GENERAL CHEMISTRY

Principles and Modern Applications

TENTH EDITION

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Chapter 4:

Chemical Reactions

Chemical Compounds

CONTENTS

4-1 Chemical Reactions and Chemical Equations

As reactants are converted to products we observe:

- Color change
- Precipitate formation
- Gas evolution
- Heat absorption or evolution

Chemical evidence may be necessary

▲ FIGURE 4-1
Precipitation of silver chromate

When aqueous solutions of silver nitrate and potassium chromate are mixed,

$AgNO_3$ + $KCrO_4$ → $AgCrO_4$
yellow red

the disappearance of the distinctive yellow color of chromate ion and the appearance of the red-brown solid, silver chromate, provide physical evidence of a reaction.

Figure 4-2 Evidence of a chemical reaction

(a) (b)

(a) Evolution of a gas: When a copper penny reacts with nitric acid, the red-brown gas is evolved. (b) Evolution of heat: When iron gauze (steel wool) is ignited in an oxygen atmosphere, evolved heat and light provide physical evidence of a reaction.

Chemical Equation

-We have a symbolic or shorthand way of representing a chemical rxn — *chemical equation*—

- Formulas for the reactants are written on the left hand side of the rxn, whereas formulas for the products are written on the right. The two sides of the equation are joined by an arrow $\rightarrow$ or = sign.

- Sometimes products react to form reactants, *reversible rxn's,* shown by double arrow , $\rightleftarrows$

Chemical Equation

In order to show the shorthand representation of the rxn,

Nitrogen monoxide + oxygen → nitrogen dioxide

Step 1: **Write** the reaction using chemical symbols.

Step 2: Balance the chemical equation.

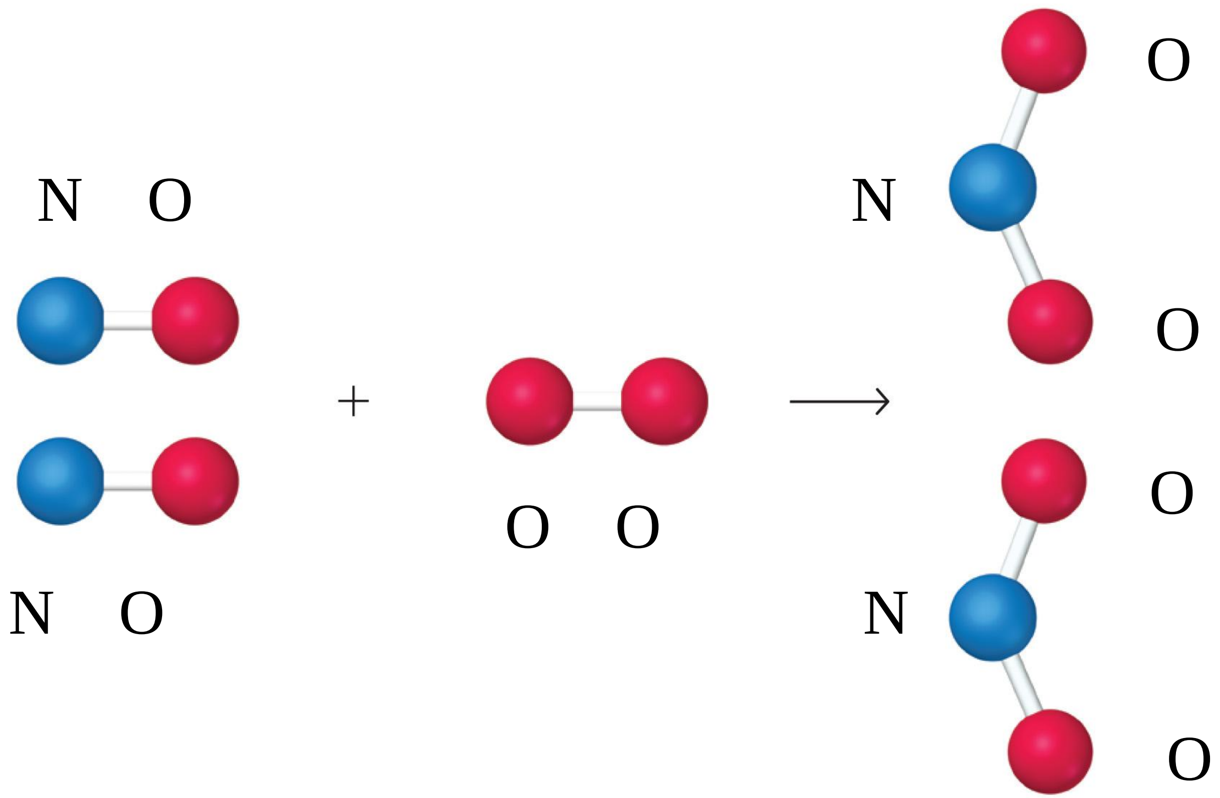
$$2\ NO + 1\ O_2 \rightarrow 2\ NO_2$$

colorless red-brown

2 molecules of NO are consumed for every 1 molecule of O_2 and 2 molecules of NO_2 are produced.

In a balanced eqn. for each element present, the total # of atoms are the same on both sides.

Molecular Representation



Balancing Equations

- An equation can be balanced only by adjusting the coefficients of formulas.

-The coefficients required to balance a chemical equation are called *stoichiometric coefficients*.

- These coefficients are essential in relating the amounts of reactants used and products formed in a chemical rxn.

In balancing a chemical rxn keep the followings in mind.

Never introduce extraneous atoms to balance.

$$NO + O_2 \rightarrow NO_2 + O$$

Never change a formula for the purpose of balancing an equation

$$NO + O_2 \rightarrow NO_3$$

Balancing Equation Strategy

Balancing by inspection. It means to adjust stoichiometric coefficients by trial and error until a balanced condition is found.

Strategies for balancing.

* Balance elements that occur in only one compound on each side first.

* Balance free elements last.

* Balance unchanged polyatomics (or other groups of atoms) as groups.

* Fractional coefficients are acceptable and can be cleared at the end by multiplication.

EXAMPLE 4-2

Writing and Balancing an Equation: The Combustion of a Carbon-Hydrogen-Oxygen Compound. Liquid triethylene glycol, C6H14O4, is used a a solvent and plasticizer for vinyl and polyurethane plastics. Write a balanced chemical equation for its complete combustion.

Chemical Equation:

$$C_6H_{14}O_4 + O_2 \rightarrow CO_2 + H_2O$$

Triethylene glycol

1. *Balance C first,* $C_6H_{14}O_4 + O_2 \rightarrow 6CO_2 + H_2O$
2. *Balance H.* $C_6H_{14}O_4 + O_2 \rightarrow 6CO_2 + 7H_2O$
3. *Balance O.* $C_6H_{14}O_4 + 15/2\ O_2 \rightarrow 6CO_2 + 7H_2O$
4. *Multiply by two*

$$2\ C_6H_{14}O_4 + 15\ O_2 \rightarrow 12\ CO_2 + 14\ H_2O$$

and check all elements.

States of Matter & Reaction conditions:

(g) Gas (l) liquid (s) solid *aqueous (aq) water*

Thus, the equation for combustion of triethylene glycol can be written as

$$2\ C_6H_{14}O_4(l) + 15\ O_2(g) \longrightarrow 12\ CO_2(g) + 14\ H_2O(l)$$

Another commonly used symbol for reactants or products dissolved in water is

(aq) *aqueous solution*

Temperature, Pressure, Catalyst

We often write rxn conditions above or below the arrow. Δ, delta means high temp is required.

$$2\ Ag_2O\ (s) \rightarrow 4\ Ag(s)\ +\ O_2\ (g)$$

Catalyst is a substance that enters into a rxn to speed up the rxn without being consumed in the rxn.

$$CO\ (g) + 2H_2\ (g) \xrightarrow[\text{340 atm}\ \ ZnO,\ Cr_2O_3]{350°C} CH_3OH\ (g)$$

4-2 Chemical Equations and Stoichiometry

Stoichiometry includes all the *quantitative* relationships involving

atomic and formula masses

chemical formulas.

chemical equations

Mole ratio or stoichiometric factor is a central conversion factor.

KEEP IN MIND

that it is important to include units and to work from a balanced chemical equation when solving stoichiometry problems.

EXAMPLE 4-3 Relating the Numbers of Moles of Reactant and Product

How many moles of CO_2 are produced in the combustion of 2.72 mol of triethylene glycol, $C_6H_{14}O_4$, in an excess of O_2?

Analyze

"An excess of O_2" means that there is more than enough O_2 available to permit the complete conversion of the triethylene glycol to CO_2 and H_2O. The factor for converting from moles of $C_6H_{14}O_4$ to moles of CO_2 is obtained from the balanced equation for the combustion reaction.

Solve

The first step in a stoichiometric calculation is to write a balanced equation for the reaction. The balanced chemical equation for the reaction is given below.

$$2\,C_6H_{14}O_4 + 15\,O_2 \longrightarrow 12\,CO_2 + 14\,H_2O$$

Thus, 12 mol CO_2 are produced for every 2 mol $C_6H_{14}O_4$ burned. The production of 12 mol CO_2 is equivalent to the consumption of 2 mol $C_6H_{14}O_4$; thus, the ratio 12 mol CO_2/2 mol $C_6H_{14}O_4$ converts from mol $C_6H_{14}O_4$ to mol CO_2.

$$?\ \text{mol } CO_2 = 2.72\ \text{mol } C_6H_{14}O_4 \times \frac{12\ \text{mol } CO_2}{2\ \text{mol } C_6H_{14}O_4} = 16.3\ \text{mol } CO_2$$

Assess

The expression above can be written in terms of two equal ratios:

$$\frac{?\ \text{mol } CO_2}{2.72\ \text{mol } C_6H_{14}O_4} = \frac{12\ \text{mol } CO_2}{2\ \text{mol } C_6H_{14}O_4}$$

You may find it easier to set up an expression in terms of ratios and then solve it for the unknown quantity.

PRACTICE EXAMPLE A: How many moles of O_2 are produced from the decomposition of 1.76 moles of potassium chlorate?

$$2\,KClO_3(s) \longrightarrow 2\,KCl(s) + 3\,O_2(g)$$

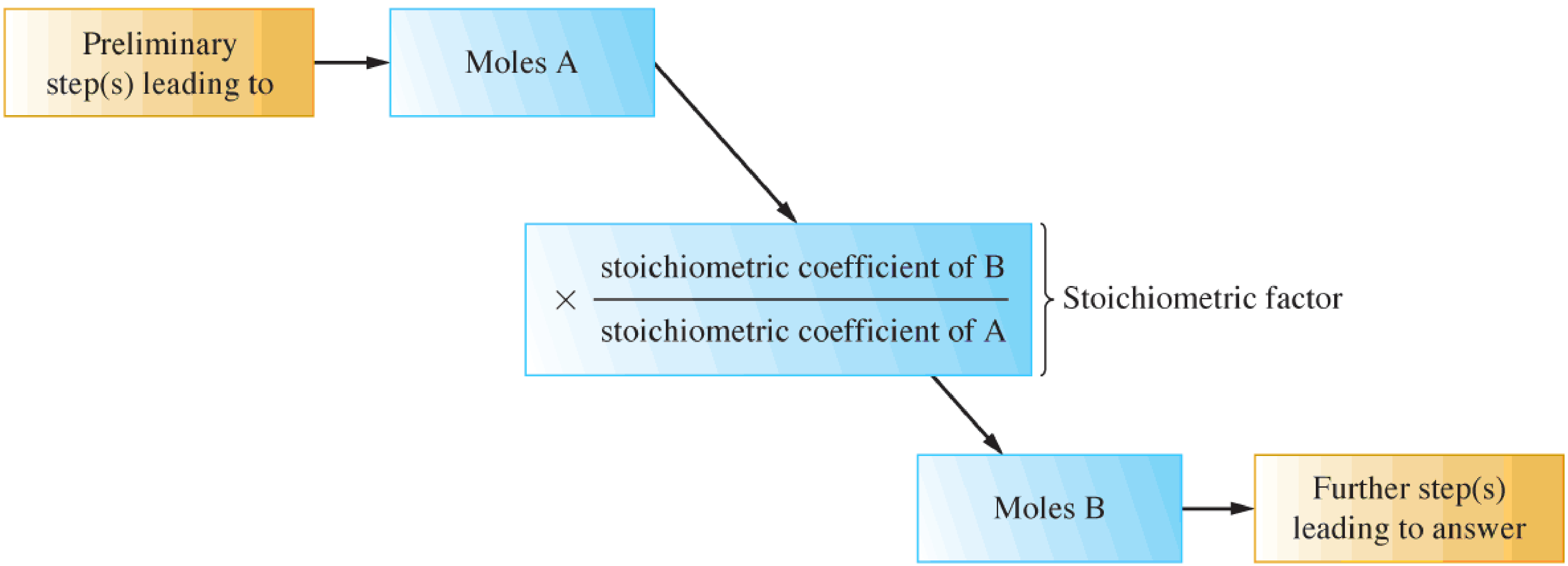


Figure 4-3 A generalized stoichiometry diagram

A key step in working stoichiometric problems is applying the appropriate stoichiometric factor (mole ratio) that converts from moles A to moles B. The stoichiometric factor is the stoichiometric coefficient of B divided by the stoichiometric coefficient of A.

EXAMPLE 4-4 Relating the Mass of a Reactant and a Product

What mass of CO_2 is formed in the reaction of 4.16 g triethylene glycol, $C_6H_{14}O_4$, with an excess of O_2?

Analyze

The general strategy involves the following conversions: (1) to moles, (2) between moles, and (3) from moles. In this example, the required conversions are g $C_6H_{14}O_4 \xrightarrow{1}$ mol $C_6H_{14}O_4 \xrightarrow{2}$ mol $CO_2 \xrightarrow{3}$ g CO_2. Each numbered arrow refers to a conversion factor that changes the unit on the left to the one on the right.

Solve

The conversions can be carried out by using either a stepwise approach or the conversion pathway approach. Using a *stepwise approach*, we proceed as follows.

Convert from grams of $C_6H_{14}O_4$ to moles of $C_6H_{14}O_4$ by using the molar mass of $C_6H_{14}O_4$ as a conversion factor.

$$? \text{ mol } C_6H_{14}O_4 = 4.16 \text{ g } C_6H_{14}O_4 \times \frac{1 \text{ mol } C_6H_{14}O_4}{150.2 \text{ g } C_6H_{14}O_4}$$
$$= 0.0277 \text{ mol } C_6H_{14}O_4$$

Convert from moles of $C_6H_{14}O_4$ to moles of CO_2 by using the stoichiometric factor.

$$? \text{ mol } CO_2 = 0.0277 \text{ mol } C_6H_{14}O_4 \times \frac{12 \text{ mol } CO_2}{2 \text{ mol } C_6H_{14}O_4}$$
$$= 0.166 \text{ mol } CO_2$$

Convert from moles of CO_2 to grams of CO_2 by using the molar mass of CO_2 as a conversion factor.

$$? \text{ g } CO_2 = 0.166 \text{ mol } CO_2 \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2}$$
$$= 7.31 \text{ g } CO_2$$

In the *conversion pathway* approach, the individual steps are combined into a single line calculation, as shown below.

$$? \text{ g } CO_2 = 4.16 \text{ g } C_6H_{14}O_4 \times \underbrace{\frac{1 \text{ mol } C_6H_{14}O_4}{150.2 \text{ g } C_6H_{14}O_4}}_{\text{converts to moles of } C_6H_{14}O_4} \times \underbrace{\frac{12 \text{ mol } CO_2}{2 \text{ mol } C_6H_{14}O_4}}_{\text{converts to moles of } CO_2} \times \underbrace{\frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2}}_{\text{converts to grams of } CO_2} = 7.31 \text{ g } CO_2$$

Note the connection between the stepwise approach and the conversion pathway approach for this problem. In the conversion pathway approach, the first conversion factor converts from grams of $C_6H_{14}O_4$ to moles of $C_6H_{14}O_4$. The second conversion factor is the stoichiometric factor, and it converts from moles of $C_6H_{14}O_4$ to moles of CO_2. The third conversion factor converts from moles of CO_2 to grams of CO_2. The conversion path-

EXAMPLE 4-5 Relating the Masses of Two Reactants to Each Other

What mass of O_2 is consumed in the complete combustion of 6.86 g of triethylene glycol, $C_6H_{14}O_4$?

Analyze

The required conversions are $\text{g } C_6H_{14}O_4 \xrightarrow{1} \text{mol } C_6H_{14}O_4 \xrightarrow{2} \text{mol } O_2 \xrightarrow{3} \text{g } O_2$.

Solve

We will first use a stepwise approach to solve this problem.

Convert from grams of $C_6H_{14}O_4$ to moles of $C_6H_{14}O_4$ by using the molar mass of $C_6H_{14}O_4$ as a conversion factor.

$$? \text{ mol } C_6H_{14}O_4 = 6.86 \text{ g } C_6H_{14}O_4 \times \frac{1 \text{ mol } C_6H_{14}O_4}{150.2 \text{ g } C_6H_{14}O_4}$$
$$= 0.0457 \text{ mol } C_6H_{14}O_4$$

Convert from moles of $C_6H_{14}O_4$ to moles of O_2 by using the stoichiometric factor.

$$? \text{ mol } O_2 = 0.0457 \text{ mol } C_6H_{14}O_4 \times \frac{15 \text{ mol } O_2}{2 \text{ mol } C_6H_{14}O_4}$$
$$= 0.0343 \text{ mol } O_2$$

Convert from moles of O_2 to grams of O_2 by using the molar mass of O_2 as a conversion factor.

$$? \text{ g } O_2 = 0.343 \text{ mol } O_2 \times \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2}$$
$$= 11.0 \text{ g } O_2$$

As in Example 4-4, the three steps can be combined into a single calculation, as shown below.

$$? \text{ g } O_2 = 6.86 \text{ g } C_6H_{14}O_4 \times \frac{1 \text{ mol } C_6H_{14}O_4}{150.2 \text{ g } C_6H_{14}O_4} \times \frac{15 \text{ mol } O_2}{2 \text{ mol } C_6H_{14}O_6} \times \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} = 11.0 \text{ g } O_2$$

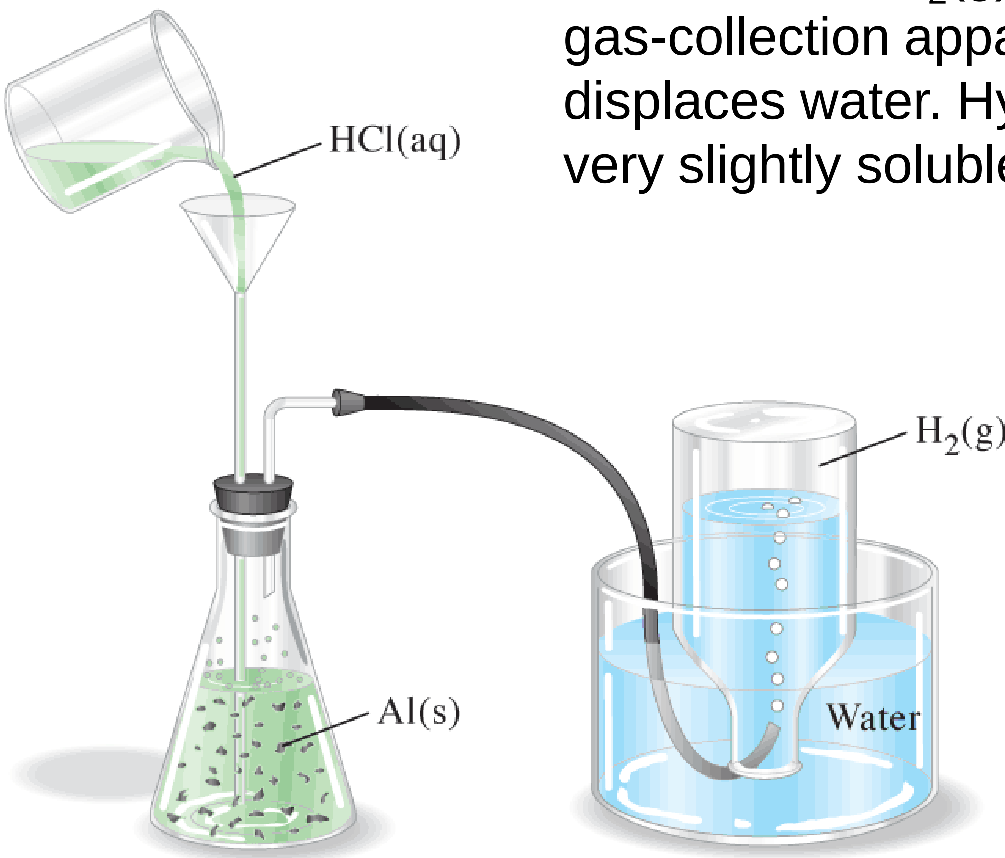
Assess

Focus on the single line calculation shown above. A quick scan of the numbers to the right of 6.86 g $C_6H_{14}O_4$ indicates that the stoichiometric factor has a value of 7.5; the product, 7.5×32.00, is about 250, which, when divided by 150.2, yields a factor of about $250/150 = 5/3$. The mass of O_2 should be about 5/3 that of the $C_6H_{14}O_4$, and it is—that is, compare 11.0 with 5/3 of 6.86, which is about 35/3 or somewhat less than 12. As in Example 4-4, note that the proper cancellation of units occurs.

PRACTICE EXAMPLE A: For the reaction in Example 4-1, how many grams of NH_3 are consumed per gram of O_2?

The reaction between solid aluminum, Al(s), and aqueous hydrochloric acid, HCl(aq), can be used for preparing small volumes of hydrogen gas, $H_2(g)$ in the laboratory. A balanced chemical equation for the reaction is shown below.

The liberated $H_2(g)$ flows into a gas-collection apparatus, where it displaces water. Hydrogen is only very slightly soluble in water.



▲ FIGURE 4-4

The reaction $2\,Al(s) + 6\,HCl(aq) \longrightarrow 2\,AlCl_3(aq) + 3\,H_2(g)$

EXAMPLE 4-6 Additional Conversion Factors in a Stoichiometric Calculation: Volume, Density, and Percent Composition

An alloy used in aircraft structures consists of 93.7% Al and 6.3% Cu by mass. The alloy has a density of 2.85 g/cm^3. A 0.691 cm^3 piece of the alloy reacts with an excess of HCl(aq). If we assume that *all* the Al but *none* of the Cu reacts with HCl(aq), what is the mass of H_2 obtained? Refer to reaction (4.2).

Analyze

A simple approach to this calculation is outlined below. Each numbered arrow refers to a conversion factor that changes the unit on the left to the one on the right.

$$\text{cm}^3\text{ alloy} \xrightarrow{1} \text{g alloy} \xrightarrow{2} \text{g Al} \xrightarrow{3} \text{mol Al} \xrightarrow{4} \text{mol H}_2 \xrightarrow{5} \text{g H}_2$$

The calculation can be done in five distinct steps, or with a single setup in which the five conversions are performed in sequence.

Solve

Using a stepwise approach, we proceed as follows.

Convert from volume of alloy to grams of alloy by using the density as a conversion factor.

$$?\text{ g alloy} = 0.691\text{ cm}^3\text{ alloy} \times \frac{2.85\text{ g alloy}}{1\text{ cm}^3\text{ alloy}} = 1.97\text{ g alloy}$$

Convert from grams of alloy to grams of Al by using the percentage by mass of Al as a conversion factor.

$$?\text{ g Al} = 1.97\text{ g alloy} \times \frac{93.7\text{ g Al}}{100\text{ g alloy}} = 1.85\text{ g Al}$$

Convert from grams of Al to moles of Al by using the molar mass as a conversion factor.

$$?\text{ mol Al} = 1.85\text{ g Al} \times \frac{1\text{ mol Al}}{26.98\text{ g Al}} = 0.0684\text{ mol Al}$$

Convert from moles of Al to moles of H_2 by using the stoichiometric factor.

$$?\text{ mol H}_2 = 0.0684\text{ mol Al} \times \frac{3\text{ mol H}_2}{2\text{ mol Al}} = 0.103\text{ mol H}_2$$

Convert from moles of Al to moles of H_2 by using the stoichiometric factor.

$$?\text{ g H}_2 = 0.103\text{ mol H}_2 \times \frac{2.016\text{ g H}_2}{1\text{ mol H}_2} = 0.207\text{ g H}_2$$

Remember to store intermediate results without rounding off. When all of the steps are combined into a single calculation, we do not have to write down intermediate results and we reduce rounding errors.

$$?\text{ g H}_2 = 0.691\text{ cm}^3\text{ alloy} \times \frac{2.85\text{ g alloy}}{1\text{ cm}^3\text{ alloy}} \times \frac{93.7\text{ g Al}}{100\text{ g alloy}} \times \frac{1\text{ mol Al}}{26.98\text{ g Al}} \times \frac{3\text{ mol H}_2}{2\text{ mol Al}} \times \frac{2.016\text{ g H}_2}{1\text{ mol H}_2} = 0.207\text{ g H}_2$$

EXAMPLE 4-7 Additional Conversion Factors in a Stoichiometric Calculation: Volume, Density, and Percent Composition of a Solution

A hydrochloric acid solution consists of 28.0% HCl by mass and has a density of 1.14 g/mL. What volume of this solution is required to react completely with 1.87 g Al in reaction (4.2)?

Analyze

The first challenge here is to determine where to begin. Although the problem refers to 28.0% HCl and a density of 1.14 g/mL, the appropriate starting point is with the given information—1.87 g Al. The goal of our calculation is a solution volume—mL HCl solution.

$$\text{g Al} \xrightarrow{1} \text{mol Al} \xrightarrow{2} \text{mol HCl} \xrightarrow{3} \text{g HCl} \xrightarrow{4} \text{g HCl solution} \xrightarrow{5} \text{mL HCl solution}$$

The conversion factors in the calculation involve (1) the molar mass of Al, (2) stoichiometric coefficients from equation (4.2), (3) the molar mass of HCl, (4) the percent composition of the HCl solution, and (5) the density of the HCl solution.

Solve

Using a stepwise approach, we proceed as follows.

Convert from grams of Al to moles of Al by using the molar mass of Al.

$$?\ \text{mol Al} = 1.87\ \text{g Al} \times \frac{1\ \text{mol Al}}{26.98\ \text{g Al}} = 0.0693\ \text{mol Al}$$

Convert from moles of Al to moles of HCl by using the stoichiometric factor.

$$?\ \text{mol HCl} = 0.0693\ \text{mol Al} \times \frac{6\ \text{mol HCl}}{2\ \text{mol Al}} = 0.208\ \text{mol HCl}$$

Convert from moles of HCl to grams of HCl by using the molar mass of HCl.

$$?\ \text{g HCl} = 0.208\ \text{mol HCl} \times \frac{36.46\ \text{g HCl}}{1\ \text{mol HCl}} = 7.58\ \text{g HCl}$$

Convert from grams of HCl to grams of HCl solution by using the percentage by mass.

$$?\ \text{g HCl soln} = 7.58\ \text{g HCl} \times \frac{100\ \text{g HCl soln}}{28.0\ \text{g HCl}} = 27.1\ \text{g HCl soln}$$

Convert from grams of HCl solution to milliliters of HCl solution by using the density.

$$?\ \text{mL HCl soln} = 27.1\ \text{g HCl soln} \times \frac{1\ \text{mL HCl soln}}{1.14\ \text{g HCl soln}} = 23.8\ \text{mL HCl soln}$$

In the conversion pathway approach, we combine the individual steps into a single line.

$$\left(\text{g Al} \xrightarrow{1} \text{mol Al} \xrightarrow{2} \text{mol HCl} \xrightarrow{3} \text{g HCl} \xrightarrow{4} \text{g HCl soln} \xrightarrow{5} \text{mL HCl soln}\right)$$

$$?\ \text{mL HCl soln} = 1.87\ \text{g Al} \times \frac{1\ \text{mol Al}}{26.98\ \text{g Al}} \times \frac{6\ \text{mol HCl}}{2\ \text{mol Al}} \times \frac{36.46\ \text{g HCl}}{1\ \text{mol HCl}} \times \frac{100.0\ \text{g HCl soln}}{28.0\ \text{g HCl}} \times \frac{1\ \text{mL HCl soln}}{1.14\ \text{g HCl soln}} = 23.8\ \text{mL HCl soln}$$

4-3 Chemical Reactions in Solution

Close contact between atoms, ions and molecules necessary for a reaction to occur.

Components of a solution are;

-solvent, determines whether the solution exists as a solid, liq, or gas. We usually use *aqueous* (aq) solution.

-solutes, a material dissolved in the solvent.

NaCl (aq)→ liquid water (solvent) , NaCl (solute)

Molarity:

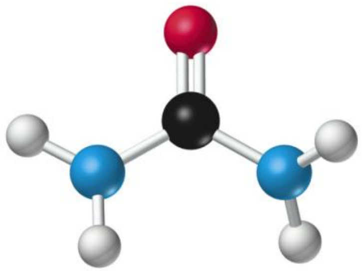
<u>Molarity, M</u> is a solution property, concentration unit

$$Molarity\ (M) = \frac{Amount\ of\ solute\ (in\ moles)}{Volume\ of\ solution\ (in\ liters)}$$

$$\boxed{M = \frac{n}{V}}$$

If 0.440 mol of urea is dissolved in enough water to make 1.0 L of solution the concentration is:

$$c_{\text{urea}} = \frac{0.440\ \text{mol urea}}{1.000\ \text{L}} = 0.440\ \text{M}\ \ CO(NH_2)_2$$

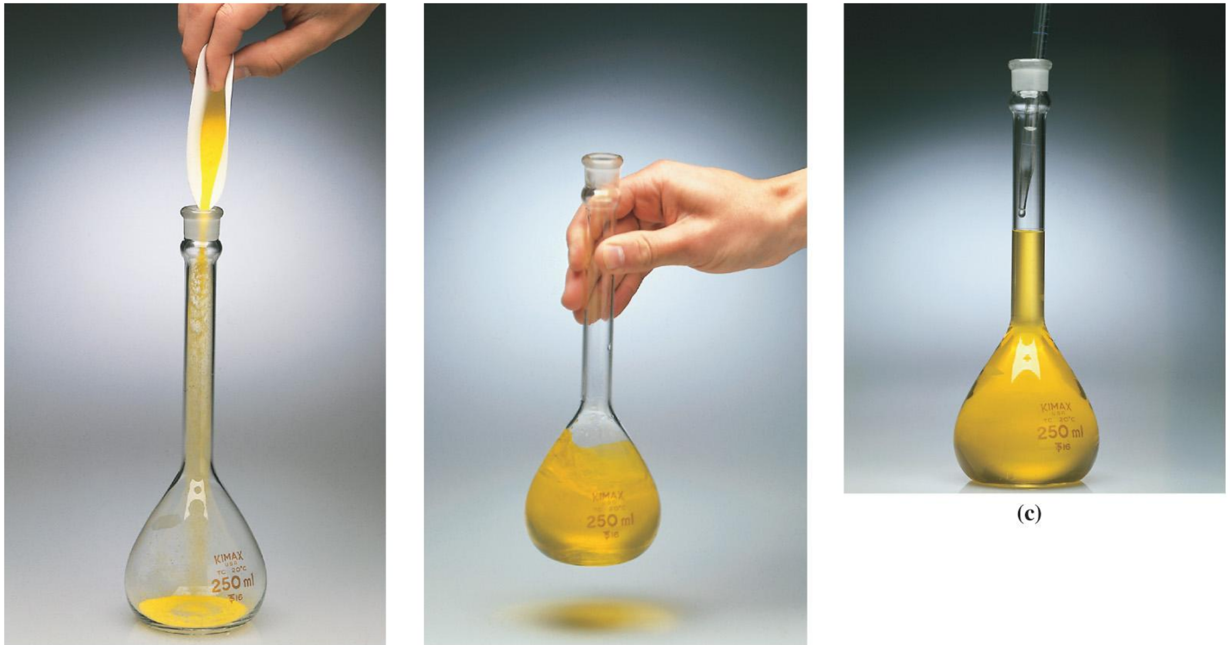


Urea

Alternatively, if 0.110 mol urea is present in 250.0 mL of solution, the solution is also 0.440 M.

$$c_{\text{urea}} = \frac{0.110\ \text{mol urea}}{0.250\ \text{L}} = 0.440\ \text{M}\ \ CO(NH_2)_2$$

Preparation of a Solution



(a) (b) (c)

-Weigh the solid sample.
-Dissolve it in a volumetric flask partially filled with solvent.
-Carefully fill to the mark.

EXAMPLE 4-8 Calculating Molarity from Measured Quantities

A solution is prepared by dissolving 25.0 mL ethanol, CH_3CH_2OH ($d = 0.789$ g/mL), in enough water to produce 250.0 mL solution. What is the molarity of ethanol in the solution?

Analyze

We must first calculate how many moles of ethanol are in a 25.0 mL sample of pure ethanol. This calculation requires the following conversions: mL ethanol $\xrightarrow{1}$ g ethanol $\xrightarrow{2}$ mol ethanol. The first conversion uses the density as a conversion factor and the second conversion uses molar mass as a conversion factor. The molarity of the solution is then calculated by using equation (4.3).

Solve

The number of moles of ethanol in a 25.0 mL sample of pure ethanol is calculated below in a single line.

$$? \text{ mol } CH_3CH_2OH = 25.0 \text{ mL } CH_3CH_2OH \times \frac{0.789 \text{ g } CH_3CH_2OH}{1 \text{ mL } CH_3CH_2OH} \times \frac{1 \text{ mol } CH_3CH_2OH}{46.07 \text{ g } CH_3CH_2OH}$$

$$= 0.428 \text{ mol } CH_3CH_2OH$$

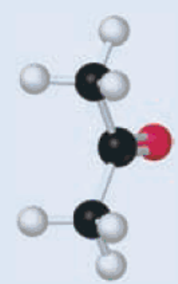
To apply the definition of molarity given in expression (4.3), note that 250.0 mL = 0.2500 L.

$$\text{molarity} = \frac{0.428 \text{ mol } CH_3CH_2OH}{0.2500 \text{ L soln}} = 1.71 \text{ M } CH_3CH_2OH$$

Assess

It is important to include the units in this calculation to ensure that we obtain the correct units for the final answer. When dealing with liquid solutes, be careful to distinguish between mL solute and mL soln.

PRACTICE EXAMPLE A: A 22.3 g sample of acetone (see the model here) is dissolved in enough water to produce 1.25 L of solution. What is the molarity of acetone in this solution?



Acetone

PRACTICE EXAMPLE B: If 15.0 mL of acetic acid, CH_3COOH ($d = 1.048$ g/mL), is dissolved in enough water to produce 500.0 mL of solution, then what is the molarity of acetic acid in the solution?

EXAMPLE 4-9

Calculating the Mass of solute in a solution of Known Molarity. *We want to prepare exactly 0.2500 L (250 mL) of an 0.250 M K_2CrO_4 solution in water. What mass of K_2CrO_4 should we use?*

Plan strategy:

$$\text{Volume} \xrightarrow{\frac{mol}{L}} \text{moles} \xrightarrow{\frac{g}{mol}} \text{mass}$$

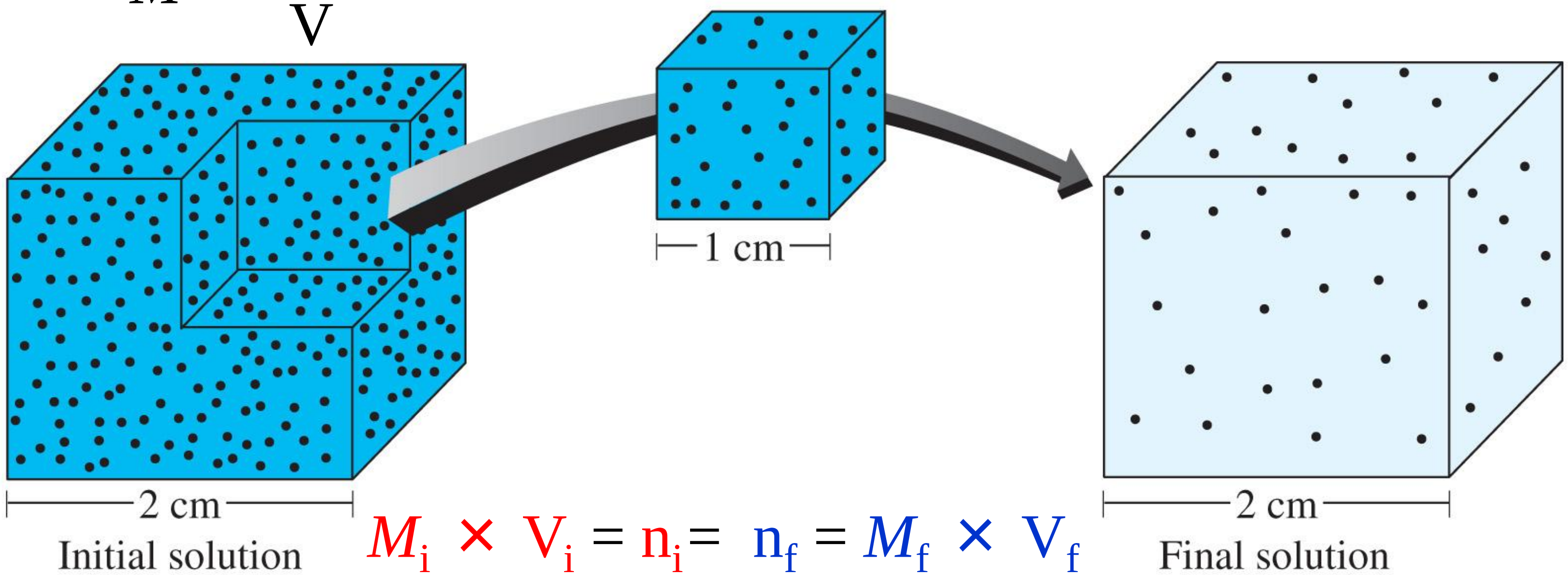
Write equation and calculate: *We need 2 conversion factors!*

$$m_{K_2CrO_4} = 0.2500\ L \times \frac{0.250\ mol}{1.00\ L} \times \frac{194.02\ g}{1.00\ mol} = 12.1\ g$$

Solution Dilution:

$$M = \frac{n}{V}$$

$$M_i \times V_i \qquad\qquad M_f \times V_f$$



$$M_i \times V_i = n_i = n_f = M_f \times V_f$$

$$M_f = \frac{M_i \times V_i}{V_f} = M_i \frac{V_i}{V_f}$$

▲ Figure 4-6

Visualizing the dilution of a solution

EXAMPLE 4-10

Preparing a solution by dilution: What volume of 0.250 M K_2CrO_4 should we use to prepare 0.250 L of 0.0100 M K_2CrO_4?

Plan strategy:

$$M_f = M_i \frac{V_i}{V_f} \qquad V_i = V_f \frac{M_f}{M_i}$$

Calculate:

$$V_{K_2CrO_4} = 0.2500\ \mathrm{L} \times \frac{0.0100\ \mathrm{mol}}{1.00\ \mathrm{L}} \times \frac{1.000\ \mathrm{L}}{0.250\ \mathrm{mol}} = 0.0100\ \mathrm{L}$$

(a) (b) (c)

▲ FIGURE 4-7
Preparing a solution by dilution—Example 4-10 illustrated

(a) A pipet is used to withdraw a 10.0 mL sample of 0.250 M K_2CrO_4.
(b) The pipetful of 0.250 M K_2CrO_4 is discharged into a 250.0 mL volumetric flask.
(c) Water is then added to bring the level of the solution to the calibration mark on the neck of the flask.

At this point, the solution is 0.0100 M K2CrO4.

EXAMPLE 4-11 Relating the Mass of a Product to the Volume and Molarity of a Reactant Solution

A 25.00 mL pipetful of 0.250 M K_2CrO_4 is added to an excess of $AgNO_3(aq)$. What mass of Ag_2CrO_4 will precipitate from the solution?

$$K_2CrO_4(aq) + 2\ AgNO_3(aq) \longrightarrow Ag_2CrO_4(s) + 2\ KNO_3(aq)$$

Analyze

The fact that an excess of $AgNO_3(aq)$ is used tells us that all of the K_2CrO_4 in the 25.00 mL sample of $K_2CrO_4(aq)$ is consumed. The calculation begins with a volume of 25.00 mL and ends with a mass of Ag_2CrO_4 expressed in grams. The conversion pathway is mL soln $\longrightarrow$ L soln $\longrightarrow$ mol K_2CrO_4 $\longrightarrow$ mol Ag_2CrO_4 $\longrightarrow$ g Ag_2CrO_4.

Solve

Let's solve this problem by using a stepwise approach.

Convert the volume of $K_2CrO_4(aq)$ from milliliters to liters, and then use molarity as a conversion factor between volume of solution and moles of solute (as in Example 4-10).

$$? \text{ mol } K_2CrO_4 = 25.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.250 \text{ mol } K_2CrO_4}{1 \text{ L}} = 6.25 \times 10^{-3} \text{ mol } K_2CrO_4$$

Use a stoichiometric factor from the equation to convert from moles of K_2CrO_4 to moles of Ag_2CrO_4.

$$? \text{ mol } Ag_2CrO_4 = 6.25 \times 10^{-3} \text{ mol } K_2CrO_4 \times \frac{1 \text{ mol } Ag_2CrO_4}{1 \text{ mol } K_2CrO_4} = 6.25 \times 10^{-3} \text{ mol } Ag_2CrO_4$$

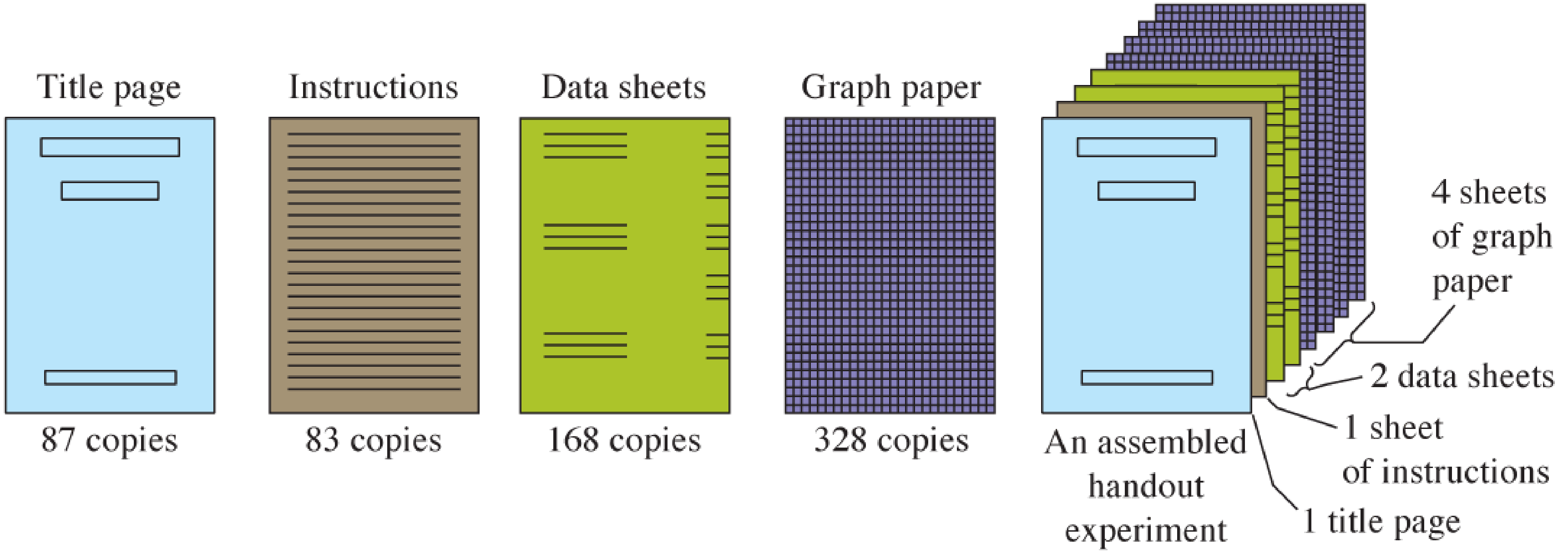
Use the molar mass to convert from moles to grams of Ag_2CrO_4.

$$? \text{ g } Ag_2CrO_4 = 6.25 \times 10^{-3} \text{ mol } Ag_2CrO_4 \times \frac{331.7 \text{ g } Ag_2CrO_4}{1 \text{ mol } Ag_2CrO_4} = 2.07 \text{ g } Ag_2CrO_4$$

The same final answer can be obtained more directly by combining the steps into a single line calculation.

$$? \text{ g } Ag_2CrO_4 = 25.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.250 \text{ mol } K_2CrO_4}{1 \text{ L}} \times \frac{1 \text{ mol } Ag_2CrO_4}{1 \text{ mol } K_2CrO_4} \times \frac{331.7 \text{ g } Ag_2CrO_4}{1 \text{ mol } Ag_2CrO_4}$$
$$= 2.07 \text{ g } Ag_2CrO_4$$

4-4 Determining Limiting Reagent



▲ FIGURE 4-8
An analogy to determining the limiting reactant in a chemical reaction—assembling a handout experiment

<u>Limiting Reactant</u>: The reactant that is completely consumed determines the quantities of the products formed.

4-4 Determining Limiting Reagent

Limiting Reactant: The reactant that is completely consumed determines the quantities of the products formed.

EXAMPLE 4-12

Phosphorus trichloride, PCl_3, is a commercially important compound used in the manufacture of pesticides, gasoline additives, and a number of other products. It is made by the direct combination of phosphorus and chlorine

$$P_4\,(s) + 6\,Cl_2\,(g) \rightarrow 4\,PCl_3\,(l)$$

What mass of PCl_3 forms in the reaction of 125 g P_4 with 323 g Cl_2?

$$n_{Cl_2} = 323 \text{ g } Cl_2 \times \frac{1 \text{ mol } Cl_2}{70.91 \text{ g } Cl_2} = 4.56 \text{ mol } Cl_2$$

$$n_{P_4} = 125 \text{ g } P_4 \times \frac{1 \text{ mol } P_4}{123.9 \text{ g } P_4} = 1.01 \text{ mol } P_4$$

6 moles of Cl_2 is required for 1 mole of P_4 to be completely used, we do not have that much, so, *Cl_2 is the limiting reactant.*

# of moles of Cl_2 used = 4.56 moles

# of moles of P_4 used = 4.56/6 =0.76 moles

# of moles of PCl_3 formed= (4.56/6) x 4 moles= 3.04 moles

mass PCl_3 formed = 3.04 moles PCl_3 x $\frac{137.3 \text{ g } PCl_3}{1 \text{ mole } PCl_3}$

= 417 g PCl_3

P_4 left = (1.0 - 0.76)=0.24 mol

= 0.24 x $\frac{123.9 \text{ g}}{1 \text{ mole } P_4}$

= 30.1 gr P_4

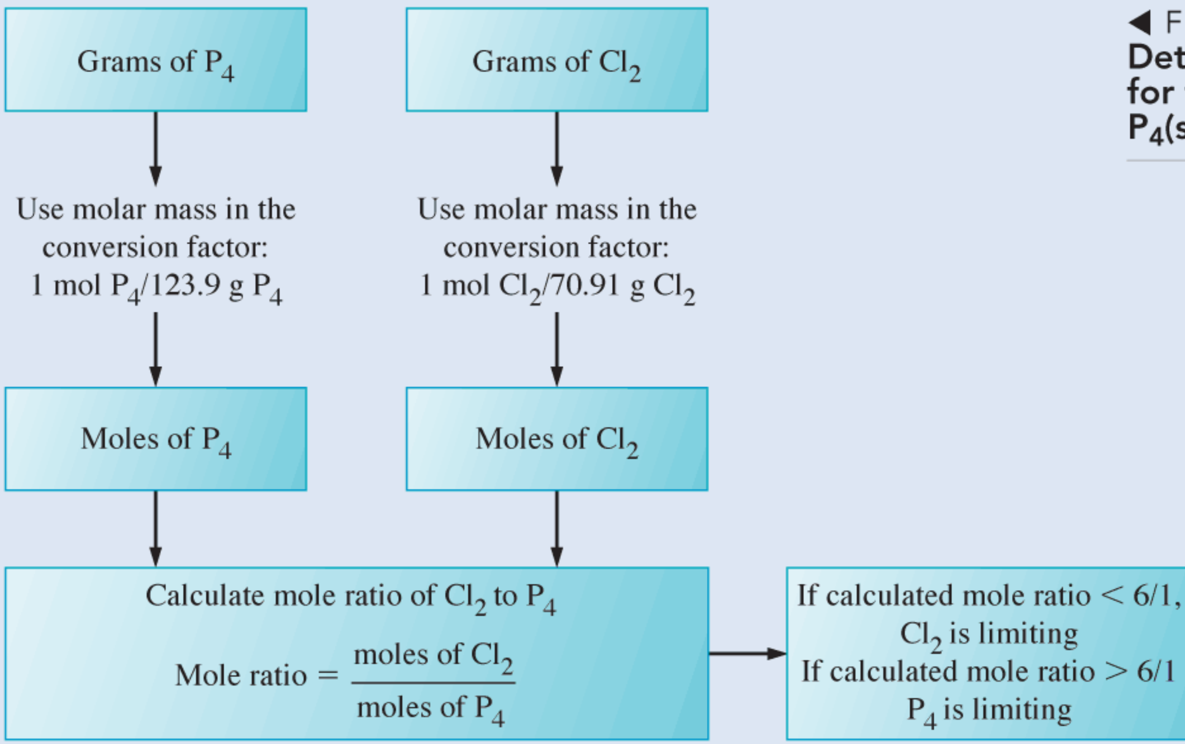


FIGURE 4-9
Determining the limiting reactant for the reaction
$$P_4(s) + 6\ Cl_2(g) \longrightarrow 4\ PCl_3(l)$$

4-5 Other Practical Matters in Reaction Stoichiometry

Theoretical yield is the expected yield from a reactant.
Actual yield is the amount of product actually produced.

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

Example: 4.15

What mass of cyclohexanol must we use to obtain 25 g cyclohexene, C_6H_{10}, if percent yield is 83% .

$$C_6H_{11}OH\ (l) \longrightarrow C_6H_{10}\ (l) + H_2O\ (l).$$

$$\text{Theor. Yield} = \frac{\text{Actual yield}}{\text{Percent yield}} \text{ x100\%}$$

$$= \frac{25 \text{ g x } 100\ \%}{83\ \%} = 30 \text{ gr } C_6H_{10}$$

$$?\ \text{g } C_6H_{11}OH = 30 \text{ g } C_6H_{10} \text{ x } \frac{1 \text{ mol } C_6H_{10}}{82.1 \text{ g } C_6H_{10}} \text{ x } \frac{1 \text{ mol } C_6H_{11}OH}{1 \text{ mol } C_6H_{10}} \text{ x } \frac{100.2 \text{ g } C_6H_{11}OH}{1 \text{ mol } C_6H_{11}OH}$$

$$= 37 \text{ g } C_6H_{11}OH$$

Theoretical, Actual and Percent Yield

When actual yield = % yield,

the reaction is said to be quantitative

Side reactions reduce the percent yield.

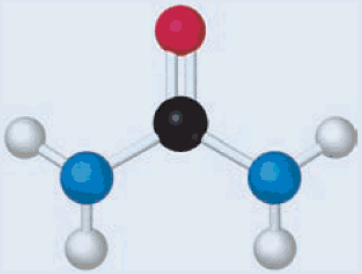
By-products are formed by side reactions.

EXAMPLE 4-14 Determining Theoretical, Actual, and Percent Yields

Billions of kilograms of urea, $CO(NH_2)_2$, are produced annually for use as a fertilizer. A ball-and-stick model of urea is shown here. The reaction used is given below.

$$2\,NH_3(g) + CO_2(g) \longrightarrow CO(NH_2)_2(s) + H_2O(l)$$

The typical starting reaction mixture has a 3:1 mole ratio of NH_3 to CO_2. If 47.7 g urea forms *per mole* of CO_2 that reacts, what is the **(a)** theoretical yield; **(b)** actual yield; and **(c)** percent yield?



Urea

Analyze

The reaction mixture contains fixed amounts of NH_3 and CO_2, and so we must first determine which reactant is the limiting reactant. The stoichiometric proportions are 2 mol NH_3:1 mol CO_2. In the reaction mixture, the mole ratio of NH_3 to CO_2 is 3:1. Therefore, NH_3 is the excess reactant and CO_2 is the limiting reactant. The calculation of the theoretical yield of urea must be based on the amount of CO_2, the limiting reactant. Because the quantity of urea is given per mole of CO_2, we should base the calculation on 1.00 mol CO_2. The following conversions are required: mol $CO_2 \rightarrow$ mol $CO(NH_2)_2 \rightarrow$ g $CO(NH_2)_2$.

Solve

(a) Let's calculate the theoretical yield by using a stepwise approach.

Convert from mol CO_2 to mol $CO(NH_2)_2$ by using the stoichiometric factor.

$$?\text{ mol } CO(NH_2)_2 = 1.00\text{ mol } CO_2 \times \frac{1\text{ mol } CO(NH_2)_2}{1\text{ mol } CO_2} = 1.00\text{ mol } CO(NH_2)_2$$

Convert from mol $CO(NH_2)_2$ to g $CO(NH_2)_2$ by using the molar mass of $CO(NH_2)_2$.

$$?\text{ g } CO(NH_2)_2 = 1.00\text{ mol } CO(NH_2)_2 \times \frac{60.1\text{ g } CO(NH_2)_2}{1\text{ mol } CO(NH_2)_2} = 60.1\text{ g } CO(NH_2)_2$$

Thus, 1.00 mol CO_2 is expected to yield 60.1 g $CO(NH_2)_2$, and so the theoretical yield of $CO(NH_2)_2$ is 60.1 g. As has been the case in all our examples, we could have combined the steps into a single line calculation.

$$\text{theoretical yield} = 1.00\text{ mol } CO_2 \times \frac{1\text{ mol } CO(NH_2)_2}{1\text{ mol } CO_2} \times \frac{60.1\text{ g } CO(NH_2)_2}{1\text{ mol } CO(NH_2)_2} = 60.1\text{ g } CO(NH_2)_2$$

(b) actual yield $= 47.7$ g $CO(NH_2)_2$

(c) $$\text{percent yield} = \frac{47.7\text{ g } CO(NH_2)_2}{60.1\text{ g } CO(NH_2)_2} \times 100\% = 79.4\%$$

Inc.

Consecutive Reactions, Simultaneous Reactions

- Multistep synthesis is often unavoidable.
- Reactions carried out in sequence to yield a final product are called consecutive reactions.
- In **simultaneous reactions,** two or more substances react independently of one another in separate reactions occurring at the same time.

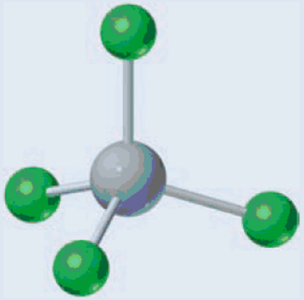
Overall Reactions and Intermediates

An **intermediate** is a substance produced in one step and consumed in another during a multistep synthesis.

• We can combine a series of chemical equations for *consecutive reactions* to obtain a single equation to represent the **overall reaction.**

The Overall Reaction is a chemical equation that expresses all the reactions occurring in a single overall equation.

EXAMPLE 4-16 Calculating the Quantity of a Substance Produced by Reactions Occurring Consecutively


Titanium tetrachloride

Titanium dioxide, TiO_2, is the most widely used white pigment for paints, having displaced most lead-based pigments, which are environmental hazards. Before it can be used, however, naturally occurring TiO_2 must be freed of colored impurities. One process for doing this converts impure $TiO_2(s)$ to $TiCl_4(g)$, which is then converted back to pure $TiO_2(s)$. The process is based on the following reactions, the first of which generates $TiCl_4$.

$$2\,TiO_2\,(\text{impure}) + 3\,C(s) + 4\,Cl_2(g) \longrightarrow 2\,TiCl_4(g) + CO_2(g) + 2\,CO(g)$$
$$TiCl_4(g) + O_2(g) \longrightarrow TiO_2(s) + 2\,Cl_2(g)$$

What mass of carbon is consumed in producing 1.00 kg of pure $TiO_2(s)$ in this process?

Analyze

In this calculation, we begin with the product, TiO_2, and work backward to one of the reactants, C. The following conversions are required.

$$\text{kg } TiO_2 \longrightarrow \text{g } TiO_2 \longrightarrow \text{mol } TiO_2 \xrightarrow{(a)} \text{mol } TiCl_4 \xrightarrow{(b)} \text{mol C} \longrightarrow \text{g C}$$

In the conversion from mol TiO_2 to mol $TiCl_4$, labeled (a), we focus on the second reaction. In the conversion from mol $TiCl_4$ to mol C, labeled (b), we focus on the first reaction.

Solve

Using a stepwise approach, we proceed as follows.

Convert from kg TiO_2 to g TiO_2 and then to mol TiO_2 by using the molar mass of TiO_2.

$$?\text{ mol } TiO_2 = 1.00\text{ kg } TiO_2 \times \frac{1000\text{ g } TiO_2}{1\text{ kg } TiO_2} \times \frac{1\text{ mol } TiO_2}{79.88\text{ g } TiO_2} = 12.5\text{ mol } TiO_2$$

Convert from mol TiO_2 to mol $TiCl_4$ by using the stoichiometric factor from the second reaction.

$$?\text{ mol } TiCl_4 = 12.5\text{ mol } TiO_2 \times \underbrace{\frac{1\text{ mol } TiCl_4}{1\text{ mol } TiO_2}}_{(a)} = 12.5\text{ mol } TiCl_4$$

Convert from mol $TiCl_4$ to mol C by using the stoichiometric factor from the first reaction.

$$?\text{ mol C} = 12.5\text{ mol } TiCl_4 \times \underbrace{\frac{3\text{ mol C}}{2\text{ mol } TiCl_4}}_{(b)} = 18.8\text{ mol C}$$

Convert from mol C to g C by using the molar mass of C.

$$?\text{ g C} = 18.8\text{ mol C} \times \frac{12.01\text{ g C}}{1\text{ mol C}} = 226\text{ g C}$$

The conversions given above can be combined into a single line, as shown below.

$$?\text{ g C} = 1.00\text{ kg } TiO_2 \times \frac{1000\text{ g } TiO_2}{1\text{ kg } TiO_2} \times \frac{1\text{ mol } TiO_2}{79.88\text{ g } TiO_2} \times \overset{(a)}{\frac{1\text{ mol } TiCl_4}{1\text{ mol } TiO_2}} \times \overset{(b)}{\frac{3\text{ mol C}}{2\text{ mol } TiCl_4}} \times \frac{12.01\text{ g C}}{1\text{ mol C}} = 226\text{ g C}$$

EXAMPLE 4-17 Calculating the Quantity of a Substance Produced by Reactions Occurring Simultaneously

Magnesium–aluminum alloys are widely used in aircraft construction. One particular alloy contains 70.0% Al and 30.0% Mg, by mass. How many grams of $H_2(g)$ are produced in the reaction of a 0.710 g sample of this alloy with excess HCl(aq)? Balanced chemical equations are given below for the reactions that occur.

$$2\,Al(s) + 6\,HCl(aq) \longrightarrow 2\,AlCl_3(aq) + 3\,H_2(g)$$
$$Mg(s) + 2\,HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

Analyze

The two reactions given above are simultaneous reactions. Simultaneous reactions occur independently; thus, we have two conversion pathways to consider:

$$(1)\quad \text{g alloy} \longrightarrow \text{g Al} \longrightarrow \text{mol Al} \xrightarrow{(a)} \text{mol } H_2;$$

$$(2)\quad \text{g alloy} \longrightarrow \text{g Mg} \longrightarrow \text{mol Mg} \xrightarrow{(b)} \text{mol } H_2;$$

Pathways (1) and (2) are based on the first and second reactions, respectively. The total amount of H_2 produced is obtained by adding together the amounts produced by each reaction. The conversion from mol Al to mol H_2 requires a stoichiometric factor, labeled (a). The conversion from mol Mg to mol H_2 requires a different stoichiometric factor, labeled (b).

Solve

Convert from g alloy to mol Al and from g alloy to mol Mg by using the mass percentages of Al and Mg and the molar masses of Al and Mg.

$$?\ \text{mol Al} = 0.710\ \text{g alloy} \times \frac{70.0\ \text{g Al}}{100.0\ \text{g alloy}} \times \frac{1\ \text{mol Al}}{26.98\ \text{g Al}}$$
$$= 0.0184\ \text{mol Al}$$

$$?\ \text{mol Mg} = 0.710\ \text{g alloy} \times \frac{30.0\ \text{g Mg}}{100.0\ \text{g alloy}} \times \frac{1\ \text{mol Mg}}{24.31\ \text{g Mg}}$$
$$= 8.76 \times 10^{-3}\ \text{mol Mg}$$

Convert from mol Al to mol H_2 and from mol Mg to mol H_2 by using stoichiometric factors (a) and (b). The total number of moles of H_2 is obtained by adding together the two independent contributions.

$$? \text{ mol } H_2 = 0.0184 \text{ mol Al} \times \underbrace{\frac{3 \text{ mol } H_2}{2 \text{ mol Al}}}_{(a)} +$$

$$8.76 \times 10^{-3} \text{ mol Mg} \times \underbrace{\frac{1 \text{ mol } H_2}{1 \text{ mol Mg}}}_{(b)} = 0.0364 \text{ mol } H_2$$

Convert from mol H_2 to g H_2 by using the molar mass of H_2 as a conversion factor.

$$? \text{ g } H_2 = 0.0364 \text{ mol } H_2 \times \frac{2.016 \text{ g } H_2}{1 \text{ mol } H_2} = 0.0734 \text{ g } H_2$$

An alternative approach is to combine the steps into a single line calculation, as shown below.

$$? \text{ g } H_2 = \left(0.710 \text{ g alloy} \times \frac{70.0 \text{ g Al}}{100.0 \text{ g alloy}} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \underbrace{\frac{3 \text{ mol } H_2}{2 \text{ mol Al}}}_{(a)} \times \frac{2.016 \text{ g } H_2}{1 \text{ mol } H_2}\right)$$

$$+ \left(0.710 \text{ g alloy} \times \frac{30.0 \text{ g Mg}}{100.0 \text{ g alloy}} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} \times \underbrace{\frac{1 \text{ mol } H_2}{1 \text{ mol Mg}}}_{(b)} \times \frac{2.016 \text{ g } H_2}{1 \text{ mol } H_2}\right) = 0.0734 \text{ g } H_2$$

Assess

In this example, the composition of the alloy is given and we solved for the amount of H_2 that is produced. The inverse problem, in which we are given the amount of H_2 produced and are asked to determine the amounts of Al and Mg in the alloy, is a little harder to solve. See Practice Examples A and B below.

PRACTICE EXAMPLE A: A 1.00 g sample of a magnesium-aluminum alloy yields 0.107 g H_2 when treated with an excess of HCl(aq). What is the percentage by mass of Al in the alloy? [*Hint:* This is the inverse of Example 4-17. To solve this problem, let m and $1.00 - m$ be the masses of Al and Mg, respectively, and then use these masses in the setup above to develop an equation that relates m to the total mass of H_2 obtained. Then solve for m.]

PRACTICE EXAMPLE B: A 1.500 g sample of a mixture containing only Cu_2O and CuO was treated with hydrogen to produce copper metal and water. After the water evaporated, 1.2244 g of pure copper metal was recovered. What is the percentage by mass of Cu_2O in the original mixture? Balanced chemical equations are given below for the reactions involved.

$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$$
$$Cu_2O(s) + H_2(g) \longrightarrow 2\,Cu(s) + H_2O(l)$$