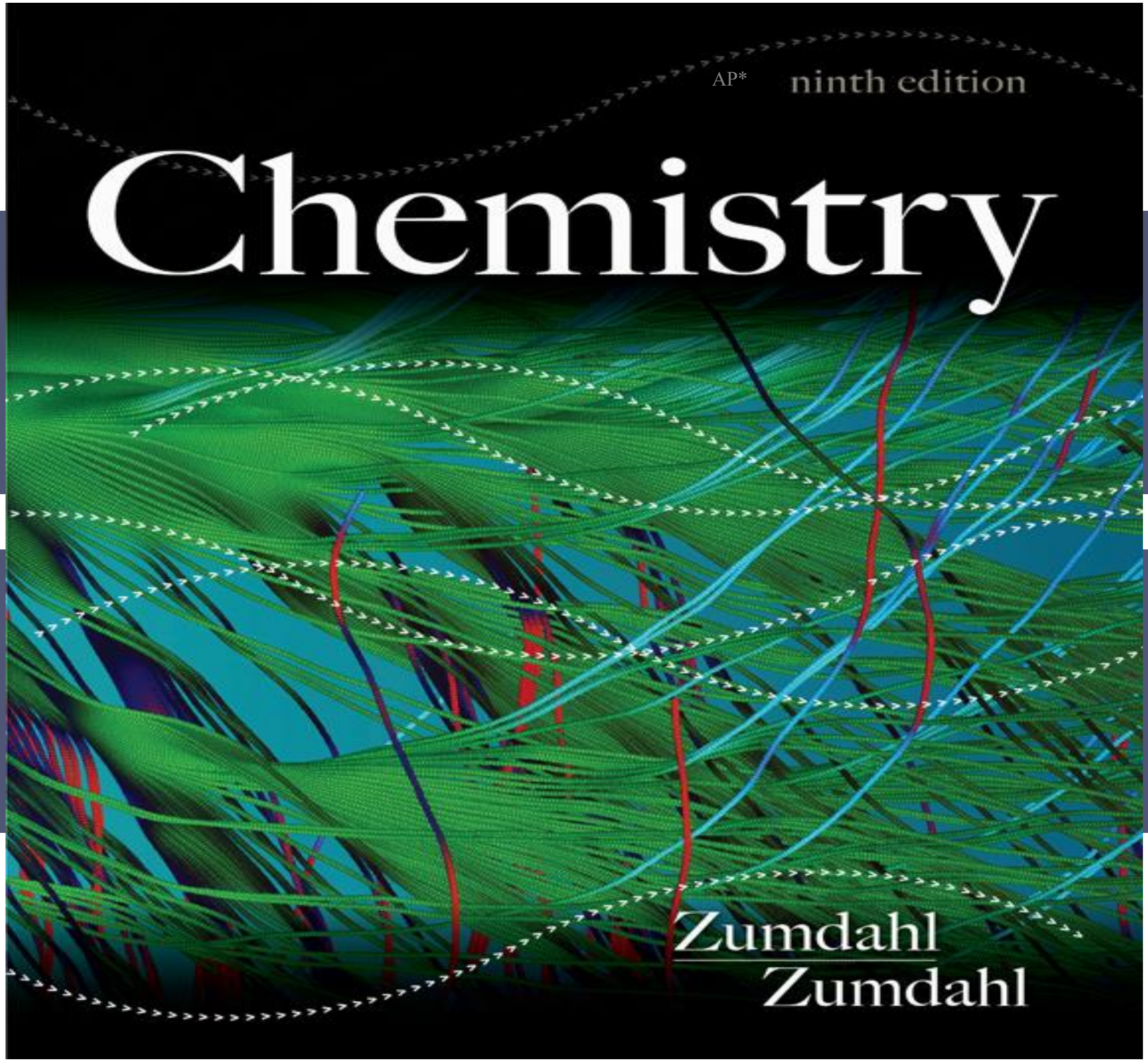


Chapter 3

Stoichiometry

Chemical Stoichiometry

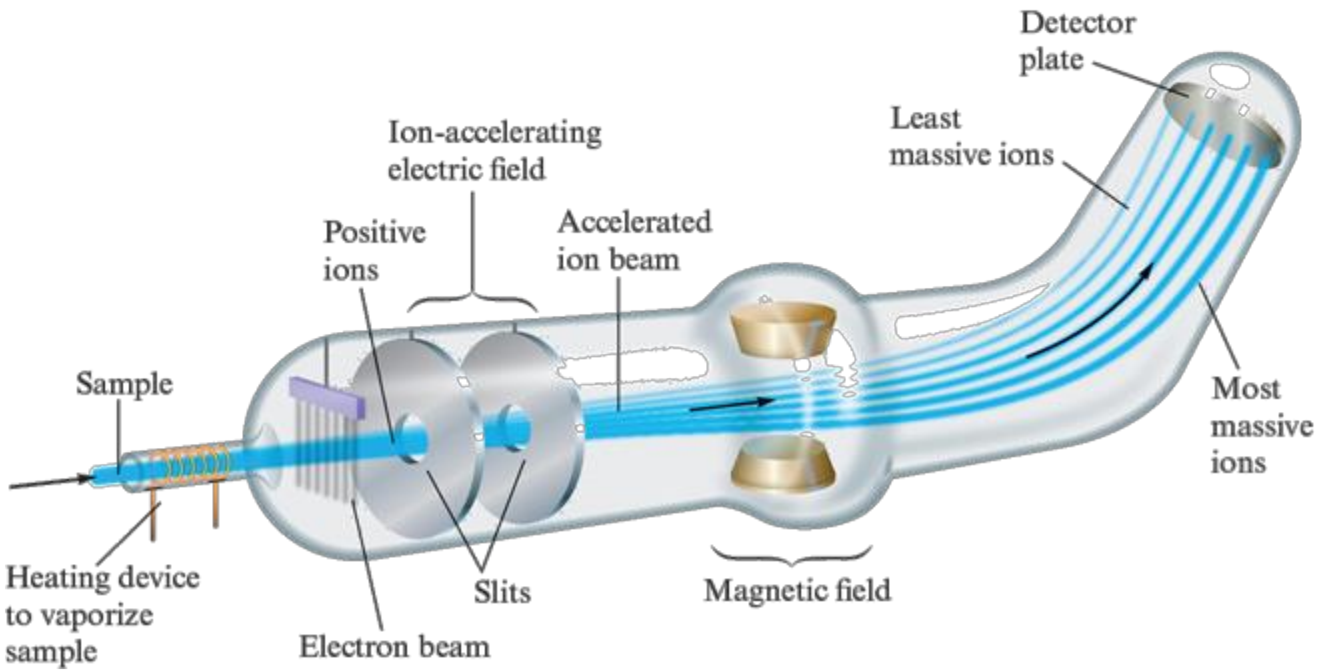
- Study of the quantities of materials consumed and produced in chemical reactions
- Requires understanding the concept of relative atomic masses

Modern System of Atomic Masses

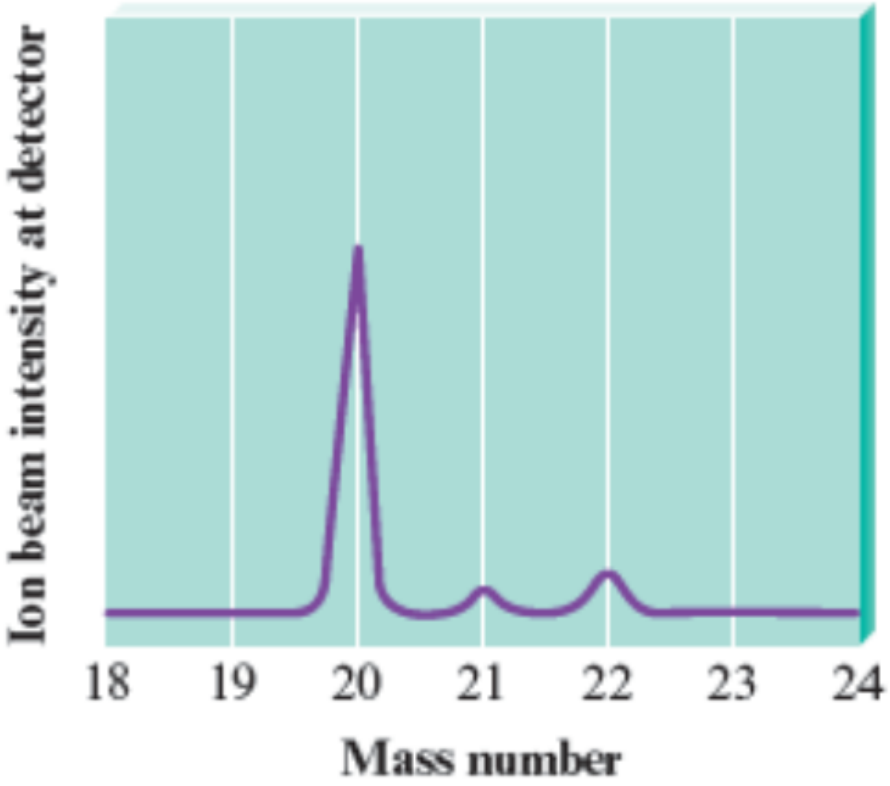
- Institued in 1961
- Standard - ^{12}C
  - ^{12}C is assigned a mass of exactly 12 atomic mass units (a.m.u) or(u) or Dalton .
    - Masses of all other atoms are given relative to this standard

Mass Spectrometer

- Helps to accurately compare the masses of atoms



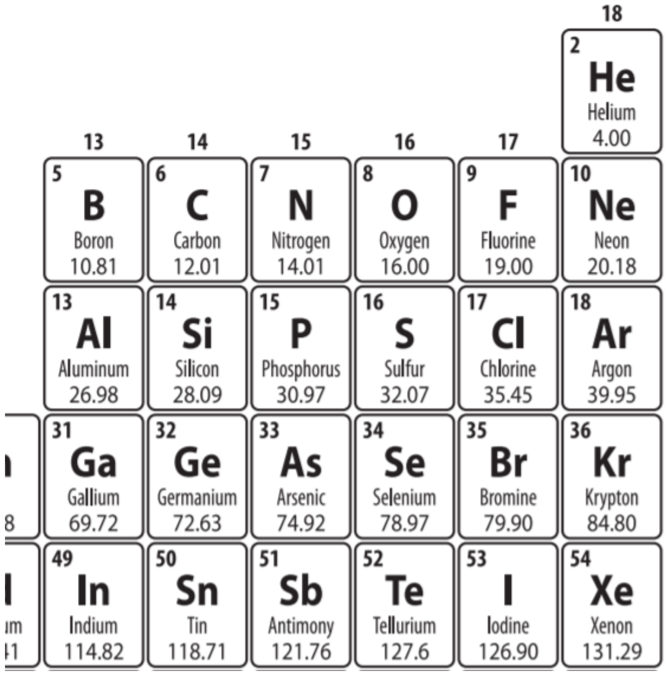
Mass Spectrometer



- 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

Average atomic mass=[atomic mass of isotope(1)*natural abundance(1)]+ [atomic mass of isotope(2)*natural abundance(2)]+ …….

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

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

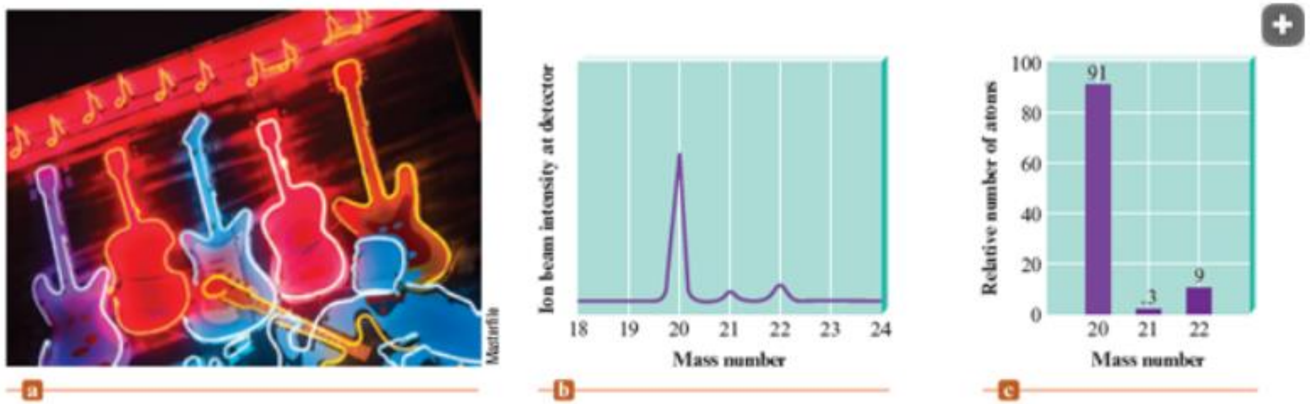
12.01 u

exact number

Mass Spectrometer

Figure 3.2

(**a**) Neon guitars on Beale Street in Memphis. The relative intensities of the signals recorded when natural neon is injected into a mass spectrometer, represented in terms of (**b**) "peaks" and (**c**) a bar graph. The relative areas of the peaks are 0.9092 (^{20}Ne), 0.00257 (^{21}Ne), and 0.0882 (^{22}Ne); natural neon is therefore 90.92% ^{20}Ne, 0.257% ^{21}Ne, and 8.82% ^{22}Ne.



Masterfile

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.

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)

EXERCISE!

Naturally occurring copper consist of ^{63}Cu(mass 62.9296 amu) and ^{65}Cu (mass 64.9278 amu) with an average atomic mass of 63.546 amu. What is the percentage of ^{63}Cu and ^{65}Cu isotopes?

- Mole is the measuring unit that represent the amount of chemical substances.

- Need a way to know quantity (amount) without counting atoms. Too many to count !… very tiny!
- 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

- Molar mass= Mass in grams of one mole of the substance (average atomic mass)
- 1 a.m.u=1g/mol

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)

Determine the number of mols of 10g of H_2O?

- 0.55 mols

Determine the mass of 3 mols of CO_2?

- 132g

Determine the number of oxygen atoms in 3 mols of oxygen molecule(O_2)

- 3.61×10^{24} oxygen atoms

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) 1.055×10^{-22} g

CONCEPT CHECK!

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

6.022×10^{23} Cu atoms

EXERCISE!

Compute the number of atoms in a 10.0g sample of Aluminum(Al).

$2.23X10^{23}$ Al atoms

EXERCISE!

Compute the number of atoms in a 10.0g sample of Aluminum(Al).

CONCEPT CHECK!

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

a) 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

EXERCISE!

- A sample of xenon fluoride has a structure of (XeF_n) where n is a whole number. If 8.06x10^{20} molecules of this compound have the mass of 0.227g. Calculate the value of n?

- The atomic composition of chemical compounds can be described using a variety of notations including molecular, empirical, and structural formulas. Another convenient way to describe atomic composition is to *examine* *the percent composition of a compound* by mass.

- H_2O has 11% H and 89% O (definite proportion from chap.2)
- % element in the compound= [molar mass of the element /molar mass of the molecule]X 100
- %H =2(1)/18X100=11%
- %O=1x(16)/18 X100= 89%

EXERCISE!

- For iron in iron(III) oxide, (Fe_2O_3): the % composition is:

  %Fe=2(55.85)/2(55.85)+3(16.00)=

  11.70/159.70 X100%

  = 69.94%

EXERCISE!

- Calculate the % of P and O in $Ba_3(PO_4)_2$

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

EXERCISE!

- Name the following compound $Na_3PO_4.10H_2O$ then determine % O in this compound.

EXERCISE!

- Substance A_3B has the composition by mass of 60% A and 40%B. What is the composition of A and B in AB_2 ?

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

EXERCISE!

Assume you have 100g of this compound S_xO_y given that mass of O=50% and mass of S=50%

- What is the empirical formula?

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:
- Rules to write a chemical equation:

1. Determine what reaction is occurring (combustion, precipitation…etc.
2. Knowing what are the Reactants (only placed on the left side of the arrow), and the products(only placed on the right side of the arrow).
3. Writ the un balanced chemical equation correctly in both sides.
4. Balance the chemical equation correctly by making sure that all atoms present in the reactants are accounted for in the products(using coefficient) .

- The balanced equation represents an overall ratio of reactants and products, not what actually “happens” during a reaction.
- It can be either mole ratio or particle ratio.

- 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.

Example:

A 1 mole or molecule of Ethanol(C_2H_5OH) reacts (burns) with 3 moles or molecules of Oxygen to produce 2 moles or molecules of carbon Dioxide and 3 moles or molecules of water.

What is the chemical equation of this reaction? Is it balanced or not?

$$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O \text{ (combustion)}$$

reactants products

- Yes balanced.

AP Learning Objectives, Margin Notes and References

- Learning Objectives
  - **LO 1.17:** The student is able to express the law of conservation of mass quantitatively and qualitatively using symbolic representations and particular drawings.
  - **LO 1.18:** The student is able to apply conservation of atoms to the rearrangements of atoms in various processes.

Writing and Balancing the Equation for a Chemical Reaction

1. 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.

2. The link bellow will show you how to balance a chemical equations.

https://www.youtube.com/watch?v=eNsVaUCzvLA

Balance the following chemical reactions:

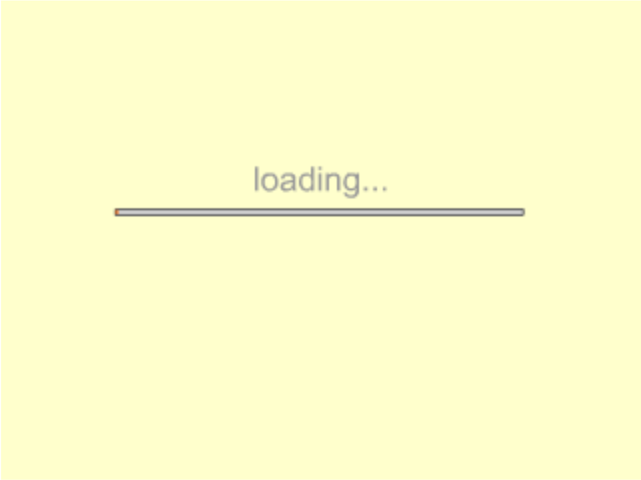
1. $Na_3PO_4 + BaCl_2 \rightarrow NaCl + Ba_3(PO_4)_2$
2. $Ba(OH)_2 + HCl \rightarrow BaCl_2 + H_2O$
3. $CH_4 + O_2 \rightarrow CO_2 + H_2O$

Balance the following chemical reactions:

1. $C_6H_6 + O_2 \qquad CO_2 + H_2O \rightarrow$

2. $C_6H_{14} + O_2 \qquad CO_2 + H_2O \rightarrow$

Section 3.9
Balancing Chemical Equations



To play movie you must be in Slide Show Mode
PC Users: Please wait for content to load, then click to play
Mac Users: **CLICK HERE**

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.

$CaO + C \rightarrow CaC_2 + CO_2$

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

II. $2CaO + 5C \rightarrow 2CaC_2 + CO_2$

III. $CaO + (2.5)C \rightarrow CaC_2 + (0.5)CO_2$

IV. $4CaO + 10C \rightarrow 4CaC_2 + 2CO_2$

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

- Calculation related to any balanced chemical equation.
- All ratio depends on the coefficient in the balanced equation.

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.

- Potassium chlorate decomposes to potassium chloride and oxygen gas.
  - Write the chemical equation for this reaction.
  - If 25g of potassium chlorate decomposes, what mass of potassium chloride and oxygen gas does it produce?

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

Given is $KClO_3$ mass =25g

Required is KCl and O_2 masses.

1- From mass to mole (m/M=n)

n of $KClO_3$= m/M=25/122.5=0.204 mol.

2- Mole ratio from equation.

a- 2mole of $KClO_3$ → 2 mole of KCl

0.204 of $KClO_3$ → X mole of KCl

X=0.204*2/2= 0.204 mol of KCl.

b- 2mole of $KClO_3$ → 3 mole of O_2

0.204 of $KClO_3$ → X mole of O_2

X=0.204*3/2= 0.306 mol of O_2

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

Given is $KClO_3$ mass =25g

Required is KCl and O_2 masses.

3- From mole to mass(n*M=m)

a- mass of KCl= n*M=0.204*74= 15g

b- mass of O_2= 0.306*32=9.8g

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!

Calculate the no. of Hydrogen atoms that will be produced from the combustion of $1.2x10^{24}$

Molecule of Ethane (C_2H_6)

EXERCISE!

- Calculate the mass of oxygen that is required to produce 0.5g of water according to the following equation?

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

EXERCISE!

- Calculate the mass of ammonia (NH3) that will be produce after a complete reaction of 2×10^{23} atoms of nitrogen with excess hydrogen according to the following equation?

$$N_2 + 3H_2 \rightarrow 2NH_3$$

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.

Limiting R

Stoichiometric amounts of sandwich ingredients for this recipe are bread and cheese slices in a 2:1 ratio. Provided with 28 slices of bread and 11 slices of cheese, one may prepare 11 sandwiches per the provided recipe, using all the provided cheese and having six slices of bread left over. In this scenario, the number of sandwiches prepared has been *limited* by the number of cheese slices, and the bread slices have been provided in *excess*.

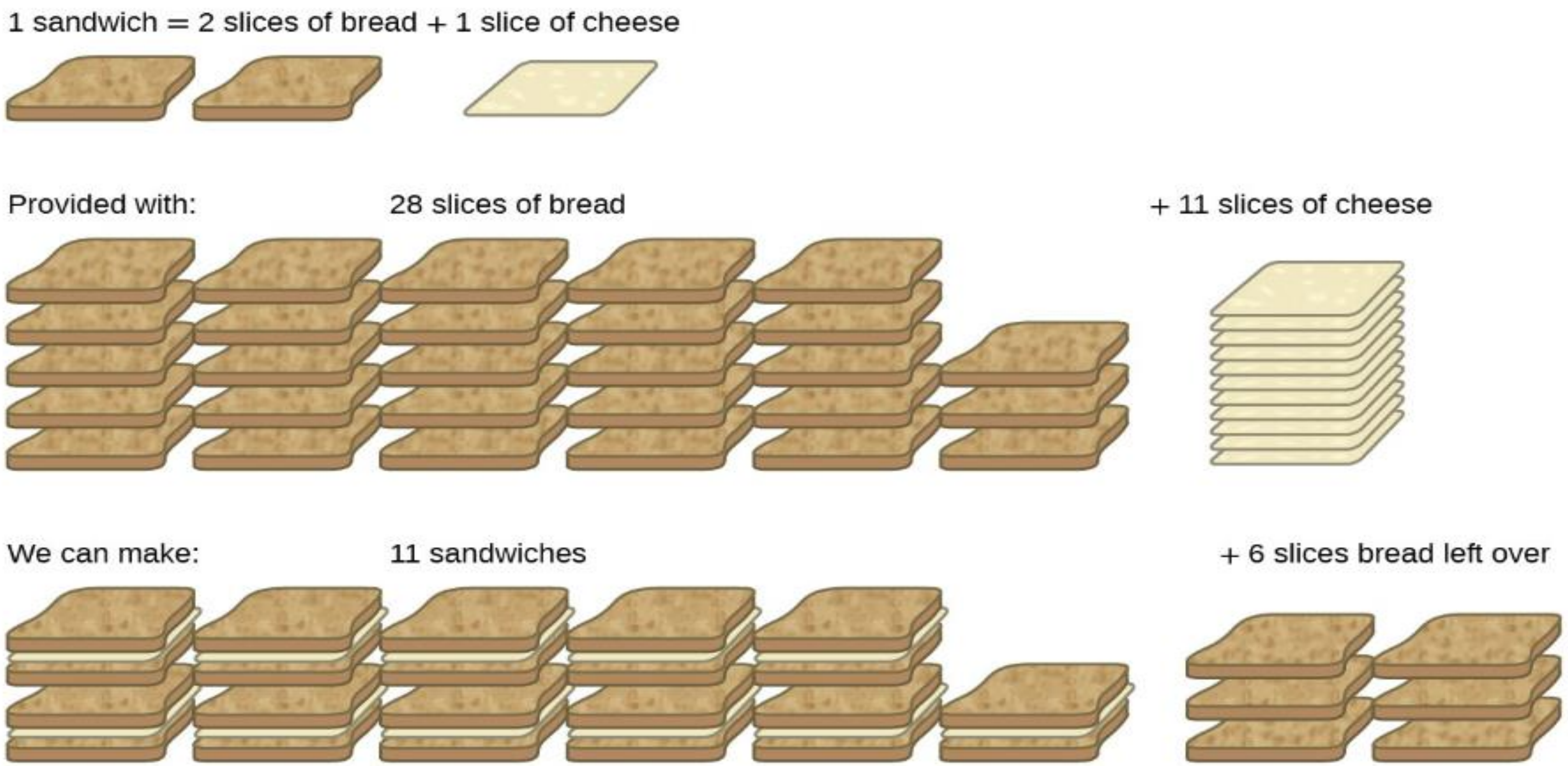
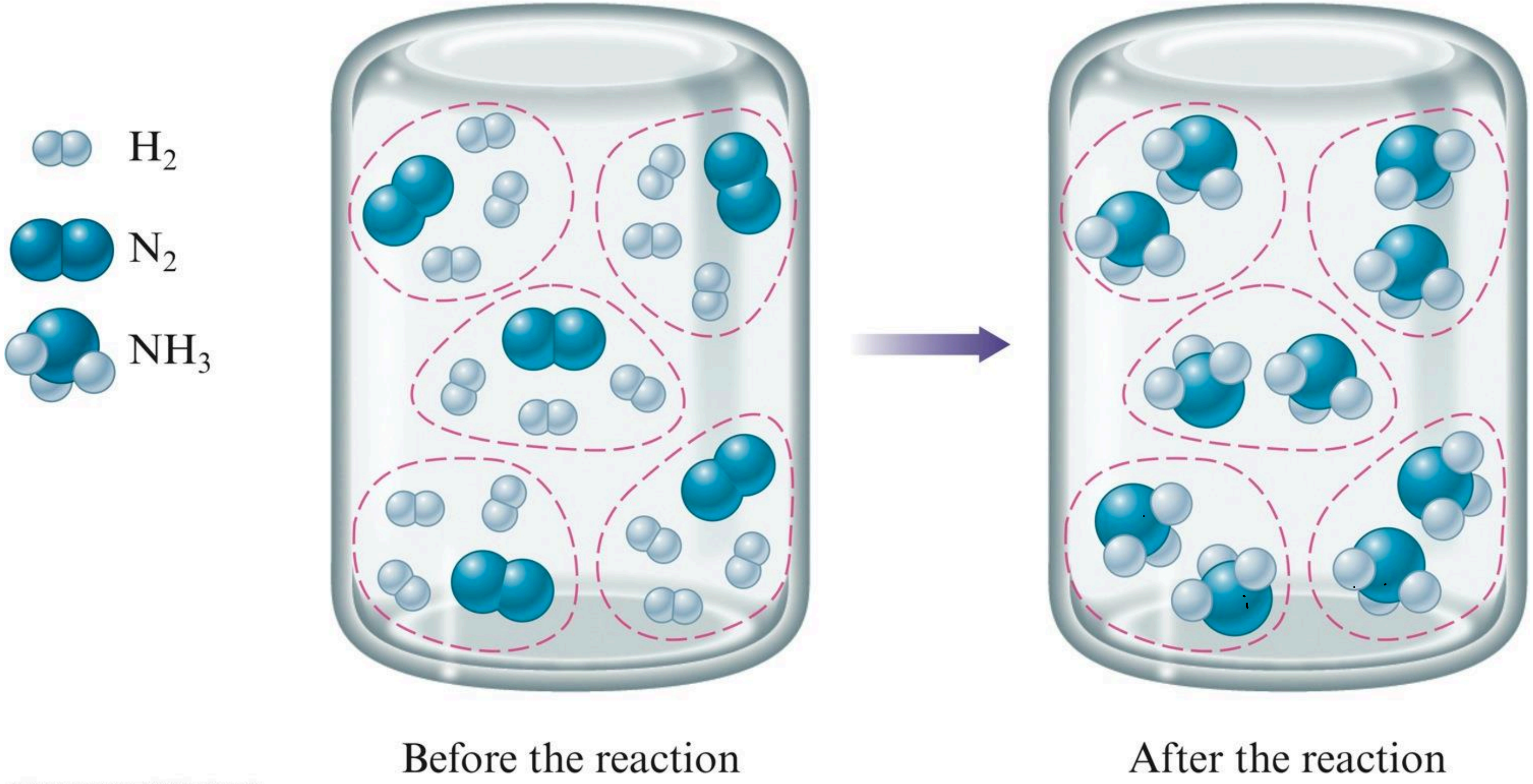


Figure 1. Sandwich making can illustrate the concepts of limiting and excess reactants.

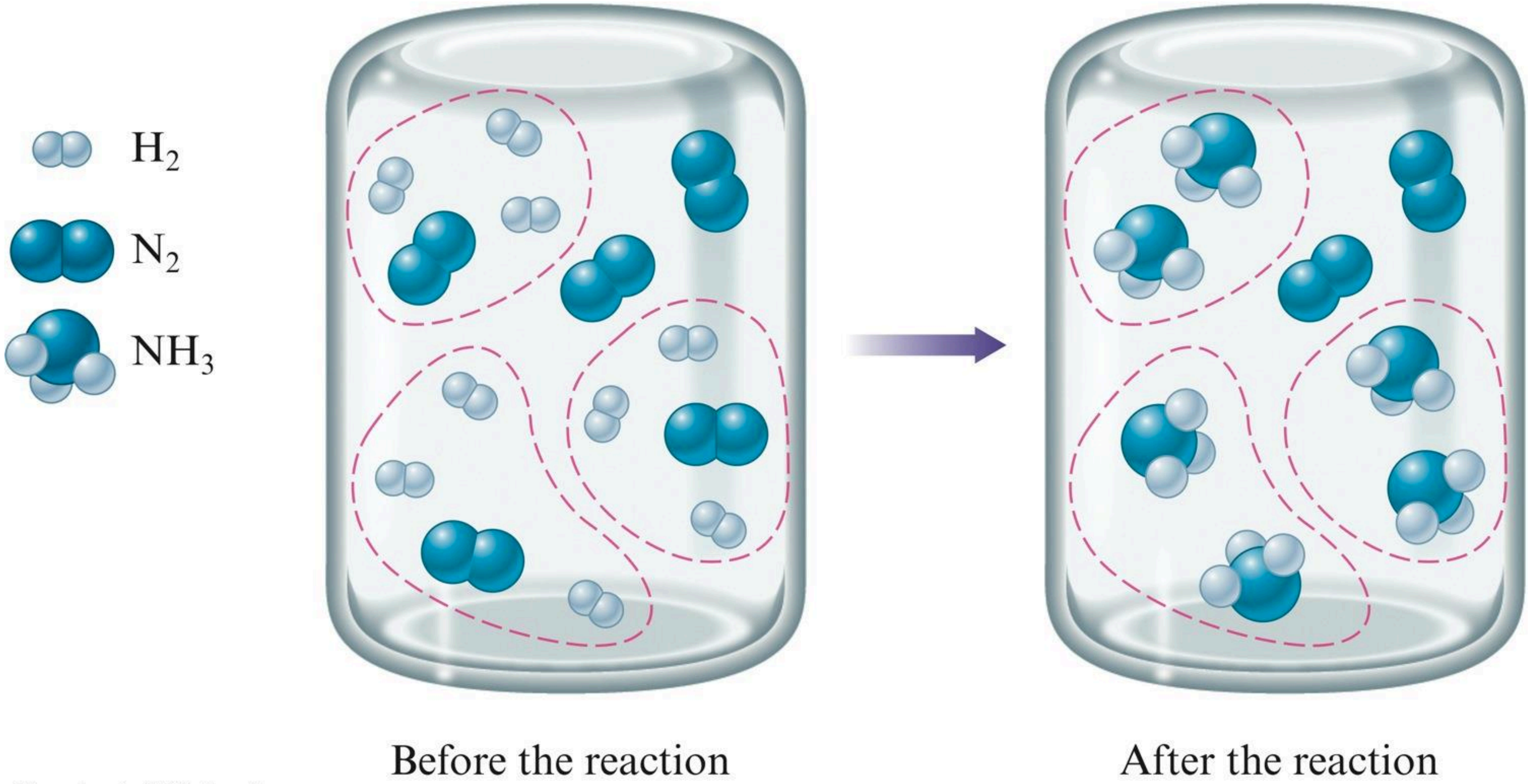
A. The Concept of Limiting Reactants

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



A. The Concept of Limiting Reactants

- Limiting reactant mixture



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

How to calculate Limiting Reactant

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

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

Calculate the mass of Al_2O_3 formed?

g Al ⟶ mol Al ⟶ mol Fe_2O_3 needed ⟶ g Fe_2O_3 needed

OR

g Fe_2O_3 ⟶ mol Fe_2O_3 ⟶ mol Al needed ⟶ g Al needed

$$124 \text{ g Al} \times \frac{1 \text{ mol Al}}{27.0 \text{ g Al}} \times \frac{1 \text{ mol } Fe_2O_3}{2 \text{ mol Al}} \times \frac{160. \text{ g } Fe_2O_3}{1 \text{ mol } Fe_2O_3} = \mathbf{367} \text{ g } Fe_2O_3$$

Start with 124 g Al ⟶ need 367 g Fe_2O_3

Have more Fe_2O_3 (601 g) so Al is limiting reagent

Continue

Use limiting reagent (Al) to calculate amount of product that can be formed.

$$\text{g Al} \longrightarrow \text{mol Al} \longrightarrow \text{mol } Al_2O_3 \longrightarrow \text{g } Al_2O_3$$

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

$$124 \text{ g Al} \times \frac{1 \text{ mol Al}}{27.0 \text{ g Al}} \times \frac{1 \text{ mol } Al_2O_3}{2 \text{ mol Al}} \times \frac{102. \text{ g } Al_2O_3}{1 \text{ mol } Al_2O_3} = 234 \text{ g } Al_2O_3$$

Example

- When 35.50 grams of nitrogen react with 25.75 grams of hydrogen, how many grams of ammonia are produced?

- How many grams of excess reagent remain in the reaction vessel?

$3H_{2(g)} + N_{2(g)} \rightarrow 2NH_{3(g)}$

Continue

1.27 mole of N_2 12.75 mole of H_2

1- Assume the N_2 is the limiting reactant

2- Calculate how many moles of H_2 needed

$$\text{Moles of } H_2 \text{ needed} = \text{moles of } N_2 \times \frac{3 \text{ Moles of } H_2}{1 \text{ Moles of } N_2}$$

$$1.27 \text{ mole } N_2 \times \frac{3 \text{ Moles of } H_2}{1 \text{ Moles of } N_2} = \mathbf{3.81 \text{ mole } H_2}$$

N_2 is the limiting reactant

Continue

$3H_{2(g)} + N_{2(g)} \rightarrow 2NH_{3(g)}$

g N_2 → moles N_2 → moles NH_3 → g NH_3

$$\text{Moles of } NH_3 = 1.27 \text{ mole } N_2 \times \frac{2 \text{ Moles of } NH_3}{1 \text{ Moles of } N_2} = 2.54 \text{ mole } NH_3$$

Mass of NH_3 = 2.54 mole × 17.03 g/mole = 43.25 g

Moles of H_2 unreacted = 12.75 – 3.81 = 8.94 moles

Mass of H_2 remains = 8.94 mole × 2.02 g/mole = **18.06 g H_2**

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.

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

Percent Yield

- *The theoretical yield is the maximum amount of product you would expect from a reaction based on the amount of limiting reagent. In practice, however, chemists don't always obtain the maximum yield for many reasons.*
- *When running a reaction in the lab, loss of product often occurs during purification or isolation steps. You might even decide it is worth losing 10% of your product during an extra purification step because it is more important to have extremely pure product—as opposed to having a larger amount of less pure product.*

Example

3.75 g of zinc (Zn) reacted with **excess** hydrochloric acid (HCl), **5.58** g of zinc chloride ($ZnCl_2$) were collected. What is the percent yield for this reaction?

$$Zn_{(s)} + 2\,HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

g of Zn $\rightarrow$ moles of Zn $\rightarrow$ moles of $ZnCl_2$ $\rightarrow$ g of $ZnCl_2$

$$3.75 \text{ g of Zn} \times \frac{\text{mole}}{65.38 \text{ g}} \times \frac{1 \text{ mole of } ZnCl_2}{1 \text{ mole of Zn}} \times \frac{136.3 \text{ g } ZnCl_2}{\text{mole}}$$

= 7.82 g of $ZnCl_2$ (Theoretical Yield)

Continue

Theoretical yield = 7.82 g $ZnCl_2$

Actual yield = 5.58 g $ZnCl_2$

$$\% \text{ yield} = \frac{5.58 \text{ g of } ZnCl_2}{7.82 \text{ g of } Zncl_2} \times 100 = 71.4 \%$$

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