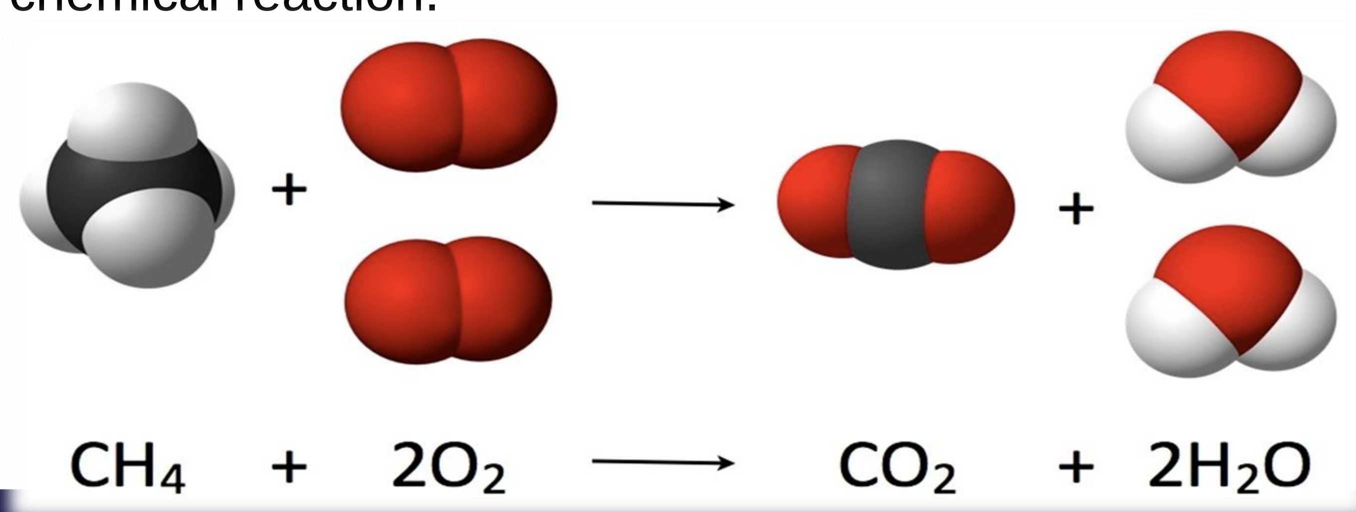
3.7
Chemical Reactions and Chemical Equations

A **chemical reaction**: a process in which a substance (or substances) is changed into one or more new substances.

A **chemical equation**: uses chemical symbols to show what happens during a chemical reaction.



Writing Chemical Equations

What happens when hydrogen gas (H_2) burns in air (which contains oxygen, O_2) to form water (H_2O). This reaction can be represented by the chemical equation

$$H_2 + O_2 \longrightarrow H_2O$$

$+$ means "reacts with"

$\longrightarrow$ means "to yield"

This symbolic expression can be read: Molecular hydrogen reacts with molecular oxygen to yield water. The reaction is assumed to proceed from left to right as the arrow indicates.

To conform with the law of conservation of mass, there must be the same number of each type of atom on both sides of the arrow (we should balance the equation).

$$2H_2 + O_2 \longrightarrow 2H_2O$$ (balanced equation)

$2H_2$	$+ O_2$	$\longrightarrow$	$2H_2O$
Two molecules	+ one molecule	$\longrightarrow$	two molecules
2 moles	+ 1 mole	$\longrightarrow$	2 moles
$2(2.02\ g) = 4.04\ g$	$+\ 32.00\ g$	$\longrightarrow$	$2(18.02\ g) = 36.04\ g$

$$\underbrace{2(2.02\ g) = 4.04\ g + 32.00\ g}_{36.04\ g\ \text{reactants}} \longrightarrow \underbrace{2(18.02\ g) = 36.04\ g}_{36.04\ g\ \text{product}}$$

$$2H_2 + O_2 \longrightarrow 2H_2O$$

In the equation:

- H_2 and O_2 are the **reactants**, which are the starting materials in a chemical reaction.
- Water is the **product**, which is the substance formed as a result of a chemical reaction.

$$\text{reactants} \longrightarrow \text{products}$$

The **states** of the reactants and products are written in parentheses to the right of each compound; (*g*) gas, (*l*) liquid, (*s*) solid or (*aq*) aqueous solution.

Examples:

$$2CO(g) + O_2(g) \longrightarrow 2CO_2(g)$$

$$2HgO(s) \longrightarrow 2Hg(l) + O_2(g)$$

$$NaCl(s) \xrightarrow{H_2O} NaCl(aq)$$

$$CH_{4\,(g)} + 2\,O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2\,H_2O_{(g)}$$

Reactants appear on the left side of the equation.

Products appear on the right side of the equation.

The **states** of the reactants and products are written in parentheses to the right of each compound; (*g*) gas, (*l*) liquid, (*s*) solid, (*aq*) aqueous solution.

Subscripts present within a formula and tell the number of atoms of each element in a molecule.

Coefficients are inserted in front of a formula to balance the equation.

Sometimes the conditions (such as temperature or pressure) under which the reaction proceeds appear above or below the reaction arrow. Δ refer to temperature.

Balancing Chemical Equations

In the laboratory, small amounts of oxygen gas can be prepared by heating potassium chlorate ($KClO_3$). The products are oxygen gas (O_2) and potassium chloride (KCl). From this information, we write

$$KClO_3 \longrightarrow KCl + O_2$$

$$2KClO_3 \longrightarrow KCl + 3O_2$$

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$

Heating potassium chlorate produces oxygen, which supports the combustion of wood splint.

Check the number of each element

Reactants	Products
K (2)	K (2)
Cl (2)	Cl (2)
O (6)	O (6)

EXAMPLE

The combustion (that is, burning) of the natural gas component ethane (C_2H_6) in oxygen or air, which yields carbon dioxide (CO_2) and water.

$$C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$$

$$C_2H_6 + O_2 \longrightarrow 2CO_2 + H_2O$$

$$C_2H_6 + O_2 \longrightarrow 2CO_2 + 3H_2O$$

$$C_2H_6 + \tfrac{7}{2}O_2 \longrightarrow 2CO_2 + 3H_2O$$

However, we normally prefer to express the coefficients as whole numbers rather than as fractions.

$$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$$

EXAMPLE

When aluminum metal is exposed to air, a protective layer of aluminum oxide (Al_2O_3) forms on its surface.

$$Al + O_2 \longrightarrow Al_2O_3$$

$$2Al + O_2 \longrightarrow Al_2O_3$$

$$2Al + \frac{3}{2}O_2 \longrightarrow Al_2O_3$$

However, equations are normally balanced with the smallest set of whole number coefficients.

$$4Al + 3O_2 \longrightarrow 2Al_2O_3$$

3.8
Amounts of Reactants and Products

Stoichiometry

How much product will be formed from specific amounts of starting materials (reactants)?

How much starting material must be used to obtain a specific amount of product?

Stoichiometry is the quantitative study of reactants and products in a chemical reaction.

To interpret a reaction quantitatively, we need to apply our knowledge of molar masses and the mole concept.

Mole method, which means that the stoichiometric coefficients in a chemical equation can be interpreted as the number of moles of each substance.

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

The stoichiometric coefficients show that one molecule of N_2 reacts with three molecules of H_2 to form two molecules of NH_3. It follows that the relative numbers of moles are the same as the relative number of molecules:

$N_2(g)$	+	$3H_2(g)$	$\longrightarrow$	$2NH_3(g)$
1 molecule		3 molecules		2 molecules
6.022×10^{23} molecules		$3(6.022 \times 10^{23}$ molecules)		$2(6.022 \times 10^{23}$ molecules)
1 mol		3 mol		2 mol

This equation can also be read as "1 mole of N_2 gas combines with 3 moles of H_2 gas to form 2 moles of NH_3 gas".

In stoichiometric calculations,

$$3 \text{ mol } H_2 \simeq 2 \text{ mol } NH_3$$

$$1 \text{ mol } N_2 \simeq 2 \text{ mol } NH_3$$

$$1 \text{ mol } N_2 \simeq 3 \text{ mol } H_2$$

the symbol $\simeq$ means "stoichiometrically equivalent to".

This relationship enables us to write the conversion factors

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

$$\boxed{N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)}$$

e.g., 6.0 moles of H_2 react completely with N_2 to form NH_3. Calculate the amount of NH_3 produced in moles?

$$\begin{aligned}\text{moles of } NH_3 \text{ produced} &= 6.0\ \cancel{\text{mol } H_2} \times \frac{2 \text{ mol } NH_3}{3\ \cancel{\text{mol } H_2}} \\ &= 4.0 \text{ mol } NH_3\end{aligned}$$

e.g., suppose 16.0 g of H_2 react completely with N_2 to form NH_3. How many grams of NH_3 will be formed?

the link between H_2 and NH_3 is the mole ratio from the balanced equation. So we need to first convert grams of H_2 to moles of H_2, then to moles of NH_3, and finally to grams of NH_3.

$$\begin{aligned}\text{moles of } H_2 &= 16.0\ \cancel{\text{g } H_2} \times \frac{1 \text{ mol } H_2}{2.016\ \cancel{\text{g } H_2}} \\ &= 7.94 \text{ mol } H_2\end{aligned}$$

/

$$\begin{aligned}\text{moles of } NH_3 &= 7.94\ \cancel{\text{mol } H_2} \times \frac{2 \text{ mol } NH_3}{3\ \cancel{\text{mol } H_2}} \\ &= 5.29 \text{ mol } NH_3\end{aligned}$$

/

$$\begin{aligned}\text{grams of } NH_3 &= 5.29\ \cancel{\text{mol } NH_3} \times \frac{17.03 \text{ g } NH_3}{1\ \cancel{\text{mol } NH_3}} \\ &= 90.1 \text{ g } NH_3\end{aligned}$$

$$\begin{aligned}\text{grams of } NH_3 &= 16.0\ \cancel{\text{g } H_2} \times \frac{1\ \cancel{\text{mol } H_2}}{2.016\ \cancel{\text{g } H_2}} \times \frac{2\ \cancel{\text{mol } NH_3}}{3\ \cancel{\text{mol } H_2}} \times \frac{17.03 \text{ g } NH_3}{1\ \cancel{\text{mol } NH_3}} \\ &= 90.1 \text{ g } NH_3\end{aligned}$$

Similarly, we can calculate the mass in grams of N_2 consumed in this reaction.

$$\begin{aligned}\text{grams of } N_2 &= 16.0\ \cancel{\text{g } H_2} \times \frac{1\ \cancel{\text{mol } H_2}}{2.016\ \cancel{\text{g } H_2}} \times \frac{1\ \cancel{\text{mol } N_2}}{3\ \cancel{\text{mol } H_2}} \times \frac{28.02 \text{ g } N_2}{1\ \cancel{\text{mol } N_2}} \\ &= 74.1 \text{ g } N_2\end{aligned}$$

EXAMPLE

The food we eat is degraded, or broken down, in our bodies to provide energy for growth and function. A general overall equation for this very complex process represents the degradation of glucose ($C_6H_{12}O_6$) to carbon dioxide (CO_2) and water (H_2O):

$$C_6H_{12}O_6 + 6O_2 \longrightarrow 6CO_2 + 6H_2O$$

If 856 g of $C_6H_{12}O_6$ is consumed by a person over a certain period, what is the mass of CO_2 produced?

$$\text{grams of } C_6H_{12}O_6 \longrightarrow \text{moles of } C_6H_{12}O_6 \longrightarrow \text{moles of } CO_2 \longrightarrow \text{grams of } CO_2$$

$$856\ \cancel{\text{g } C_6H_{12}O_6} \times \frac{1 \text{ mol } C_6H_{12}O_6}{180.2\ \cancel{\text{g } C_6H_{12}O_6}} = 4.750 \text{ mol } C_6H_{12}O_6$$

$$4.750\ \cancel{\text{mol } C_6H_{12}O_6} \times \frac{6 \text{ mol } CO_2}{1\ \cancel{\text{mol } C_6H_{12}O_6}} = 28.50 \text{ mol } CO_2$$

$$28.50\ \cancel{\text{mol } CO_2} \times \frac{44.01 \text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}} = 1.25 \times 10^3 \text{ g } CO_2$$

$$\text{mass of } CO_2 = 856\ \cancel{\text{g } C_6H_{12}O_6} \times \frac{1\ \cancel{\text{mol } C_6H_{12}O_6}}{180.2\ \cancel{\text{g } C_6H_{12}O_6}} \times \frac{6\ \cancel{\text{mol } CO_2}}{1\ \cancel{\text{mol } C_6H_{12}O_6}} \times \frac{44.01 \text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}}$$
$$= \boxed{1.25 \times 10^3 \text{ g } CO_2}$$

EXAMPLE

All alkali metals react with water to produce hydrogen gas and the corresponding alkali metal hydroxide. A typical reaction is that between lithium and water:

$$2Li(s) + 2H_2O(l) \longrightarrow 2LiOH(aq) + H_2(g)$$

How many grams of Li are needed to produce 9.89 g of H_2?

$$\text{grams of } H_2 \longrightarrow \text{moles of } H_2 \longrightarrow \text{moles of Li} \longrightarrow \text{grams of Li}$$

$$9.89\ \cancel{g\ H_2} \times \frac{1\ \cancel{mol\ H_2}}{2.016\ \cancel{g\ H_2}} \times \frac{2\ \cancel{mol\ Li}}{1\ \cancel{mol\ H_2}} \times \frac{6.941\ g\ Li}{1\ \cancel{mol\ Li}} = \boxed{68.1\ g\ Li}$$

Practice Exercise

Methanol (CH_3OH) burns in air according to the equation

$$2CH_3OH + 3O_2 \longrightarrow 2CO_2 + 4H_2O$$

If 209 g of methanol are used up in a combustion process, what is the mass of H_2O produced?

Practice Exercise

The reaction between nitric oxide (NO) and oxygen to form nitrogen dioxide (NO_2) is a key step in photochemical smog formation:

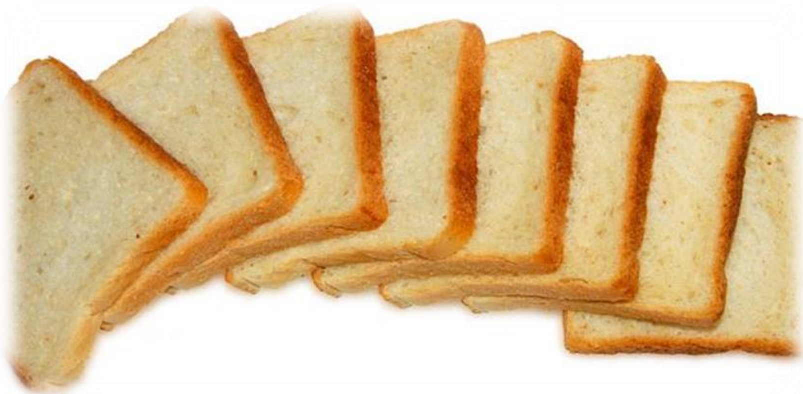
$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$

How many grams of O_2 are needed to produce 2.21 g of NO_2?

3.9
Limiting Reagents

Limiting Reagents

How Many Cheese Sandwiches Can I Make?



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The amount of available bread limits the number of sandwiches.

Limiting reagent **(limiting reactant):** the reactant used up first in a reaction, because the maximum amount of product formed depends on how much of this reactant was originally present. When this reactant is used up, no more product can be formed.

Limiting reagent is completely consumed in a reaction (present in the smallest stoichiometric amount). Its called **limiting regent**; because it determines or limits the amount of product formed.

Excess reagents are the reactants present in quantities greater than necessary to react with the quantity of the limiting reagent.

EXAMPLE

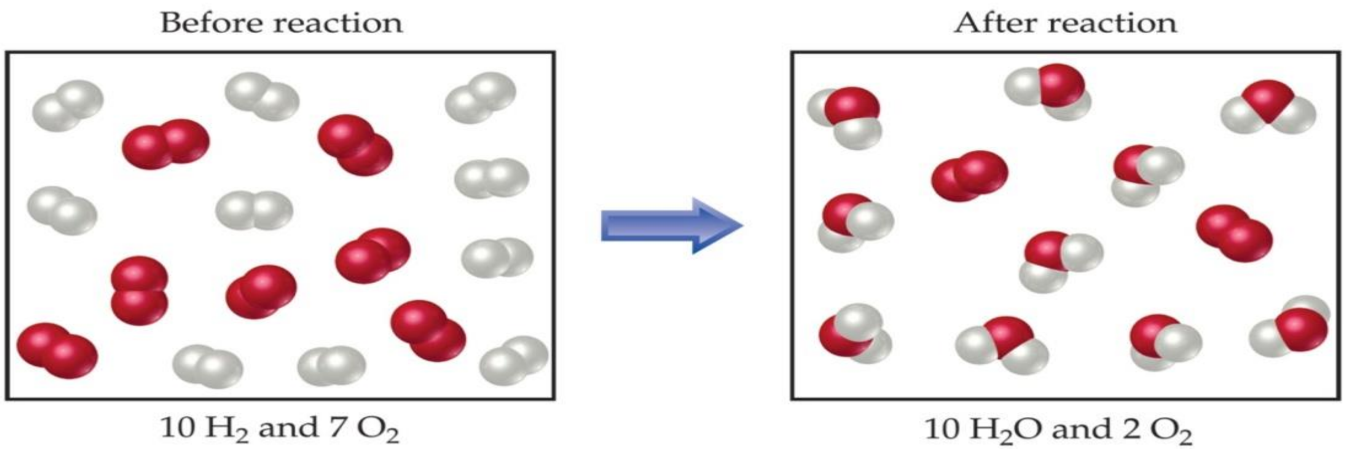
$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$$

$$2 \text{ mol } H_2 \simeq 1 \text{ mol } O_2$$

Suppose, mixture of 10 mole H_2 and 7 mole O_2 react to form water.

The number of O_2 needed to react with all the H_2 is:

$$\text{Moles } O_2 = (10 \cancel{\text{ mol } H_2})\left(\frac{1 \text{ mol } O_2}{2 \cancel{\text{ mol } H_2}}\right) = 5 \text{ mol } O_2$$



In this example, H_2 would be the **limiting reactant**, which means that once all the H_2 has been consumed the reaction stops. And O_2 would be the **excess reactant**, and some is left over when the reaction stops.

EXAMPLE

Consider the industrial synthesis of methanol (CH_3OH) from carbon monoxide and hydrogen at high temperatures:

$$CO(g) + 2H_2(g) \longrightarrow CH_3OH(g)$$

Suppose initially we have 4 moles of CO and 6 moles of H_2.

In stoichiometric calculations involving limiting reagents, the first step is to decide which reactant is the limiting reagent.

One way to determine the limiting reagent is to calculate the number of moles of CH_3OH obtained based on the initial quantities of CO and H_2; the limiting reagent will yield the **smaller** amount of the product.

$$4\ \cancel{\text{mol CO}} \times \frac{1\ \text{mol CH}_3\text{OH}}{1\ \cancel{\text{mol CO}}} = 4\ \text{mol CH}_3\text{OH}$$

$$6\ \cancel{\text{mol H}_2} \times \frac{1\ \text{mol CH}_3\text{OH}}{2\ \cancel{\text{mol H}_2}} = 3\ \text{mol CH}_3\text{OH}$$

Because H_2 results in a smaller amount of CH_3OH, it must be the limiting reagent. Therefore, CO is the excess reagent.

EXAMPLE

Urea [$(NH_2)_2CO$] is prepared by reacting ammonia with carbon dioxide:

$$2NH_3(g) + CO_2(g) \longrightarrow (NH_2)_2CO(aq) + H_2O(l)$$

In one process, 637.2 g of NH_3 are treated with 1142 g of CO_2.

(a) Which of the two reactants is the limiting reagent?

(b) Calculate the mass of $(NH_2)_2CO$ formed.

(c) How much excess reagent (in grams) is left at the end of the reaction?

(a)

from NH_3

$$\text{moles of } (NH_2)_2CO = 637.2\ \cancel{\text{g } NH_3} \times \frac{1\ \cancel{\text{mol } NH_3}}{17.03\ \cancel{\text{g } NH_3}} \times \frac{1 \text{ mol } (NH_2)_2CO}{2\ \cancel{\text{mol } NH_3}}$$
$$= 18.71 \text{ mol } (NH_2)_2CO$$

from CO_2

$$\text{moles of } (NH_2)_2CO = 1142\ \cancel{\text{g } CO_2} \times \frac{1\ \cancel{\text{mol } CO_2}}{44.01\ \cancel{\text{g } CO_2}} \times \frac{1 \text{ mol } (NH_2)_2CO}{1\ \cancel{\text{mol } CO_2}}$$
$$= 25.95 \text{ mol } (NH_2)_2CO$$

NH_3 is the limiting reagent because it produces a smaller amount of $(NH_2)_2CO$.

(b) We determined the moles of $(NH_2)_2CO$ produced in part (a), using NH_3 as the limiting reagent. The molar mass of $(NH_2)_2CO$ is 60.06 g.

$$\text{mass of } (NH_2)_2CO = 18.71 \cancel{\text{mol } (NH_2)_2CO} \times \frac{60.06 \text{ g } (NH_2)_2CO}{1 \cancel{\text{mol } (NH_2)_2CO}}$$
$$= \boxed{1124 \text{ g } (NH_2)_2CO}$$

(c) We can determine the amount of CO_2 that reacted to produce 18.71 moles of $(NH_2)_2CO$. The amount of CO_2 left over is the difference between the initial amount and the amount reacted.

Starting with 18.71 moles of $(NH_2)_2CO$, we can determine the mass of CO_2 that reacted

$$\text{mass of } CO_2 \text{ reacted} = 18.71 \cancel{\text{mol } (NH_2)_2CO} \times \frac{1 \cancel{\text{mol } CO_2}}{1 \cancel{\text{mol } (NH_2)_2CO}} \times \frac{44.01 \text{ g } CO_2}{1 \cancel{\text{mol } CO_2}}$$
$$= 823.4 \text{ g } CO_2$$

The amount of CO_2 remaining (in excess) is the difference between the initial amount and the amount reacted:

$$\text{mass of } CO_2 \text{ remaining} = 1142 \text{ g} - 823.4 \text{ g} = \boxed{319 \text{ g}}$$

Practice Exercise

The reaction between aluminum and iron(III) oxide can generate temperatures approaching 3000ºC and is used in welding metals:

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

In one process, 124 g of Al are reacted with 601 g of Fe_2O_3.

(a) Calculate the mass (in grams) of Al_2O_3 formed.

(b) How much of the excess reagent is left at the end of the reaction?

3.10
Reaction Yield

Theoretical yield of the reaction, is the amount of product that would result if all the limiting reagent reacted. The theoretical yield, then, is the maximum obtainable yield, predicted by the balanced equation.

Actual yield is the amount of product actually obtained (in practice) from a reaction, is almost always less than the theoretical yield.

To determine how efficient a given reaction is, chemists often figure the **percent yield**, which describes the proportion of the actual yield to the theoretical yield. It is calculated as follows:

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

Percent yields may range from a fraction of 1 percent to 100 percent.

- The theoretical yield is the yield that you calculate using the balanced equation.
- The actual yield is the yield obtained by carrying out the reaction.

EXAMPLE

Titanium is a strong, lightweight, corrosion-resistant metal that is used in rockets, aircraft, jet engines, and bicycle frames. It is prepared by the reaction of titanium(IV) chloride with molten magnesium between 950°C and 1150°C:

$$TiCl_4(g) + 2Mg(l) \longrightarrow Ti(s) + 2MgCl_2(l)$$

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

Because there are two reactants, this is likely to be a **limiting reagent** problem.

(a) from $TiCl_4$

$$\begin{aligned}\text{moles of Ti} &= 3.54 \times 10^7 \cancel{\text{g TiCl}_4} \times \frac{1 \cancel{\text{mol TiCl}_4}}{189.7 \cancel{\text{g TiCl}_4}} \times \frac{1 \text{ mol Ti}}{1 \cancel{\text{mol TiCl}_4}} \\ &= 1.87 \times 10^5 \text{ mol Ti}\end{aligned}$$

from Mg

$$\begin{aligned}\text{moles of Ti} &= 1.13 \times 10^7 \cancel{\text{g Mg}} \times \frac{1 \cancel{\text{mol Mg}}}{24.31 \cancel{\text{g Mg}}} \times \frac{1 \text{ mol Ti}}{2 \cancel{\text{mol Mg}}} \\ &= 2.32 \times 10^5 \text{ mol Ti}\end{aligned}$$

Therefore, $TiCl_4$ is the limiting reagent because it produces a smaller amount of Ti. The mass of Ti formed is

$$1.87 \times 10^5 \cancel{\text{mol Ti}} \times \frac{47.88 \text{ g Ti}}{1 \cancel{\text{mol Ti}}} = \boxed{8.95 \times 10^6 \text{ g Ti}}$$

(b) The mass of Ti determined in part (a) is the theoretical yield. The amount given in part (b) is the actual yield of the reaction.

$$\begin{aligned}\%\text{yield} &= \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \\ &= \frac{7.91 \times 10^6 \text{ g}}{8.95 \times 10^6 \text{ g}} \times 100\% \\ &= \boxed{88.4\%}\end{aligned}$$

Practice Exercise

Industrially, vanadium metal, which is used in steel alloys, can be obtained by reacting vanadium(V) oxide with calcium at high temperatures:

$$5Ca + V_2O_5 \longrightarrow 5CaO + 2V$$

In one process, 1.54 x 10^3 g of V_2O_5 react with 1.96 x 10^3 g of Ca.
(a) Calculate the theoretical yield of V.
(b) Calculate the percent yield if 803 g of V are obtained.


Thank You!