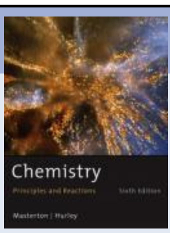

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Chapter 3
Mass Relations in Chemistry; Stoichiometry

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Outline

- Atomic Masses
- The Mole
- Mass Relations in Chemical Formulas
- Mass Relations in Reactions

Fundamental Questions

- How many atoms are there in a gram of an element?
- How much iron can be obtained from a ton of iron ore?
- How much nitrogen gas is required to form a kilogram of ammonia?

Stoichiometry

- Chemical arithmetic
- Study of mass relationships in chemistry

Atomic Masses

- Individual atoms are too light to be weighed on a balance
- Collections of atoms can be weighed and those collections can give accurate atomic masses

The Carbon-12 Scale

- The atomic mass of an element indicates how heavy, on average, an atom of an element is when compared to an atom of another element
- Unit is the atomic mass unit (amu)
- Mass of one ^{12}C atom = 12 amu (exactly)
- Note that ^{12}C and C-12 mean the same thing

Masses and the Periodic Table

- The mass of an element is indicated below the symbol for the element
- Consider He
  - On average, one He atom weighs 4.003 amu
  - This is about 1/3 the mass of a carbon-12 atom

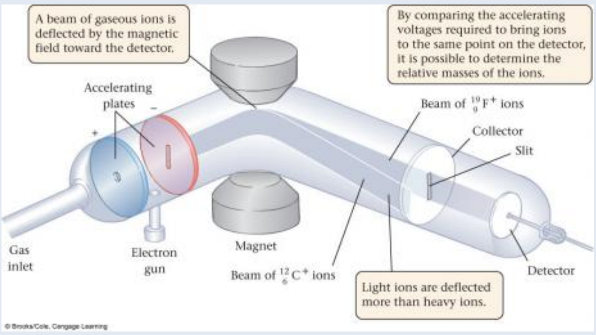
$$\frac{4.003 \text{ amu}}{12.00 \text{ amu}} = 0.3336$$

Comparing Two Atoms – Mass Ratio

$$\frac{\text{atomic mass X}}{\text{atomic mass Y}} = \frac{\text{mass of atom X}}{\text{mass of atom Y}}$$

Figure 3.1 – Mass Spectrometer

- A mass spectrometer is used to determine atomic masses



Mass Spectrometry

- Atoms are ionized at low pressure in the gas phase
- The cations that form are accelerated toward a magnetic field
- The extent to which the cation beam is deflected is inversely related to the mass of the cation

Fluorine

- Atomic fluorine exists as a single isotope
- Mass of F is exactly 19.00 amu

Carbon

- ^{12}C has a mass of exactly 12.000 amu
- Carbon in the periodic table has a mass of 12.011 amu
  - Why isn't it exactly 12.000?
  - Why are most atomic masses not whole numbers?

Isotopes

- Recall that an isotope is an atom with the same number of protons
  - Therefore, the same element
- Different mass number
  - Therefore, a different number of neutrons

Averages

- Take the simple average of two numbers
- Add the numbers and divide by 2
  - Each number counts 50%
  - (10 + 15) / 2 = 12.5
  - (0.5 X 10) + (0.5 X 15) = 12.5
- A simple average is simply a weighted average where each contributor counts 50%

Isotopic Abundance

- To determine the mass of an element, we must know the mass of each isotope and the atom percent of the isotopes (***isotopic abundance***)
- The mass spectrometer can determine the isotopic abundance

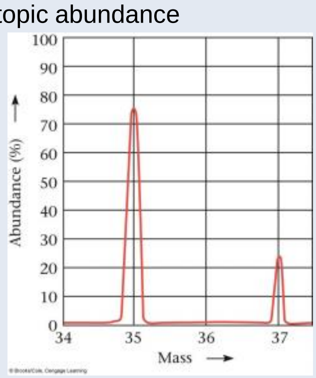
Chlorine

- Two isotopes
  - Cl-35
  - Cl-37

	Atomic Mass	Abundance
Cl-35	34.97 amu	75.53%
Cl-37	36.97 amu	24.47%

Figure 3.2 – Mass Spectrum of Cl

- The area under the peak in the mass spectrogram gives the isotopic abundance



Average Atomic Masses

- The atomic mass quoted in the periodic table is the weighted average of the atomic masses of all isotopes of that element

Weighted Averages

- For an element with two isotopes, Y_1 and Y_2:

$$\text{atomic mass } Y = (\text{atomic mass } Y_1)\frac{\%Y_1}{100} + (\text{atomic mass } Y_2)\frac{\%Y_2}{100}$$

- For chlorine:

$$\text{atomic mass Cl} = 34.97 \text{ amu}\left(\frac{75.53}{100}\right) + 36.97\left(\frac{35.46}{100}\right)\text{amu}$$

Example 3.1

Example 3.1 Bromine is a red-orange liquid with an average atomic mass of 79.90 amu. Its name is derived from the Greek word *bromos* ($\beta\rho o\mu o\varsigma$), which means stench. It has two naturally occurring isotopes: Br-79 (78.92 amu) and Br-81 (80.92 amu). What is the abundance of the heavier isotope?

Strategy The key to solving this problem is to realize that *the abundances of the isotopes have to add to 100%*. If we let x = abundance of Br-81, then the abundance of Br-79 must be $(100 - x)$.

SOLUTION Substituting in Equation 3.1,

$$79.90 \text{ amu} = 78.92 \text{ amu}\frac{(100 - x)}{100} + 80.92 \text{ amu}\frac{(x)}{100}$$

$$= 78.92 \text{ amu} + 2.00 \text{ amu}\frac{(x)}{100}$$

Solving: $\dfrac{x}{100} = \dfrac{(79.90 - 78.92)\text{ amu}}{2.00 \text{ amu}} = 0.49; \quad x = 100(0.49) = 49\%$

Reality Check The atomic mass of Br, 79.90, is just about halfway between the masses of the two isotopes, 78.92 and 80.92. So, it is reasonable that it should contain nearly equal amounts of the two isotopes.

Masses of Individual Atoms

- It is usually sufficient to know the relative masses of atoms
  - One He atom is about four times as heavy as one H atom
  - Therefore
    - The mass of 100 He atoms is about four times the mass of 100 H atoms
    - The mass of a million He atoms is about four times the mass of a million H atoms

Avogadro's Number

- There is a number that corresponds to a collection of atoms where the mass of that collection in grams is numerically equal to the same number in amu, for a single atom
- $N_A = 6.022 \times 10^{23}$
- Number of atoms of an element in a sample whose mass is numerically equal to the mass of a single atom
- By knowing Avogadro's number and the atomic mass, it is now possible to calculate the mass of a single atom in grams

Figure 3.3 – One Mole of Several Substances

Example 3.2

Example 3.2 Consider titanium (Ti), the "space-age" metal discussed at the end of Chapter 1. Taking Avogadro's number to be 6.022×10^{23}, calculate

(a) the mass of a titanium atom.
(b) the number of atoms in a ten-gram sample of the metal.
(c) the number of protons in 0.1500 lb of titanium.

Strategy From the periodic table, the atomic mass of titanium is 47.87 amu. It follows that

$$6.022 \times 10^{23} \text{ Ti atoms} = 47.87 \text{ g Ti}$$

The relation yields the required conversion factor for (a) and (b). Also from the periodic table, we see that the atomic number for titanium is 22, the number of protons in an atom of titanium.

Example 3.2 (cont'd)

SOLUTION

(a) $\text{mass of Ti atom} = 1\ \text{Ti atom} \times \dfrac{47.87\ \text{g Ti}}{6.022 \times 10^{23}\ \text{Ti atoms}} = \boxed{7.949 \times 10^{-23}\ \text{g}}$

(b) $\text{no. of Ti atoms} = 10.00\ \text{g} \times \dfrac{6.022 \times 10^{23}\ \text{Ti atoms}}{47.87\ \text{g Ti}} = \boxed{1.258 \times 10^{23}\ \text{atoms}}$

(c) First, change the given mass of Ti (0.1500 lb) to grams:

$$0.1500\ \text{lb} \times \frac{453.6\ \text{g}}{1\ \text{lb}} = 68.04\ \text{g}$$

$$\text{no. of protons} = 68.04\ \text{g} \times \frac{6.022 \times 10^{23}\ \text{Ti atoms}}{47.87\ \text{g Ti}} \times \frac{22\ \text{protons}}{1\ \text{atom}}$$

$$= \boxed{1.883 \times 10^{25}\ \text{protons}}$$

Reality Check Because atoms are so tiny, we expect their mass to be very small: 7.949×10^{-23} g sounds reasonable. Conversely, it takes a lot of atoms, in this case 1.258×10^{23}, to weigh ten grams.

The Mole

- A ***mole*** is Avogadro's number of items
- Very large number
  - Avogadro's number of pennies is enough to pay all the expenses of the United States for a billion years or more, without accounting for inflation
- The molar mass
  - The molar mass, MM, in grams/mole, is numerically equal to the sum of the masses (in amu) of the atoms in the formula

Molar Masses of Some Substances

Formula	Sum of Atomic Masses	Molar Mass, MM
O	16.00 amu	16.00 g/mol
O_2	2(16.00 amu) = 32.00 amu	32.00 g/mol
H_2O	2(1.008 amu) + 16.00 amu = 18.02 amu	18.02 g/mol
NaCl	22.99 amu + 35.45 amu = 58.44 amu	58.44 g/mol

The Significance of the Mole

- In the laboratory, substances are weighed on balances, in units of grams
- The mole allows us to relate the number of grams of a substance to the number of atoms or molecules of a substance

Mole-Gram Conversions

- m = MM X n
  - m = mass
  - MM = molar mass
  - n = number of moles

Example 3.3

Example 3.3 Acetylsalicylic acid, $C_9H_8O_4$, is the active ingredient of aspirin.

(a) What is the mass in grams of 0.509 moles of acetylsalicylic acid?
(b) A one-gram sample of aspirin contains 75.2% by mass of $C_9H_8O_4$. How many moles of acetylsalicylic acid are in the sample?
(c) How many molecules of $C_9H_8O_4$ are there in 12.00 g of acetylsalicylic acid? How many carbon atoms?

Strategy Find the molar mass of $C_9H_8O_4$ and use it to convert moles to grams in (a), grams to moles in (b), and (with Avogadro's number) grams to molecules in (c).

Example 3.3 (cont'd)

SOLUTION The molar mass of $C_9H_8O_4$ is

$$MM = [9(12.01) + 8(1.008) + 4(16.00)]\ g/mol = 180.15\ g/mol$$

(a) $0.509\ mol \times \frac{180.15\ g}{1\ mol} = \boxed{91.7\ g}$

(b) $1.00\ g\ aspirin \times \frac{75.2\ g\ C_9H_8O_4}{100\ g\ aspirin} \times \frac{1\ mol}{180.15\ g} = \boxed{4.17 \times 10^{-3}\ mol}$

(c) $12.00\ g\ C_9H_8O_4 \times \frac{1\ mol}{180.15\ g} \times \frac{6.022 \times 10^{23}\ molecules}{1\ mol} = \boxed{4.011 \times 10^{22}\ molecules}$

Since there are 9 carbon atoms in one molecule of $C_9H_8O_4$,

$$\text{no. of C atoms} = 4.011 \times 10^{22}\ molecules \times \frac{9\ C\ atoms}{1\ molecule} = \boxed{3.610 \times 10^{23}}$$

Reality Check Because every known substance has a molar mass greater than 1 g/mol, the mass of the sample in grams is always larger than the number of moles.

Chemical Formulas

- In chapter 2, we learned that the chemical formula tells us the number of atoms of each element in a compound, whether that is a molecular compound or an ionic compound
- We can now combine that knowledge with the knowledge of molar mass to begin relating elements by mass

Mass Relations in Chemical Formulas

- Percent composition from formula
  - The percent composition of a compound is stated as then number of grams of each element in 100 g of the compound
  - By knowing the formula, the mass percent of each element can be readily calculated

Example 3.4

Example 3.4 Graded

Metallic iron is most often extracted from hematite ore, which consists of iron(III) oxide mixed with impurities such as silicon dioxide, SiO_2.

*(a) What are the mass percents of iron and oxygen in iron(III) oxide?
**(b) How many grams of iron can be extracted from one kilogram of Fe_2O_3?
***(c) How many metric tons of hematite ore, 66.4% Fe_2O_3, must be processed to produce one kilogram of iron?

SOLUTION

(a) In one mole of Fe_2O_3, there are two moles of iron and three moles of oxygen.

$$\begin{aligned} \text{mass Fe} &= 2\ mol(55.85\ g/mol) = 111.7\ g \\ \text{mass O} &= 3\ mol(16.00\ g/mol) = 48.00\ g \\ & \qquad\qquad\qquad\qquad\quad 159.7\ g \end{aligned}$$

$$\%\ Fe = \frac{111.7\ g}{159.7\ g} \times 100\% = 69.94;\ \%\ O = 100.00 - 69.94 = \boxed{30.06}$$

(b) $1.000 \times 10^3\ g\ Fe_2O_3 \times \frac{111.7\ g\ Fe}{159.7\ g\ Fe_2O_3} = \boxed{699.4\ g\ Fe}$

(c) Perhaps the simplest way to proceed here is to start with one kilogram of Fe, convert that to kilograms of Fe_2O_3, then convert to kilograms of hematite and finally to metric tons.

$$\text{mass hematite} = 1.00\ kg\ Fe \times \frac{159.7\ g\ Fe_2O_3}{111.7\ g\ Fe} \times \frac{100.0\ g\ hematite}{66.4\ g\ Fe_2O_3} \times \frac{1\ metric\ ton}{10^3\ kg}$$

$$= \boxed{2.15 \times 10^{-3}\ metric\ ton}$$

Chemical Analysis

- Experimentation can give data that lead to the determination of the formula of a compound
  - Masses of elements in the compound
  - Mass percents of elements in the compound
  - Masses of products obtained from the reaction of a weighed sample of the compound

Simplest Formula from Chemical Analysis

- Often, the formula is not known, but data from chemical analysis is known
  - Amount of each element in grams
  - Can be used to determine the simplest formula
    - Smallest whole-number ratio of atoms in a compound
    - H_2O is the simplest formula for water
    - H_2O_2 is the molecular formula for HO

Example 3.5 – Simplest Formula from Masses of Elements

Example 3.5 A 25.00-g sample of an orange compound contains 6.64 g of potassium, 8.84 g of chromium, and 9.52 g of oxygen. Find the simplest formula.

Strategy First (1), convert the masses of the three elements to moles. Knowing the number of moles (n) of K, Cr, and O, you can then (2) calculate the mole ratios. Finally (3), equate the mole ratio to the atom ratio, which gives you the simplest formula.

SOLUTION

(1) $n_K = 6.64\text{ g K} \times \dfrac{1\text{ mol K}}{39.10\text{ g K}} = 0.170\text{ mol K}$

$n_{Cr} = 8.84\text{ g Cr} \times \dfrac{1\text{ mol Cr}}{52.00\text{ g Cr}} = 0.170\text{ mol Cr}$

$n_O = 9.52\text{ g O} \times \dfrac{1\text{ mol O}}{16.00\text{ g O}} = 0.595\text{ mol O}$

(2) To find the mole ratios, divide by the smallest number, 0.170 mol K:

$$\frac{0.170\text{ mol Cr}}{0.170\text{ mol K}} = \frac{1.00\text{ mol Cr}}{1\text{ mol K}} \qquad \frac{0.595\text{ mol O}}{0.170\text{ mol K}} = \frac{3.50\text{ mol O}}{1\text{ mol K}}$$

The mole ratio is 1 mol of K : 1 mol of Cr : 3.50 mol of O.

(3) As pointed out earlier, the mole ratio is the same as the atom ratio. To find the simplest whole-number atom ratio, multiply throughout by 2:

$$2\text{ K} : 2\text{ Cr} : 7\text{ O}$$

The simplest formula of the orange compound is $K_2Cr_2O_7$.

Reality Check A mole ratio of 1.00 A : 1.00 B : 3.33 C would imply a formula $A_3B_3C_{10}$; if the mole ratio were 1.00 A : 2.50 B : 5.50 C, the formula would be $A_2B_5C_{11}$. In general, multiply through by the smallest whole number that will give integers for all the subscripts.

Example 3.6 – Simplest Formula from Mass Percents

- When dealing with percentages, assume 100 g of the compound
- By doing so, the unitless percentage becomes a meaningful mass

Example 3.6 The compound that gives vinegar its sour taste is acetic acid, which contains the elements carbon, hydrogen, and oxygen. When 5.00 g of acetic acid are burned in air, 7.33 g of CO_2 and 3.00 g of water are obtained. What is the simplest formula of acetic acid?

Example 3.6 – Simplest Formula from Mass Percents (cont'd)

Strategy The first step (1) is to calculate the masses of C, H, and O in the 5.00-g sample. To do this note that all of the carbon has been converted to carbon dioxide. One mole of C (12.01 g) forms one mole of CO_2 (44.01 g). Hence to find the mass of carbon in 7.33 g of CO_2, you use the conversion factor:

$$12.01\text{ g C}/44.01\text{ g CO}_2$$

By the same token, all the hydrogen ends up as water. Because there are two moles of H (2.016 g) in one mole of H_2O (18.02 g), the conversion factor is

$$2(1.008)\text{g H}/18.02\text{ g H}_2\text{O}$$

The mass of oxygen in the sample cannot be found in a similar way because some of the oxygen in the products comes from the air required to burn the reactants. Instead we find the mass of oxygen by difference:

$$\text{mass O} = \text{mass of sample} - (\text{mass of C} + \text{mass of H})$$

Once the masses of the elements are obtained, it's all downhill; follow the same path as in Example 3.5.

Example 3.6 – Simplest Formula from Mass Percents (cont'd)

SOLUTION

(1) Find the mass of each element in the sample.

$$\text{mass C} = 7.33\text{ g CO}_2 \times \frac{12.01\text{ g C}}{44.01\text{ g CO}_2} = 2.00\text{ g C}$$

$$\text{mass H} = 3.00\text{ g H}_2\text{O} \times \frac{2.016\text{ g H}}{18.02\text{ g H}_2\text{O}} = 0.336\text{ g H}$$

$$\text{mass O} = 5.00\text{ g sample} - 2.00\text{ g C} - 0.336\text{ g H} = 2.66\text{ g O}$$

(2) Find the number of moles of each element.

$$n_C = 2.00\text{ g C} \times \frac{1\text{ mol C}}{12.01\text{ g C}} = 0.167\text{ mol C}$$

$$n_H = 0.336\text{ g H} \times \frac{1\text{ mol H}}{1.008\text{ g H}} = 0.333\text{ mol H}$$

$$n_O = 2.66\text{ g O} \times \frac{1\text{ mol O}}{16.00\text{ g O}} = 0.166\text{ mol O}$$

(3) Find the mole ratios and then the simplest formula.

$$\frac{0.167\text{ mol C}}{0.166\text{ mol O}} = \frac{1.01\text{ mol C}}{1.00\text{ mol O}} \qquad \frac{0.333\text{ mol H}}{0.166\text{ mol O}} = \frac{2.01\text{ mol H}}{1.00\text{ mol O}}$$

Rounding off to whole numbers, the mole ratios are

$$1\text{ mol C} : 2\text{ mol H} : 1\text{ mol O}$$

The simplest formula of acetic acid is CH_2O.

Molecular Formula from Simplest Formula

- The relationship between simplest and molecular formula is a whole number
- The whole number relates the molecular mass to the mass of the simplest formula as well

Example 3.7

Example 3.7 The molar mass of acetic acid, as determined with a mass spectrometer, is about 60 g/mol. Using that information along with the simplest formula found in Example 3.6, determine the molecular formula of acetic acid.

Strategy Calculate the molar mass corresponding to the simplest formula, CH_2O. Then find the multiple by dividing the actual molar mass, 60 g/mol, by the molar mass of CH_2O.

SOLUTION

$$\text{MM CH}_2\text{O} = 12.01\text{ g/mol} + 2(1.008\text{ g/mol}) + 16.00\text{ g/mol} = 30.03\text{ g/mol}$$

The ratio of the actual molar mass, 60 g/mol, to that of CH_2O is

$$\frac{60\text{ g/mol}}{30.03\text{ g/mol}} = 2; \qquad \text{molecular formula} = C_2H_4O_2$$

Mass Relations in Reactions

- Chemical equations represent chemical reactions
  - Reactants appear on the left
  - Products appear on the right
  - Equation must be balanced
    - Number of atoms of each element on the left …
    - …equals the number of atoms of each element on the right

How are Equations Written?

- We must know the reactants and the products for a reaction for which an equation is to be written
  - It is often necessary to do an experiment and an analysis to determine the products of a reaction
  - Determining the products is often time consuming and difficult

Writing Chemical Equations

1. Write a skeleton equation for the reaction.
2. Indicate the physical state of each reactant and product.
3. Balance the equation
   - Only the coefficients can be changed; subscripts are fixed by chemical nature of the reactants and products
   - It is best to balance atoms that appear only once on each side of the equation first

Example 3.8

Example 3.8 Crystals of sodium hydroxide (lye) react with carbon dioxide from air to form a colorless liquid, water, and a white powder, sodium carbonate, which is commonly added to detergents as a softening agent. Write a balanced equation for this chemical reaction.

Strategy To translate names into formulas, recall the discussion in Section 2.6, Chapter 2. The physical states are given or implied. To balance the equation, you could start with either sodium or hydrogen.

SOLUTION The skeleton equation is

$$NaOH(s) + CO_2(g) \longrightarrow Na_2CO_3(s) + H_2O(l)$$

Because there are two Na atoms on the right, a coefficient of 2 is written in front of NaOH:

$$2NaOH(s) + CO_2(g) \longrightarrow Na_2CO_3(s) + H_2O(l)$$

Careful inspection shows that all atoms are now balanced. There are two Na atoms, four O atoms, two H atoms, and one C atom on both sides of the equation.

Mass Relations from Equations

- The coefficients of a balanced equation represent the numbers of moles of reactants and products
  - $2\ N_2H_4\ (l) + N_2O_4\ (l) \rightarrow 3\ N_2\ (g) + 4\ H_2O\ (l)$
  - $2\ mol\ N_2H_4 + 1\ mol\ N_2O_4 \rightarrow 3\ mol\ N_2 + 4\ mol\ H_2O$

Example 3.9

Example 3.9 Graded

Ammonia is used to make fertilizers for lawns and gardens by reacting nitrogen gas with hydrogen gas.

* (a) Write a balanced equation with smallest whole-number coefficients for this reaction.
* (b) How many moles of ammonia are formed when 1.34 mol of nitrogen react?
** (c) How many grams of hydrogen are required to produce 2.75×10^3 g of ammonia?
** (d) How many molecules of ammonia are formed when 2.92 g of hydrogen react?
*** (e) How many grams of ammonia are produced when 15.0 L of air (79% by volume nitrogen) react with an excess of hydrogen? The density of nitrogen at the conditions of the reaction is 1.25 g/L.

SOLUTION

(a) The equation for the reaction is

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

Example 3.9 (cont'd)

(b) The reaction gives you the stoichiometric mole ratio: 2 moles NH_3/1 mole N_2. Thus,

$$n_{NH_3} = 1.34 \text{ mol } N_2 \times \frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2} = 2.68 \text{ mol } NH_3$$

(c) Here we follow the scheme

$$m_{NH_3} \xrightarrow{MM} n_{NH_3} \xrightarrow{ratio} n_{H_2} \xrightarrow{MM} m_{H_2}$$

$$\text{mass } H_2 = 2.75 \times 10^3 \text{ g } NH_3 \times \frac{1 \text{ mol } NH_3}{17.03 \text{ g } NH_3} \times \frac{3 \text{ mol } H_2}{2 \text{ mol } NH_3} \times \frac{2.016 \text{ g } H_2}{1 \text{ mol } H_2} = 488 \text{ g } H_2$$

(d) We add the Avogadro conversion factor N_A to the scheme:

$$m_{H_2} \xrightarrow{MM} n_{H_2} \xrightarrow{ratio} n_{NH_3} \xrightarrow{N_A} \text{molecules } NH_3$$

$$\text{molecules } NH_3 = 2.92 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2} \times \frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol } NH_3}$$

$$= 5.81 \times 10^{23} \text{ molecules}$$

(e) We first find the mass of nitrogen in air and then proceed using a scheme similar to those above.

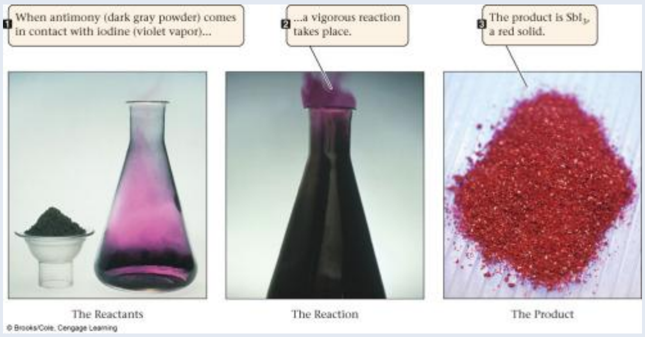
$$m_{N_2} \xrightarrow{MM} n_{N_2} \xrightarrow{ratio} n_{NH_3} \xrightarrow{MM} m_{NH_3}$$

$$\text{mass } N_2 = 15.0 \text{ L air} \times \frac{79 \text{ L } N_2}{100 \text{ L air}} \times \frac{1.25 \text{ g } N_2}{1 \text{ L } N_2} = 14.8 \text{ g } N_2$$

$$m_{NH_3} = 14.8 \text{ g } N_2 \times \frac{1 \text{ mol } N_2}{28.02 \text{ g } N_2} \times \frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2} \times \frac{17.03 \text{ g } NH_3}{1 \text{ mol } NH_3} = 18 \text{ g } NH_3$$

Limiting Reactant and Theoretical Yield

- 2 Sb (s) + 3 I_2 (s) → 2 SbI_3 (s)



Interpreting by Mass

- Reactants
  - One mole Sb (243.6 g)
  - Three moles I_2 (761.4 g)
  - Two moles SbI_3 (1005.0 g)
- All of the reactants are converted to product

In the Laboratory

- Reactants are usually not mixed in exact ratios
- An excess of one reactant is often used
  - Usually the less (or least) expensive reactant
  - One reactant will then limit the amount of product that will form

Sb-I_2 with a limiting reactant

- Suppose the mixture is
  - 3.00 mol Sb
  - 3.00 mol I_2
- In this case
  - 1.00 mol Sb will be left over
  - 2.00 mol of Sb will be used
    - React with 3.00 mol I_2
    - Form 2.00 mol SbI_3

Approach to Limiting Reactant Problems

1. Calculate the amount of product that will form if the first reactant were completely consumed.
2. Repeat the calculation for the second reactant in the same way.
3. Choose the smaller amount of product and relate it to the reactant that produced it. This is the limiting reactant and the resulting amount of product is the ***theoretical yield***.
4. From the theoretical yield, determine how much of the reactant in excess is used, and subtract from the starting amount.

Example 3.10

Example 3.10 Consider the reaction

$$2Sb(s) + 3I_2(s) \longrightarrow 2SbI_3(s)$$

Determine the limiting reactant and the theoretical yield when

(a) 1.20 mol of Sb and 2.40 mol of I_2 are mixed.
(b) 1.20 g of Sb and 2.40 g of I_2 are mixed. What mass of excess reactant is left when the reaction is complete?

Strategy Follow the four steps outlined above. In steps (1) and (2), follow the strategy described in Example 3.9 to calculate the amount of product formed. In (a), a simple one-step conversion is required; in (b), the path is longer because you have to go from mass of reactant to mass of product. To find the mass of reactant left over in (b), calculate how much is required to give the theoretical yield of product. Subtract that from the starting amount to find the amount left.

Example 3.10

SOLUTION

(a) (1) n_{SbI_3} from Sb $= 1.20 \text{ mol Sb} \times \dfrac{2 \text{ mol } SbI_3}{2 \text{ mol Sb}} = 1.20 \text{ mol } SbI_3$

(2) n_{SbI_3} from $I_2 = 2.40 \text{ mol } I_2 \times \dfrac{2 \text{ mol } SbI_3}{3 \text{ mol } I_2} = 1.60 \text{ mol } SbI_3$

(3) Because 1.20 mol is the smaller amount of product, that is the theoretical yield of SbI_3. This amount of SbI_3 is produced by the antimony, so Sb is the limiting reactant.

(b) (1) mass of SbI_3 from Sb $= 1.20 \text{ g Sb} \times \dfrac{1 \text{ mol Sb}}{121.8 \text{ g Sb}} \times \dfrac{2 \text{ mol } SbI_3}{2 \text{ mol Sb}} \times \dfrac{502.5 \text{ g } SbI_3}{1 \text{ mol } SbI_3} = 4.95 \text{ g } SbI_3$

(2) mass of SbI_3 from $I_2 = 2.40 \text{ g } I_2 \times \dfrac{1 \text{ mol } I_2}{253.8 \text{ g } I_2} \times \dfrac{2 \text{ mol } SbI_3}{3 \text{ mol } I_2} \times \dfrac{502.5 \text{ g } SbI_3}{1 \text{ mol } SbI_3} = 3.17 \text{ g } SbI_3$

(3) The reactant that yields the smaller amount (3.17 g) of SbI_3 is I_2. Hence I_2 is the limiting reactant. The smaller amount, 3.17 g of SbI_3, is the theoretical yield.

(4) mass of Sb required $= 3.17 \text{ g } SbI_3 \times \dfrac{121.8 \text{ g Sb}}{502.5 \text{ g } SbI_3} = 0.768 \text{ g Sb}$

mass of Sb left $= 1.20 \text{ g} - 0.768 \text{ g} = 0.43 \text{ g}$

Verifying the Limiting Reactant

- Once the limiting reactant has completely reacted, there is no more left to react with the excess of the other reactant

The Pancake Analogy

- Consider a recipe for pancakes. To make 16 pancakes, you need
  - 2 cups flour
  - 2 teaspoons baking powder
  - 2 eggs
  - 1 cup milk

Pancakes

- Now start with
  - 2 cups flour
  - 2 teaspoons baking powder
  - 1 egg
  - 1 cup milk
- It is clear that the egg will limit you to 8 pancakes and that you'll have a cup of flour, a teaspoon of baking powder and a half a cup of milk left over

Pancakes

- The egg is the limiting reactant and the theoretical yield is 8 pancakes

Experimental Yield

- Experimental yields are always lower than theoretical yields
  - Some product is lost to competing reactions
  - Some product is lost to handling
  - Some product may be lost in separating it from the reaction mixture
- The ***actual yield*** is the quantity of product you measure after you have done the reaction in the laboratory

Percent Yield

- The percent yield is defined as

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

Example 3.11

Example 3.11 Consider again the reaction discussed in Example 3.10:

$$2Sb(s) + 3I_2(s) \longrightarrow 2SbI_3(s)$$

Suppose that in part (a) the percent yield is 78.2%. How many grams of SbI_3 are formed?

Strategy Use Equation 3.3 to find the number of moles of SbI_3 and then the mass in grams (1 mol SbI_3 = 502.5 g SbI_3).

SOLUTION

$$\text{experimental yield } SbI_3 = 1.20 \text{ mol } SbI_3 \times \frac{78.2}{100} \times \frac{502.5 \text{ g } SbI_3}{1 \text{ mol } SbI_3} = 472 \text{ g } SbI_3$$

Key Concepts

1. Relate the atomic mass of an element to isotopic masses and abundances.
2. Use Avogadro's number to calculate the mass of an atom or molecule (in grams).
3. Use molar mass to relate
   - Moles to mass of a substance
   - Molecular formula to simplest formula
4. Use the formula of a compound to find its percent composition or the equivalent

Key Concepts

5. Find the simplest formula of a compound from chemical analysis
6. Balance chemical equations by inspection
7. Use a balanced equation to
   - Relate masses of reactants and products
   - Find the limiting reactant, theoretical yield, and percent yield of a reaction.