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Chapter 3
Mass Relations in Chemistry: Stoichiometry

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Outline

- The Mole
  - The mole and solutions: Molarity
- Mass Relations in Chemical Formulas
- Mass Relations in Reactions

Counting

- Different specialties use different counting numbers
  - Doughnuts and eggs are sold by the dozen
  - Pennies are wrapped in rolls of 50
  - ATMs dispense money in units of $20 but Congress spends by the million

Avogadro's Number: 1 mole = 6.022×10^{23} Particles

- 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

The Mole

- A ***mole*** is Avogadro's number of items
- The mole is a 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

Significance of the Mole

- By knowing Avogadro's number and the atomic mass, it is now possible to calculate the mass of a single atom in grams

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

Example 3.1

EXAMPLE 3.1

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 (ASA)?

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?

continued

Example 3.1, (Cont'd)

a

ANALYSIS	
Information given:	moles of acetylsalicylic acid (0.509) formula for acetylsalicylic acid ($C_9H_8O_4$)
Information implied:	molar mass (MM) of acetylsalicylic acid
Asked for:	mass of acetylsalicylic acid (ASA)
STRATEGY	
Substitute into Equation 3.1 mass = MM × n	
SOLUTION	
MM of $C_9H_8O_4$	9(12.01) + 8(1.008) + 4(16.00) = 180.15 g/mol
mass	mass = MM × n = 0.509 mol × $\frac{180.15\ \text{g}}{1\ \text{mol}}$ = 91.7 g

Example 3.1, (Cont'd)

b

ANALYSIS	
Information given:	mass of aspirin (1.000 g) mass percent of ASA in aspirin (75.2%) formula for acetylsalicylic acid ($C_9H_8O_4$)
Information implied:	molar mass (MM) of acetylsalicylic acid
Asked for:	moles of ASA in the sample of acetylsalicylic acid
STRATEGY	
Follow the plan outlined in Figure 3.2. aspirin $\xrightarrow{\%}$ mass $\xrightarrow{\text{MM}}$ moles	
SOLUTION	
moles ASA	1.00 g aspirin × $\frac{75.2\ \text{g ASA}}{100\ \text{g aspirin}}$ × $\frac{1\ \text{mol ASA}}{180.15\ \text{g ASA}}$ = 4.17×10^{-3} mol ASA

c

ANALYSIS	
Information given:	mass of ASA (12.00 g) formula for acetylsalicylic acid ($C_9H_8O_4$)
Information implied:	molar mass (MM) of acetylsalicylic acid Avogadro's number (N_A)
Asked for:	number of molecules of ASA number of carbon atoms

continued

Example 3.1, (Cont'd)

STRATEGY	
Follow the plan outlined in Figure 3.2. mass $\xrightarrow{\text{MM}}$ moles $\xrightarrow{N_A}$ molecules $\xrightarrow{9\text{ C atoms}}$ C atoms	
SOLUTION	
number of molecules	12.00 g ASA × $\frac{6.022 \times 10^{23}\ \text{molecules ASA}}{180.15\ \text{g ASA}}$ = 4.011×10^{22}
number of C atoms	4.011×10^{22} molecules ASA × $\frac{9\ \text{C atoms}}{1\ \text{molecule}}$ = 3.610×10^{23}
END POINT	
The mass of the sample in grams is always larger than the number of moles because every known substance has a molar mass greater than 1 g/mol.	

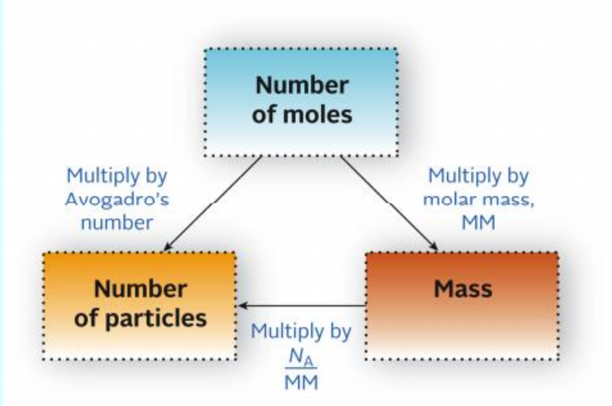
The Significance of the Mole

- 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

Figure 3.2 – Schema for Working With Moles and Masses



Reactions in the Laboratory

- Because water is common everywhere, most chemical reactions take place in aqueous solution
  - Water is called the universal solvent
  - Three common types of reactions in solution:
    - Precipitation reactions
    - Acid-base reactions
    - Oxidation-reduction reactions
- In Chapter 4, we will examine these reactions in detail – for now, we will look at the concentration of solutions in terms of their *molarity*

Solute Concentrations - Molarity

- Definition of molarity
  - Molarity = moles of solute/liters of solution
  - Symbol is M
  - Square brackets are used to indicate concentration in M
    - $[Na^+]$ = 1.0 M
- Consider a solution prepared from 1.20 mol of substance A, diluted to a total volume of 2.50 L
  - Concentration is 1.20 mol/2.50 L or 0.480 M

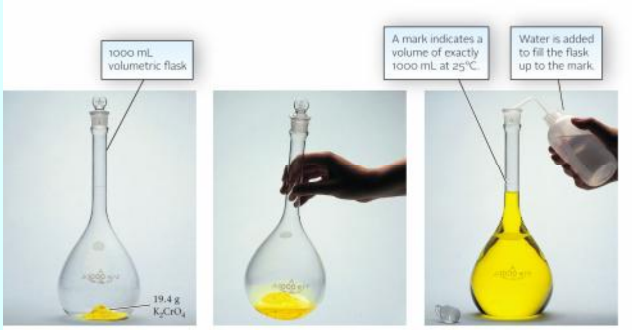
Additivity

- Masses are additive; volumes are not
- The ***total mass*** of a solution is the ***sum of the mass of the solute and the solvent***
- The ***total volume*** of a solution is ***not the sum of the volumes of the solute and solvent***

Volumetric Glassware

- Volumetric pipets, burets and flasks are made so that they contain a known volume of liquid at a given temperature
- Preparing solutions with concentrations in M involves using volumetric glassware

Figure 3.3 – Preparation of Molar Solution



Molarity as a Conversion Factor

- The molarity can be used to calculate
  - The number of moles of solute in a given volume of solution
  - The volume of solution containing a given number of moles of solute

Example 3.2

EXAMPLE 3.2

Nitric acid, HNO_3, is extensively used in the manufacture of fertilizer. A bottle containing 75.0 mL of nitric acid solution is labeled 6.0 *M* HNO_3.

(a) How many moles of HNO_3 are in the bottle?

(b) A reaction needs 5.00 g of HNO_3. How many mL of solution are required?

(c) Ten mL of water are added to the solution. What is the molarity of the resulting solution? (Assume volumes are additive.)

(a)

ANALYSIS

Information given:	*V* (75.0 mL) and *M* (6.0 *M*) of HNO_3 in the bottle
Information implied:	molar mass (MM) of HNO_3
Asked for:	moles of HNO_3 in the bottle

STRATEGY

1. Do not forget to change the volume unit given (mL) to L.
2. Use the molarity of HNO_3 as a conversion factor:

$$\frac{6.0 \text{ mol } HNO_3}{1 \text{ L solution}}$$

continued

Example 3.2, (Cont'd)

SOLUTION

moles HNO_3	$75.0 \text{ mL solution} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{6.0 \text{ mol } HNO_3}{1 \text{ L solution}} = 0.45 \text{ mol}$

(b)

ANALYSIS

Information given:	*V* (75.0 mL) and *M* (6.0 *M*) of HNO_3 in the bottle mass of HNO_3 required (5.00 g)
Information implied:	molar mass (MM) of HNO_3
Asked for:	volume of HNO_3 required

STRATEGY

1. Use the molarity and the molar mass of HNO_3 as conversion factors.

$$\frac{6.0 \text{ mol } HNO_3}{1 \text{ L solution}} \qquad \frac{1 \text{ mol } HNO_3}{63.02 \text{ g } HNO_3}$$

2. Follow the plan:

$$\text{mass} \xrightarrow{MM} \text{moles} \xrightarrow{M} \text{volume}$$

SOLUTION

volume HNO_3	$5.00 \text{ g } HNO_3 \times \frac{1 \text{ mol } HNO_3}{63.02 \text{ g } HNO_3} \times \frac{1 \text{ L solution}}{6.0 \text{ mol } HNO_3} = 0.013 \text{ L} = 13 \text{ mL}$

Example 3.2, (Cont'd)

(c)

ANALYSIS

Information given:	*V* (75.0 mL) and *M* (6.0 *M*) of HNO_3 in the bottle volume of water added (10.00 mL) Assume volumes are additive.
Information implied:	molar mass (MM) of HNO_3
Asked for:	molarity (*M*) of diluted solution

STRATEGY

1. Adding water does not change the number of moles of solute. It does change the volume of solution.
2. Substitute into Equation 3.2. Use (75.0 mL + 10.0 mL) as the total volume.

SOLUTION

M	$\frac{0.45 \text{ mol}}{0.0750 \text{ L} + 0.0100/\text{L}} = 5.3 \text{ mol/L} = 5.3\ M$

END POINT

The molarity of a solution decreases when water is added to the solution, but the moles of solute in solution remain the same.

Dissolving Ionic Solids

- When an ionic solid is dissolved in a solvent, the ions separate from each other
  - $MgCl_2$ (s) → Mg^{2+} (aq) + 2 Cl^- (aq)
- The concentrations of ions are related to each other by the formula of the compound:
  - Molarity $MgCl_2$ of = molarity of Mg^{2+}
  - Molarity of Cl^- = 2 X molarity of $MgCl_2$
  - Total number of moles of ions per mole of $MgCl_2$ is 3

Determining Moles of Ions

- By knowing the charge on ions, the formula of a compound can be quickly determined
- The formula of the compound is key to determining the concentration of ions in solution

3	4	5	6	7	8	9	10	11	12
			Cr^{3+}	Mn^{2+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+}	Cu^{2+}	Zn^{2+}
								Ag^{+}	Cd^{2+}
									Hg^{2+}

Example 3.3

EXAMPLE 3.3

Potassium dichromate, $K_2Cr_2O_7$, is used in the tanning of leather. A flask containing 125 mL of solution is labeled 0.145 *M* $K_2Cr_2O_7$.

(a) What is the molarity of each ion in solution?

(b) A sample containing 0.200 moles of K^+ is added to the solution. Assuming no volume change, what is the molarity of the new solution?

(a)

ANALYSIS

Information given:	volume, *V* (125 mL), and molarity, *M* (0.145 *M*), of solution
Information implied:	number of each ion in the parent compound
Asked for:	molarity of each ion in solution

continued

Example 3.3, (Cont'd)

STRATEGY

1. Distinguish between the parent compound and the ions.
2. Count the number of ions in the parent compound (recall that it is the subscript of the ion), and use it as a conversion factor.

$$\frac{2 \text{ mol } K^+}{1 \text{ mol } K_2Cr_2O_7} \qquad \frac{1 \text{ mol } Cr_2O_7^{2-}}{1 \text{ mol } K_2Cr_2O_7}$$

SOLUTION

$[K^+]$	$\frac{0.145 \text{ mol } K_2Cr_2O_7}{1 \text{ L}} \times \frac{2 \text{ mol } K^+}{1 \text{ mol } K_2Cr_2O_7} = 0.290 \text{ mol/L} = 0.290\ M$
$[Cr_2O_7^{2-}]$	$\frac{0.145 \text{ mol } K_2Cr_2O_7}{1 \text{ L}} \times \frac{1 \text{ mol } Cr_2O_7^{2-}}{1 \text{ mol } K_2Cr_2O_7} = 0.145 \text{ mol/L} = 0.145\ M$

(b)

ANALYSIS

Information given:	volume, *V* (125 mL), and molarity, *M* (0.145 *M*), of solution moles of K^+ added (0.200)
Information implied:	number of each ion in the parent compound
Asked for:	molarity of $K_2Cr_2O_7$ after the addition of K^+ ions

Example 3.3, (Cont'd)

STRATEGY

1. Use the molarity of K^+ found in (a) and substitute into Equation 3.2 to find the moles of K^+ initially.
2. Find the moles of K^+ after the addition.
3. Find $[K^+]$ again by substituting into Equation 3.2.
4. Find $[K_2Cr_2O_7]$ by relating $[K^+]$ to $K_2Cr_2O_7$.
5. The overall plan is:

$$[K^+_{\text{initial}}] \xrightarrow{V} (\text{mol } K^+)_{\text{initial}} \xrightarrow{+0.200 \text{ mol } K^+} (\text{mol } K^+)_{\text{final}} \xrightarrow{V} [K^+]_{\text{final}} \xrightarrow{2\,K^+/K_2Cr_2O_7} [K_2Cr_2O_7]_{\text{final}}$$

SOLUTION

$(\text{mol } K^+)_{\text{initial}}$	$0.125 \text{ L} \times \frac{0.290 \text{ mol } K^+}{1 \text{ L}} = 0.03625$
$(\text{mol } K^+)_{\text{final}}$	$0.03625 + 0.200 = 0.236$
$[K^+]_{\text{final}}$	$\frac{0.236 \text{ mol } K^+}{0.125 \text{ L}} = 1.89 \text{ mol/L} = 1.89\ M$
$[K_2Cr_2O_7]_{\text{final}}$	$\frac{1.89 \text{ mol } K^+}{1 \text{ L}} \times \frac{1 \text{ mol } K_2Cr_2O_7}{2 \text{ mol } K^+} = 0.945 \text{ mol/L} = 0.945\ M$

END POINT

The concentration of K^+ should be twice that of $Cr_2O_7^{2-}$ in either solution. It is!

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?

a

ANALYSIS

Information given:	formula of the iron oxide (Fe_2O_3)
Information implied:	molar mass (MM) of Fe_2O_3
Asked for:	mass % of Fe and O in Fe_2O_3

STRATEGY

1. To find the mass percent of Fe follow the plan outlined below. Start with one mole of Fe_2O_3.

$$n_{Fe_2O_3} \xrightarrow{\text{subscript}} n_{Fe} \xrightarrow{\text{MM Fe}} \text{mass Fe} \xrightarrow{\text{MM } Fe_2O_3} \text{\% element}$$

2. Find the mass percent of oxygen by difference. Note that the compound is made up of only two elements, Fe and O.

SOLUTION

mass % Fe	$1\ \text{mol } Fe_2O_3 \times \frac{2\ \text{mol Fe}}{1\ \text{mol } Fe_2O_3} \times \frac{55.85\ \text{g}}{1\ \text{mol Fe}} = 111.7\ \text{g Fe}$ $\frac{111.7\ \text{g Fe}}{159.7\ \text{g } Fe_2O_3} = 69.94\%$
mass % O	mass % O = 100% − mass % Fe = 100.00% − 69.94% = 30.06%

continued

Example 3.4, (Cont'd)

b

ANALYSIS

Information given:	mass of Fe_2O_3 (1.000 kg = 1.000×10^3 g)
Information implied:	from (a): mass % of Fe in Fe_2O_3 (69.94% = 69.94 g Fe/100.00 g Fe_2O_3)
Asked for:	mass of Fe obtained from 1.000 kg of Fe_2O_3

STRATEGY

mass of Fe = (mass Fe_2O_3)(mass Fe/100 g Fe_2O_3)

SOLUTION

mass Fe	$1.000 \times 10^3\ \text{g } Fe_2O_3 \times \frac{69.94\ \text{g Fe}}{100\ \text{g } Fe_2O_3} = 699.4\ \text{g Fe}$

c

ANALYSIS

Information given:	mass % of Fe_2O_3 in the ore (66.4% = 66.4 g Fe_2O_3/100.0 g ore) mass of Fe needed (1.000 kg)
Information implied:	from (a): mass % of Fe in Fe_2O_3 (69.94% = 69.94 g Fe/100.00 g Fe_2O_3) factor for converting metric tons to grams
Asked for:	mass of hematite needed to produce 1.00 kg of Fe

STRATEGY

1. Distinguish between the mass % of Fe_2O_3 in the ore (66.4%) and the mass % of Fe in Fe_2O_3 (69.94%).
2. Start with 1000 g of Fe and use the mass percents as conversion factors to go from mass of Fe to mass of the ore.

$$\frac{69.94\ \text{g Fe}}{100.00\ \text{g } Fe_2O_3} \qquad \frac{66.4\ \text{g } Fe_2O_3}{100.00\ \text{g ore}}$$

3. Convert grams to metric tons.

1 metric ton = 1×10^6 g

SOLUTION

mass of hematite needed	$1000\ \text{g Fe} \times \frac{100\ \text{g } Fe_2O_3}{69.94\ \text{g Fe}} \times \frac{100\ \text{g ore}}{66.4\ \text{g } Fe_2O_3} \times \frac{1\ \text{metric ton}}{1 \times 10^6\ \text{g}} = 2.15 \times 10^{-3}\ \text{metric tons}$

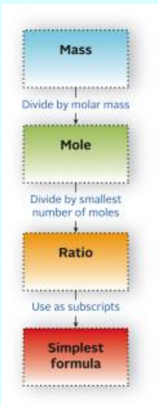
Chemical Analysis

- 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 and the molecular formula for water
    - HO is the simplest formula for hydrogen peroxide; the molecular formula is H_2O_2

Figure 3.5: Flowchart for Determining Simplest Formula



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.

ANALYSIS

Information given:	mass of sample (25.00 g) mass of each element in the compound: K (6.64 g); Cr (8.84 g); O (9.52 g)
Information implied:	atomic masses of the elements
Asked for:	simplest formula for the compound.

STRATEGY

Follow the schematic in Figure 3.5.

SOLUTION

moles of each element	K: $\frac{6.64\ \text{g K}}{39.10\ \text{g/mol}} = 0.170\ \text{mol}$ Cr: $\frac{8.84\ \text{g Cr}}{52.00\ \text{g/mol}} = 0.170\ \text{mol}$

continued

Example 3.5, (Cont'd)

ratios	O: $\frac{9.52\ g\ O}{16.00\ g/mol} = 0.595\ mol$ K: $\frac{0.170\ mol}{0.170\ mol} = 1$; Cr: $\frac{0.170\ mol}{0.170\ mol} = 1$; O: $\frac{0.595\ mol}{0.170}mol = 3.5$ Since we need smallest *whole* number ratios, we multiply all ratios by 2. 2 K : 2 Cr : 7 O
simplest formula	$K_2Cr_2O_7$

END POINT

1. 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.
2. Note that in this case, the mass of the sample was not used to find the simplest formula.

Example 3.6

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?

ANALYSIS

Information given:	elements in acetic acid (C, H, and O) mass of acetic acid (5.00 g) result of combustion analysis (7.33 g CO_2, 3.00 g H_2O)
Information implied:	molar masses (MM) of CO_2 and H_2O
Asked for:	simplest formula of acetic acid

STRATEGY

1. Find the mass of C and H by using the conversion factors:

$$\frac{12.01\ g\ C}{44.01\ g\ CO_2} \qquad \frac{2(1.008)\ g\ H}{18.02\ g\ H_2O}$$

2. Find the mass of O by difference:

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

3. Follow the schematic pathway shown in Figure 3.5.

Example 3.6, (Cont'd)

SOLUTION

1. mass of C	$7.33\ g\ CO_2 \times \frac{12.01\ g\ C}{44.01\ g\ CO_2} = 2.00\ g\ C$
mass of H	$3.00\ g\ H_2O \times \frac{2(1.008)\ g\ H}{18.02\ g\ H_2O} = 0.336\ g\ H$
2. mass of O	mass of O = mass of sample − (mass of C + mass of H) = 5.00 g − (2.00 g + 0.336 g) = 2.66 g
3. mol of each element	C: $\frac{2.00\ g\ C}{12.01\ g/mol} = 0.167$; H: $\frac{0.336\ g\ H}{1.008\ g/mol} = 0.333$; O: $\frac{2.66\ g\ O}{16.00\ g/mol} = 0.166$
ratios	C: $\frac{0.167\ mol}{0.166\ mol} = 1$; H: $\frac{0.333\ mol}{0.167\ mol} = 2$; O: $\frac{0.166\ mol}{0.166\ mol} = 1$
simplest formula	Using the ratios as subscripts, the simplest formula 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 – 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 3.7

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.

ANALYSIS

Information given:	simplest formula (CH_2O) MM_a: actual molar mass (60 g/mol)
Asked for:	molecular formula

Example 3.7 – Simplest Formula from Mass Percents , Cont'd

STRATEGY

1. Determine the molar mass of the simplest formula, MM_s
2. Find the ratio

$$\frac{\text{actual molar mass}}{\text{simplest formula molar mass}} = \frac{MM_a}{MM_s}$$

3. To get the molecular formula, multiply all subscripts in the simplest formula by the ratio.

SOLUTION

1. MM_s	12.01 g C + 2(1.008) g H + 16.00 g O = 30.03 g CH_2O/mol
2. ratio	$\frac{MM_a}{MM_s} = \frac{60}{30.03} = 2$
3. molecular formula	$C_{1\times2}H_{2\times2}O_{1\times2}$ The molecular formula of acetic acid is $C_2H_4O_2$.

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

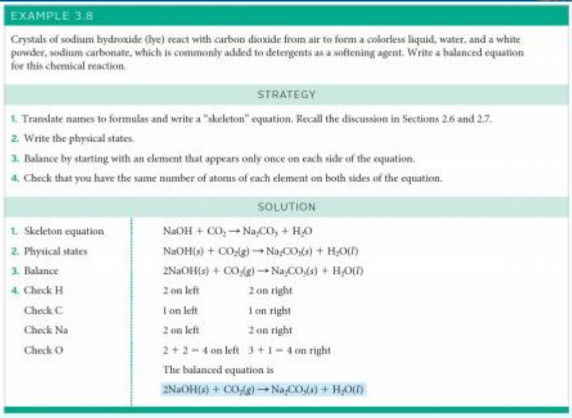
How are Equations Written?

- 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

1. Translate names to formulas and write a "skeleton" equation. Recall the discussion in Sections 2.6 and 2.7.
2. Write the physical states.
3. Balance by starting with an element that appears only once on each side of the equation.
4. Check that you have the same number of atoms of each element on both sides of the equation.

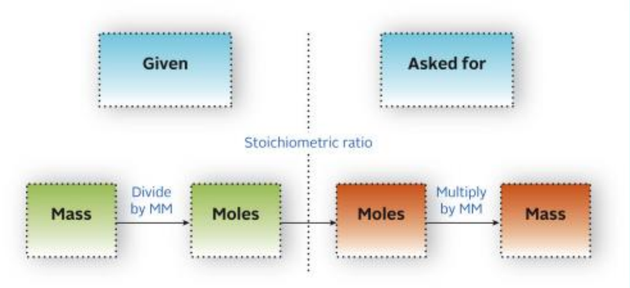
SOLUTION

1. Skeleton equation	$NaOH + CO_2 \rightarrow Na_2CO_3 + H_2O$	
2. Physical states	$NaOH(s) + CO_2(g) \rightarrow Na_2CO_3(s) + H_2O(l)$	
3. Balance	$2NaOH(s) + CO_2(g) \rightarrow Na_2CO_3(s) + H_2O(l)$	
4. Check H	2 on left	2 on right
Check C	1 on left	1 on right
Check Na	2 on left	2 on right
Check O	2 + 2 = 4 on left	3 + 1 = 4 on right
	The balanced equation is	
	$2NaOH(s) + CO_2(g) \rightarrow Na_2CO_3(s) + H_2O(l)$	

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

Figure 3.8: Flowchart for Mole-Mass Calculations



Example 3.9

EXAMPLE 3.9

Ammonia is used to make fertilizers for lawns and gardens by reacting nitrogen gas with hydrogen gas. The balanced equation for the reaction is

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

a How many moles of ammonia are formed when 1.34 mol of nitrogen react?
b How many grams of hydrogen are required to produce 2.75×10^3 g of ammonia?
c How many molecules of ammonia are formed when 2.92 g of hydrogen react?
d 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.

continued

Example 3.9, (Cont'd)

STRATEGY

For all parts of this example, follow the schematic pathway shown in Figure 3.8.

a

ANALYSIS	
Information given:	balanced equation: [$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$] moles N_2 (1.34)
Asked for:	mol NH_3 formed
SOLUTION	
mol NH_3	mol $H_2 \rightarrow$ mol NH_3 $1.34 \text{ mol } H_2 \times \frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2} = 2.68 \text{ mol } NH_3$

Example 3.9, (Cont'd)

b

ANALYSIS	
Information given:	mass of ammonia (2.75×10^3 g) balanced equation: [$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$]
Information implied:	molar masses of NH_3 and H_2
Asked for:	mass of H_2 needed
SOLUTION	
mass H_2	mass $NH_3 \rightarrow$ mol $NH_3 \rightarrow$ mol $H_2 \rightarrow$ mass H_2 $2.75 \times 10^3 \text{ g } NH_3 \times \frac{1 \text{ mol } NH_3}{17.03 \text{ g } NH_3} \times \frac{3 \text{ mol } H_2}{2 \text{ mol } NH_3} \times \frac{2.016 \text{ g } H_2}{1 \text{ mol } H_2} = 488 \text{ g } H_2$

c

ANALYSIS	
Information given:	mass of H_2 (2.92 g) balanced equation: [$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$]
Information implied:	molar mass of H_2 Avogadro's number (N_A)
Asked for:	molecules of NH_3 produced
SOLUTION	
molecules NH_3	mass $H_2 \rightarrow$ mol $H_2 \rightarrow$ mol $NH_3 \xrightarrow{N_A}$ molecules NH_3 $2.92 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2} \times \frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol } NH_3} = 5.81 \times 10^{23}$

continued

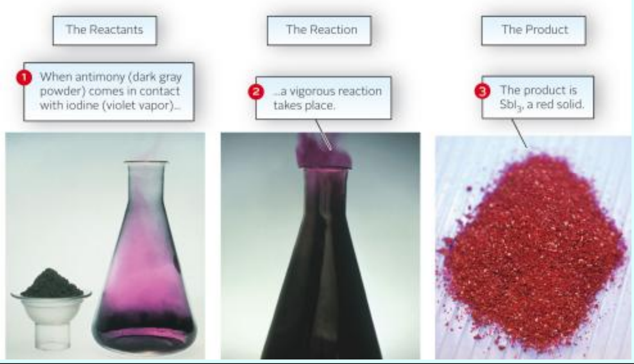
Example 3.9, (Cont'd)

d

ANALYSIS	
Information given:	balanced equation: [$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$] V_{air} (15.0 L); % by volume of N_2 in air (79%); density, *d*, of N_2 (1.25 g/L)
Information implied:	molar masses of N_2 and NH_3
Asked for:	mass of NH_3 produced
SOLUTION	
mass NH_3	$V_{air} \xrightarrow{\% N_2} V_{N_2} \xrightarrow{\text{density}}$ mass $N_2 \rightarrow$ mol $N_2 \rightarrow$ mol $NH_3 \rightarrow$ mass NH_3 $15.0 \text{ L air} \times \frac{79 \text{ L } N_2}{100 \text{ L air}} \times \frac{1.25 \text{ g } N_2}{1 \text{ L } N_2} \times \frac{1 \text{ mol } N_2}{28.02 \text{ g } N_2} \times \frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2} \times \frac{17.03 \text{ g } NH_3}{1 \text{ mol } NH_3} = 18 \text{ g } NH_3$

Limiting Reactant and Theoretical Yield

- $2 \text{ Sb (s)} + 3 \text{ I}_2 \text{ (s)} \rightarrow 2 \text{ SbI}_3 \text{ (s)}$



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?

(a)

ANALYSIS

Information given:	moles of each reactant: Sb (1.20), I_2 (2.40) balanced equation: [$2Sb(s) + 3I_2(s) \rightarrow 2SbI_3(s)$]
Asked for:	limiting reactant theoretical yield

continued

Example 3.10, (Cont'd)

STRATEGY

1. Find mol SbI_3 produced by first assuming Sb is limiting, and then assuming I_2 is limiting.
   mol Sb $\rightarrow$ mol SbI_3; mol $I_2 \rightarrow$ mol SbI_3
2. The reactant that gives the smaller amount of SbI_3 is limiting, and the smaller amount of SbI_3 is the theoretical yield.

SOLUTION

mol SbI_3	$1.20 \text{ mol Sb} \times \frac{2 \text{ mol } SbI_3}{2 \text{ mol Sb}} = 1.20 \text{ mol}$ $\quad 2.40 \text{ mol } I_2 \times \frac{2 \text{ mol } SbI_3}{3 \text{ mol } SbI_3} = 1.60 \text{ mol}$
limiting reactant	1.20 mol (Sb limiting) < 1.60 mol (I_2 limiting) The limiting reactant is Sb.
theoretical yield	1.20 mol < 1.60 mol The theoretical yield is 1.20 mol SbI_3

(b)

ANALYSIS

Information given:	mass of each reactant: Sb (1.20 g), I_2 (2.40 g) balanced equation: [$2Sb(s) + 3I_2(s) \rightarrow 2SbI_3(s)$]
Information implied:	molar masses (MM) of SbI_3 and I_2
Asked for:	limiting reactant theoretical yield mass of excess reactant not used up

Example 3.10, (Cont'd)

STRATEGY

1. Follow the plan outlined in Figure 3.8 and convert mass of Sb and mass of I_2 to mol SbI_3.
2. The smaller number of moles SbI_3 obtained is the theoretical yield. The reactant that yields the smaller amount is the limiting reactant.
3. Convert moles limiting reactant to mass of excess reactant. That is the mass of excess reactant consumed in the reaction.
4. Mass of excess reactant not used up = mass of excess reactant initally − mass excess reactant consumed

SOLUTION

1. mol SbI_3	$1.20 \text{ g Sb} \times \frac{1 \text{ mol Sb}}{121.8 \text{ g Sb}} \times \frac{2 \text{ mol } SbI_3}{2 \text{ mol Sb}} = 0.00985 \text{ mol } SbI_3$ $2.40 \text{ g } I_2 \times \frac{1 \text{ mol } I_2}{253.8 \text{ g } I_2} \times \frac{2 \text{ mol } SbI_3}{3 \text{ mol } I_2} = 0.006304 \text{ mol } SbI_3$
2. limiting reactant	0.006304 mol (I_2 limiting) < 0.00985 mol (Sb limiting); thus I_2 is the limiting reactant.
theoretical yield	0.006304 mol < 0.00985 mol The theoretical yield is 0.006304 mol (3.17 g) SbI_3. The reactant in excess is Sb.
3. mass Sb used up	$2.40 \text{ g } I_2 \times \frac{1 \text{ mol } I_2}{253.8 \text{ g } I_2} \times \frac{2 \text{ mol Sb}}{3 \text{ mol } I_2} \times \frac{121.8 \text{ g Sb}}{1 \text{ mol Sb}} = 0.768 \text{ g Sb}$
4. mass unreacted	mass unreacted = mass present initially − mass used up = 1.20 g − 0.768 g = 0.43 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?

ANALYSIS

Information given:	From Example 3.10a, theoretical yield (1.20 mol) percent yield (78.2%)
Asked for:	mass SbI_3 actually obtained

STRATEGY

1. Substitute into Equation 3.3.

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

2. Your answer will be the actual yield in moles. Convert to grams.

continued

Example 3.11, (Cont'd)

SOLUTION

actual yield	$78.2\% = \dfrac{\text{actual yield}}{1.20\ \text{mol}} \times 100\%$; actual yield $= 0.938$ mol SbI_3
mass SbI_3	$0.938\ \text{mol}\ SbI_3 \times \dfrac{502.5\ \text{g}\ SbI_3}{1\ \text{mol}\ SbI_3} = 472\ \text{g}$

END POINT

If your actual yield is larger than your theoretical yield, something's wrong!

Key Concepts

1. Use molar mass to relate
   - Moles to mass
   - Moles in solution; molarity
   - Molecular formula to simplest formula
2. Use the formula of a compound to find percent composition or its equivalent.
3. Find the simplest formula from chemical analysis data.
4. Balance chemical equations by inspection

Key Concepts, (Cont'd)

5. Use a balanced equation to
   - Relate masses of products and reactants
   - Find the limiting reactant, theoretical yield and percent yield