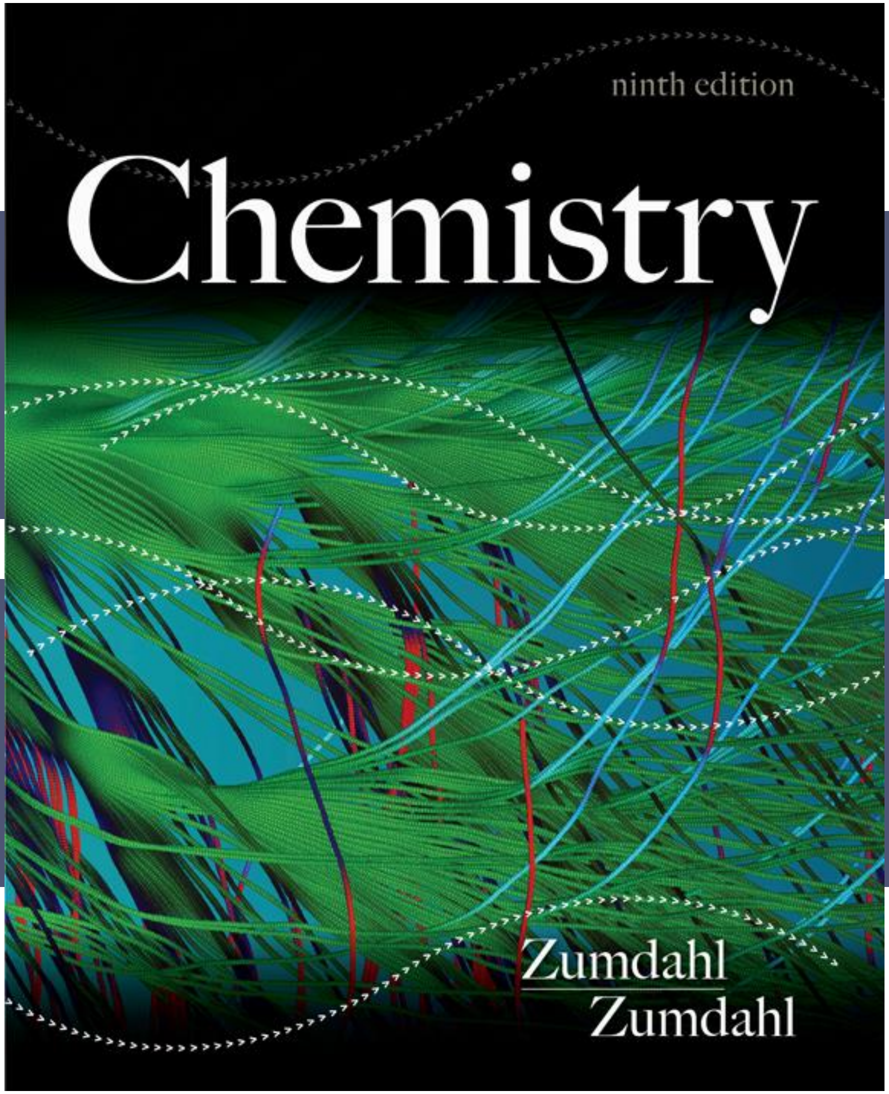


Chapter 3

Stoichiometry

Chemical Stoichiometry

- Stoichiometry – The study of quantities of materials consumed and produced in chemical reactions.

- Objects behave as though they were all identical.
- Atoms are too small to count.
- Need average mass of the object.

EXERCISE!

A pile of marbles weigh 394.80 g. 10 marbles weigh 37.60 g. How many marbles are in the pile?

$$\text{Avg. Mass of 1 Marble} = \frac{37.60\text{ g}}{10\text{ marbles}} = 3.76\text{ g / marble}$$

$$\frac{394.80\text{ g}}{3.76\text{ g}} = 105\text{ marbles}$$

- ^{12}C is the standard for atomic mass, with a mass of exactly 12 atomic mass units (u).
- The masses of all other atoms are given relative to this standard.
- Elements occur in nature as mixtures of isotopes.
- Carbon = 98.89% ^{12}C

  1.11% ^{13}C

  $< 0.01\%$ ^{14}C

Average Atomic Mass for Carbon

98.89% of 12 u + 1.11% of 13.0034 u =

exact number

(0.9889)(12 u) + (0.0111)(13.0034 u) =

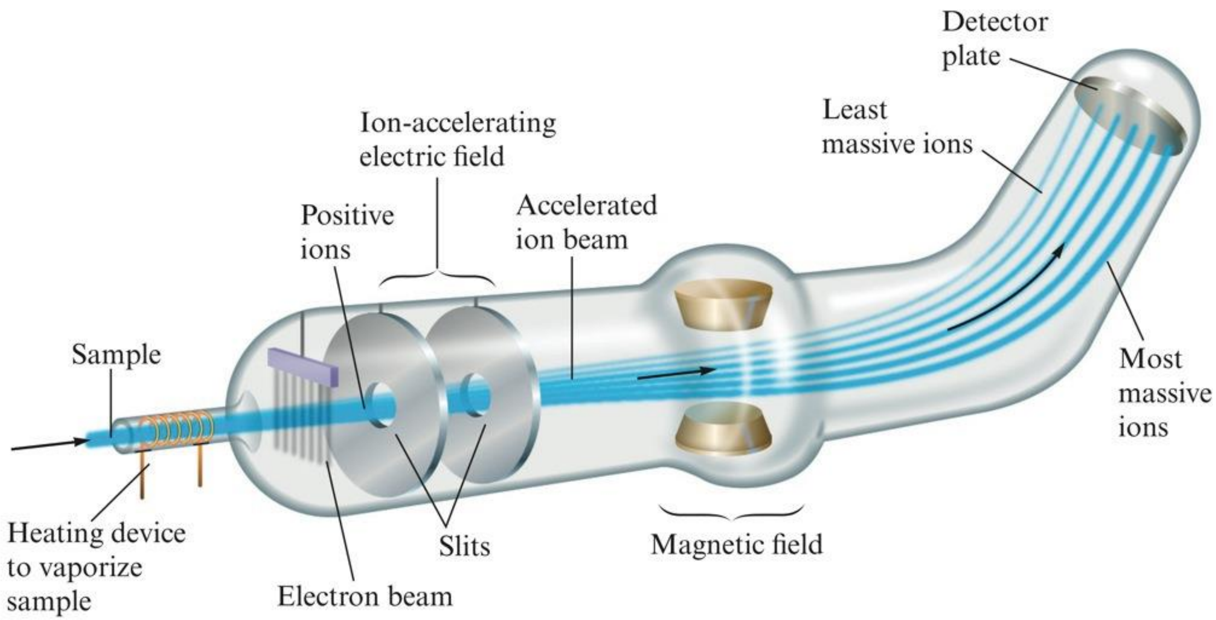
12.01 u

Average Atomic Mass for Carbon

- Even though natural carbon does not contain a single atom with mass 12.01, for stoichiometric purposes, we can consider carbon to be composed of only one type of atom with a mass of 12.01.
- This enables us to count atoms of natural carbon by weighing a sample of carbon.

Schematic Diagram of a Mass Spectrometer

Geoff Tompkinson/Photo Researchers, Inc.



EXERCISE!

An element consists of 62.60% of an isotope with mass 186.956 u and 37.40% of an isotope with mass 184.953 u.

- Calculate the average atomic mass and identify the element.

186.2 u

Rhenium (Re)

- The number equal to the number of carbon atoms in exactly 12 grams of pure ^{12}C.
- 1 mole of something consists of 6.022×10^{23} units of that substance (Avogadro's number).
- 1 mole C = 6.022×10^{23} C atoms = 12.01 g C

EXERCISE!

Calculate the number of iron atoms in a 4.48 mole sample of iron.

2.70×10^{24} Fe atoms

- Mass in grams of one mole of the substance:

  Molar Mass of N = 14.01 g/mol

  Molar Mass of H_2O = 18.02 g/mol

  (2 × 1.008 g) + 16.00 g

  Molar Mass of $Ba(NO_3)_2$ = 261.35 g/mol

  137.33 g + (2 × 14.01 g) + (6 × 16.00 g)

CONCEPT CHECK!

Which of the following is closest to the average mass of one atom of copper?

a) 63.55 g
b) 52.00 g
c) 58.93 g
d) 65.38 g
e) $\boxed{1.055 \times 10^{-22}\ g}$

CONCEPT CHECK!

Calculate the number of copper atoms in a 63.55 g sample of copper.

6.022×10^{23} Cu atoms

CONCEPT CHECK!

Which of the following 100.0 g samples contains the greatest number of atoms?

a) $\boxed{\text{Magnesium}}$
b) Zinc
c) Silver

EXERCISE!

Rank the following according to number of atoms (greatest to least):

a) 107.9 g of silver
b) 70.0 g of zinc
c) 21.0 g of magnesium

b) a) c)

EXERCISE!

Consider separate 100.0 gram samples of each of the following:

H_2O, N_2O, $C_3H_6O_2$, CO_2

- Rank them from greatest to least number of oxygen atoms.

H_2O, CO_2, $C_3H_6O_2$, N_2O

Conceptual Problem Solving

- Where are we going?
  - Read the problem and decide on the final goal.
- How do we get there?
  - Work backwards from the final goal to decide where to start.
- Reality check.
  - Does my answer make sense? Is it reasonable?

- Mass percent of an element:

$$\text{mass \%} = \frac{\text{mass of element in compound}}{\text{mass of compound}} \times 100\%$$

- For iron in iron(III) oxide, (Fe_2O_3):

$$\text{mass \% Fe} = \frac{2(55.85\text{ g})}{2(55.85\text{ g})+3(16.00\text{ g})} = \frac{111.70\text{ g}}{159.70\text{ g}} \text{ n } 100\% = 69.94\%$$

EXERCISE!

Consider separate 100.0 gram samples of each of the following:

H_2O, N_2O, $C_3H_6O_2$, CO_2

- Rank them from highest to lowest percent oxygen by mass.

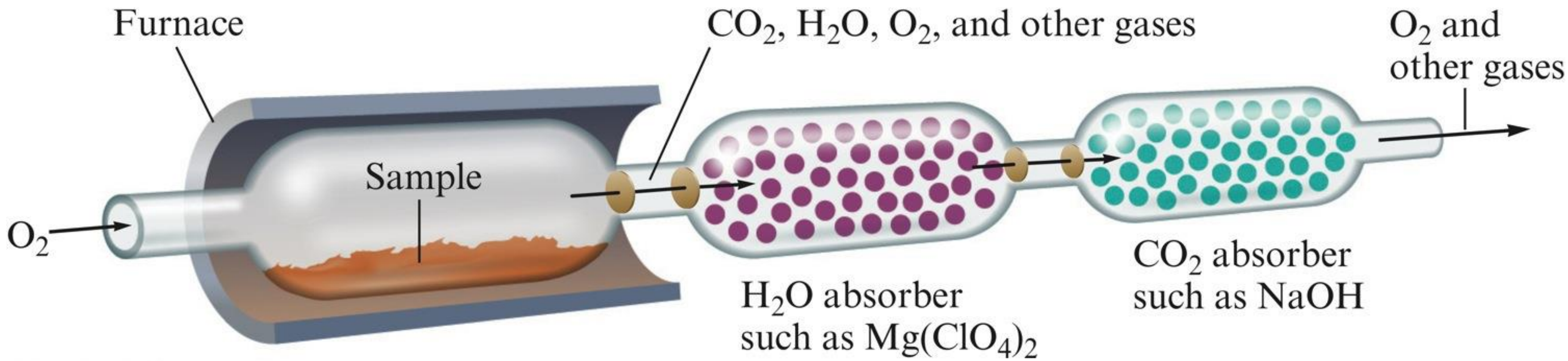
H_2O, CO_2, $C_3H_6O_2$, N_2O

Formulas

- Empirical formula = CH
  - Simplest whole-number ratio
- Molecular formula = (empirical formula)$_n$
  [n = integer]
- Molecular formula = C_6H_6 = $(CH)_6$
  - Actual formula of the compound

Analyzing for Carbon and Hydrogen

- Device used to determine the mass percent of each element in a compound.



EXERCISE!

The composition of adipic acid is 49.3% C, 6.9% H, and 43.8% O (by mass). The molar mass of the compound is about 146 g/mol.

- What is the empirical formula?

  $C_3H_5O_2$

- What is the molecular formula?

  $C_6H_{10}O_4$

- A representation of a chemical reaction:

$$C_2H_5OH + 3O_2 \longrightarrow 2CO_2 + 3H_2O$$

reactants products

- Reactants are only placed on the left side of the arrow, products are only placed on the right side of the arrow.

$$C_2H_5OH + 3O_2 \longrightarrow 2CO_2 + 3H_2O$$

- The equation is balanced.
- All atoms present in the reactants are accounted for in the products.
- 1 mole of ethanol reacts with 3 moles of oxygen to produce 2 moles of carbon dioxide and 3 moles of water.

- The balanced equation represents an overall ratio of reactants and products, not what actually "happens" during a reaction.
- Use the coefficients in the balanced equation to decide the amount of each reactant that is used, and the amount of each product that is formed.

Writing and Balancing the Equation for a Chemical Reaction

1. Determine what reaction is occurring. What are the reactants, the products, and the physical states involved?
2. Write the *unbalanced* equation that summarizes the reaction described in step 1.
3. Balance the equation by inspection, starting with the most complicated molecule(s). The same number of each type of atom needs to appear on both reactant and product sides. Do NOT change the formulas of any of the reactants or products.

EXERCISE!

Which of the following correctly balances the chemical equation given below? There may be more than one correct balanced equation. If a balanced equation is incorrect, explain what is incorrect about it.

$$CaO + C \longrightarrow CaC_2 + CO_2$$

I. $CaO_2 + 3C \longrightarrow CaC_2 + CO_2$

II. $\boxed{2CaO + 5C \longrightarrow 2CaC_2 + CO_2}$

III. $\boxed{CaO + (2.5)C \longrightarrow CaC_2 + (0.5)CO_2}$

IV. $\boxed{4CaO + 10C \longrightarrow 4CaC_2 + 2CO_2}$

CONCEPT CHECK!

Which of the following are true concerning balanced chemical equations? There may be more than one true statement.

I. The number of molecules is conserved.

II. The coefficients tell you how much of each substance you have.

III. Atoms are neither created nor destroyed.

IV. The coefficients indicate the mass ratios of the substances used.

V. The sum of the coefficients on the reactant side equals the sum of the coefficients on the product side.

Notice

- The number of atoms of each type of element must be the same on both sides of a balanced equation.
- Subscripts must not be changed to balance an equation.
- A balanced equation tells us the ratio of the number of molecules which react and are produced in a chemical reaction.
- Coefficients can be fractions, although they are usually given as lowest integer multiples.

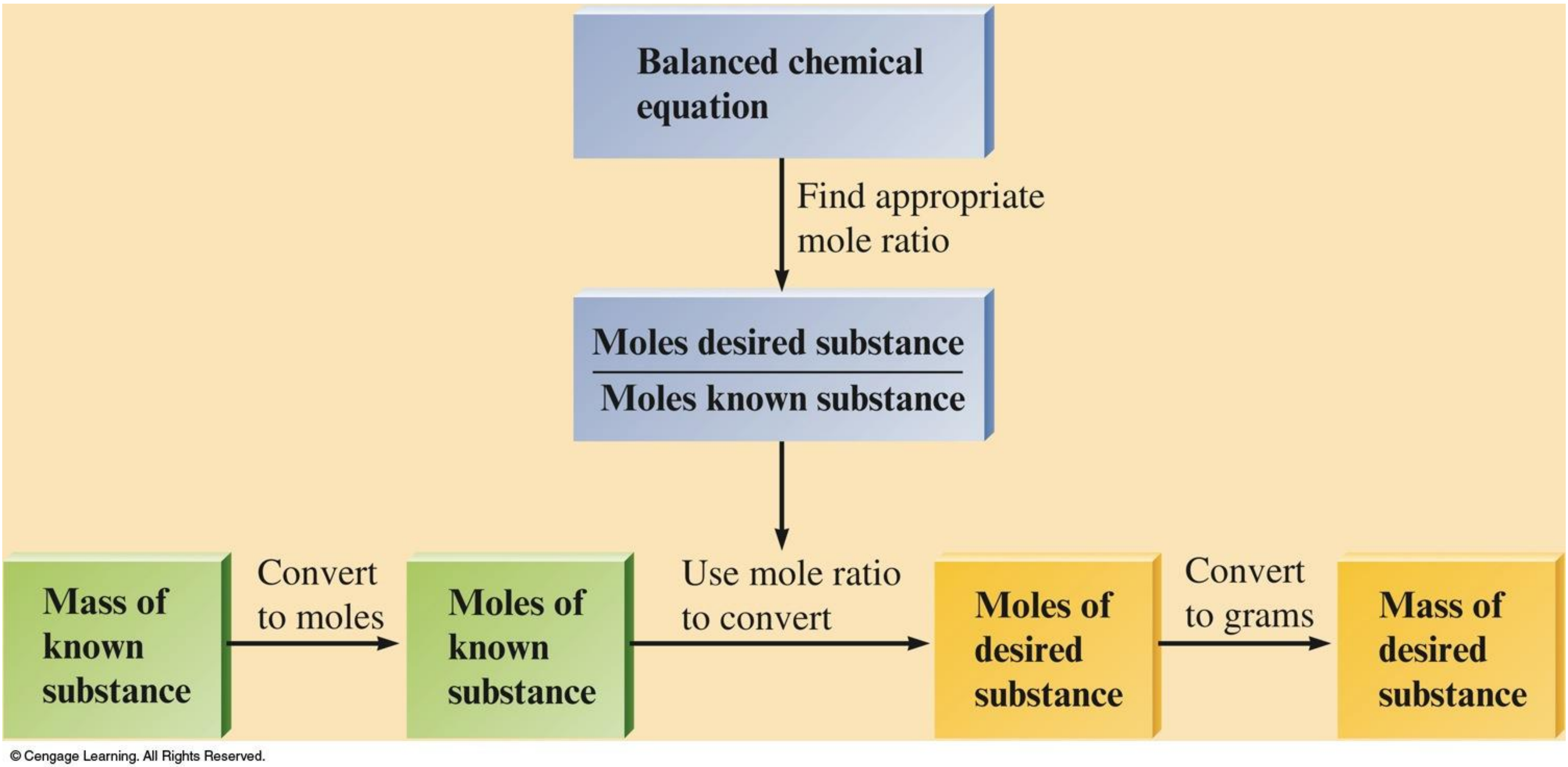
Stoichiometric Calculations

- Chemical equations can be used to relate the masses of reacting chemicals.

Calculating Masses of Reactants and Products in Reactions

1. Balance the equation for the reaction.
2. Convert the known mass of the reactant or product to moles of that substance.
3. Use the balanced equation to set up the appropriate mole ratios.
4. Use the appropriate mole ratios to calculate the number of moles of the desired reactant or product.
5. Convert from moles back to grams if required by the problem.

Calculating Masses of Reactants and Products in Reactions



EXERCISE!

Consider the following reaction:

$P_4(s) + 5\ O_2(g)$ ® $2\ P_2O_5(s)$

If 6.25 g of phosphorus is burned, what mass of oxygen does it combine with?

8.07 g O_2

EXERCISE! (Part I)

Methane (CH_4) reacts with the oxygen in the air to produce carbon dioxide and water.

Ammonia (NH_3) reacts with the oxygen in the air to produce nitrogen monoxide and water.

- Write balanced equations for each of these reactions.

EXERCISE! (Part II)

Methane (CH_4) reacts with the oxygen in the air to produce carbon dioxide and water.

Ammonia (NH_3) reacts with the oxygen in the air to produce nitrogen monoxide and water.

- What mass of ammonia would produce the same amount of water as 1.00 g of methane reacting with excess oxygen?

Let's Think About It

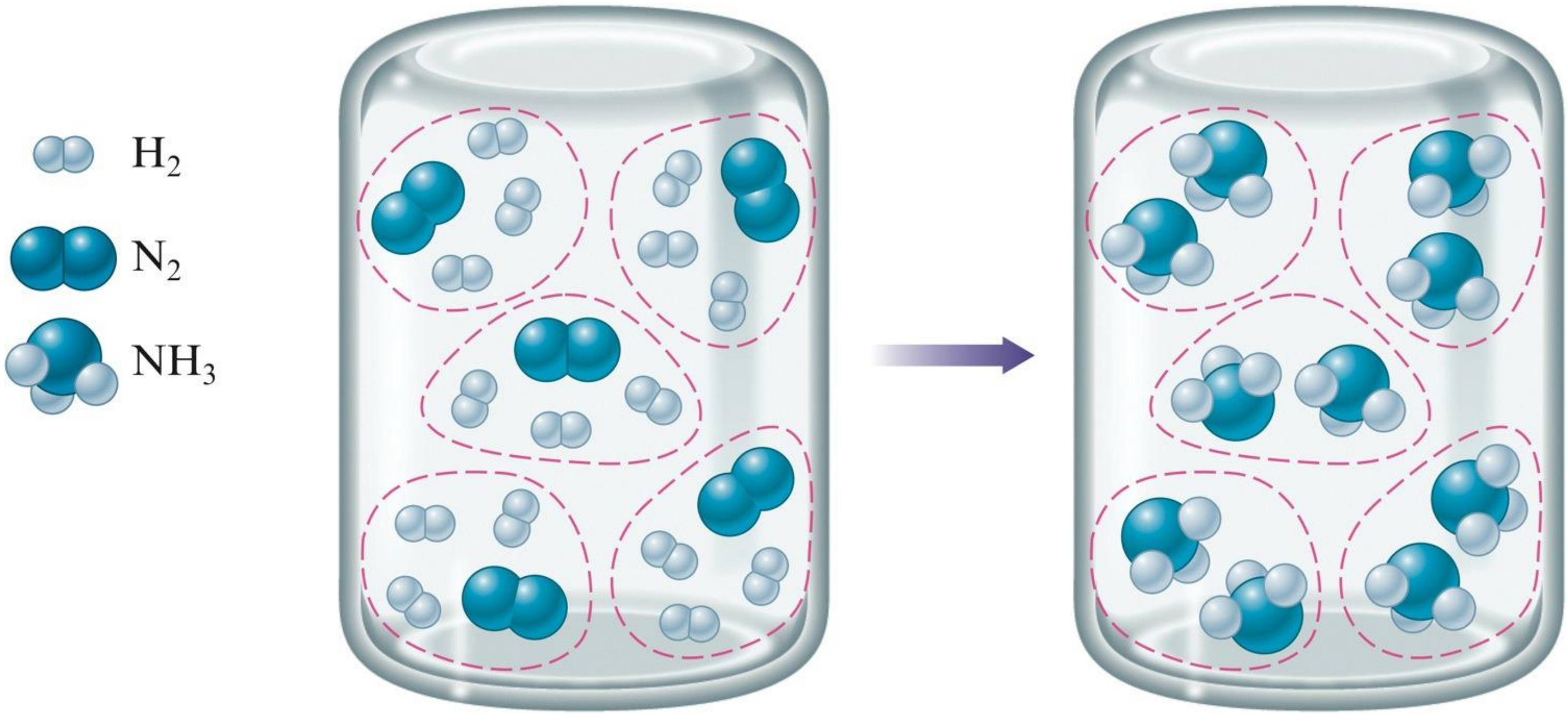
- Where are we going?
  - To find the mass of ammonia that would produce the same amount of water as 1.00 g of methane reacting with excess oxygen.
- How do we get there?
  - We need to know:
    - How much water is produced from 1.00 g of methane and excess oxygen.
    - How much ammonia is needed to produce the amount of water calculated above.

Limiting Reactants

- Limiting reactant – the reactant that runs out first and thus limits the amounts of products that can be formed.
- Determine which reactant is limiting to calculate correctly the amounts of products that will be formed.

A. The Concept of Limiting Reactants

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

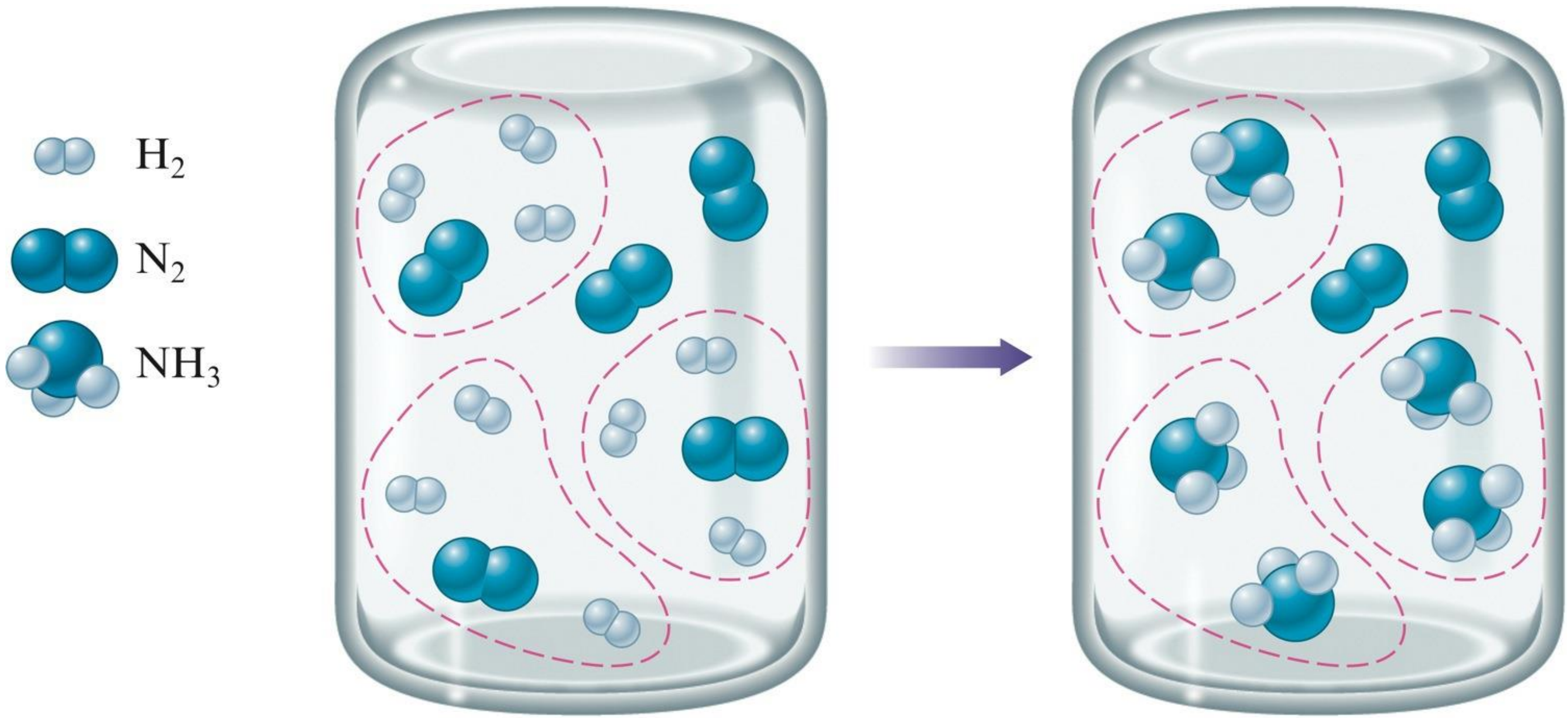


Before the reaction

After the reaction

A. The Concept of Limiting Reactants

- Limiting reactant mixture



Before the reaction

After the reaction

A. The Concept of Limiting Reactants

- Limiting reactant mixture
  - $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$
  - Limiting reactant is the reactant that runs out first.
    - H_2

Limiting Reactants

- The amount of products that can form is limited by the methane.
- Methane is the limiting reactant.
- Water is in excess.

CONCEPT CHECK!

Which of the following reaction mixtures could produce the greatest amount of product? Each involves the reaction symbolized by the equation:

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

a) 2 moles of H_2 and 2 moles of O_2
b) 2 moles of H_2 and 3 moles of O_2
c) 2 moles of H_2 and 1 mole of O_2
d) 3 moles of H_2 and 1 mole of O_2
e) Each produce the same amount of product.

Notice

- We cannot simply add the total moles of all the reactants to decide which reactant mixture makes the most product. We must always think about how much product can be formed by using what we are given, and the ratio in the balanced equation.

CONCEPT CHECK!

- You know that chemical A reacts with chemical B. You react 10.0 g of A with 10.0 g of B.
  - What information do you need to know in order to determine the mass of product that will be produced?

Let's Think About It

- Where are we going?
  - To determine the mass of product that will be produced when you react 10.0 g of A with 10.0 g of B.
- How do we get there?
  - We need to know:
    - The mole ratio between A, B, and the product they form. In other words, we need to know the balanced reaction equation.
    - The molar masses of A, B, and the product they form.

EXERCISE!

You react 10.0 g of A with 10.0 g of B. What mass of product will be produced given that the molar mass of A is 10.0 g/mol, B is 20.0 g/mol, and C is 25.0 g/mol? They react according to the equation:

$$A + 3B \longrightarrow 2C$$

Percent Yield

- An important indicator of the efficiency of a particular laboratory or industrial reaction.

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

EXERCISE!

Consider the following reaction:

$$P_4(s) + 6F_2(g) \longrightarrow 4PF_3(g)$$

- What mass of P_4 is needed to produce 85.0 g of PF_3 if the reaction has a 64.9% yield?

46.1 g P_4