Chapter 3
Stoichiometry of Formulas & Equations

- Mole - Mass Relationships in Chemical Systems
- Determining the Formula of an Unknown Compound
- Writing and Balancing Chemical Equations
- Calculating the Amounts of Reactant and Product
- Calculating Limiting Reagent & Theoretical Yield
- Fundamentals of Solution Stoichiometry

Stoichiometry

- Stoichiometry – The study of the quantitative relationships between elements, compounds, chemical formulas and chemical reactions
  - Each element in a compound has a unique atomic mass (total mass of protons & neutrons)
  - The concept of the "MOLE" was developed to relate the number of entities in a substance to the mass values we determine in the laboratory
  - From the relationship between the number of atoms and the mass of a substance we can quantify the relationship between elements and compounds in chemical reactions

Mass vs. Amount

- The standard unit of mass in the metric system is the gram (or kilogram)
- Each of the 100 or so different elements has a unique mass (atomic weight) expressed as either atomic mass units (amu) or grams determined by the number of protons and neutrons in the nucleus
- The same mass (weight) of two different substances will represent a <u>different</u> number of atoms
- A chemical equation defines the relative number of molecules of each component involved in the reaction
- The "Mole" establishes the relationship between the number of atoms of a given element and the mass of the substance used in a reaction

Mass vs. Amount

- Amounts in chemistry are expressed by the mole
  - mole – quantity of substance that contains the same number of molecules or formula units as exactly 12 g of Carbon-12
  - Number of atoms in 12 g of Carbon-12 is Avogadro's number (N_A) which equals 6.022 x 10^{23}
  - The atomic mass of one atom expressed in atomic mass units (amu) is numerically the same as the mass of 1 mole of the element expressed in grams
- Molar Mass = mass of 1 mole of substance
  - One molecule of Carbon (C) has an atomic mass of 12.0107 amu and a molar mass of 12.0107 g/mol
  - 1 mole of Carbon contains 6.022 x 10^{23} atoms
  - 1 mole of Sodium contains 6.022 x 10^{23} atoms

Molecular & Formula Weight

- Molecular Mass (also referred to as Molecular Weight (MW) is the sum of the atomic weights of all atoms in a covalently bonded molecule – organic compounds, oxides, etc.
- Formula Mass is sometimes used in a more general sense to include Molecular Mass, but its formal definition refers to the sum of the atomic weights of the atoms in ionic bonded compounds

Molecular & Formula Weight

The computation of Molecular (covalent) or Formula (ionic) molar masses is mathematically the same

Ex.

Molecular Molar Mass of Methane (CH_4) (covalent bonds)

1 mol CH_4 = 1 mol C/mol CH_4 x 12 g/mol C = 12 g

4 mol H/mol CH_4 x 1 g/mol H = 4 g

= 16 g/mol CH_4 = 6.022×10^{23} molecules

Formula Molar Mass of Aluminum Phosphate ($AlPO_4$) (ionic bonds)

1 mol $AlPO_4$ = 1 mol Al/mol $AlPO_4$ x 27 g/mol Al) = 27 g

1 mol P/mol $AlPO_4$ x 31 g/mol P = 31 g

4 mol O/mol $AlPO_4$ x 16 g/mol O) = 64 g

= 122 g/mol $AlPO_4$ = 6.022×10^{23} molecules

The Concept of Amount

- Summary of Mass Terminology

Isotopic Mass	Mass of an isotope of an element in atomic mass units (amu)
Atomic Mass (atomic weight)	Average of the masses of the naturally occurring isotopes of an element based on relative abundance 1 amu $=1.66054 \times 10^{-24}$ g Atomic Mass Oxygen (O) = 15.9994 amu Mass (O)=15.9994 amu x 1.66054×10^{-24} g/amu $=2.65676 \times 10^{-23}$g
Molecular Mass Formula Mass	Sum of the atomic masses of the atoms or ions in a molecule or formula unit
Mole	Quantity of substance that contains the same number of molecules or formula units as exactly 12 g of Carbon-12 – 6.022×10^{23} molecules
Molar Mass	Mass of 1 mole (6.022×10^{23} molecules) of a chemical entity (gram-molecular weight) atom, ion, molecule, formula unit

Mole Relationships: Example Calculations

How many molecules of H_2O are in 251 kg of water?

251 kg x (1000 g/kg) = 2.51×10^5 g H_2O

2.51×10^5 g H_2O x (1 mol H_2O/18.0153 g) = 1.39326×10^4 mol H_2O

1.39326×10^4 mol x 6.022×10^{23} atoms/mol = 8.39021×10^{27} molecules

How many total atoms are in 251 kg of water?

8.39021×10^{27} molecules x (3 atoms/1 molecule) = 2.52×10^{28} atoms

Practice Problems

What is the molar mass of Caffeine, $C_8H_{10}N_4O_2$?

C = 12.0107 g/mol H = 1.00794 g/mol

N = 14.0067 g/mol O = 15.9994 g/mol

$12.0107\ g/\cancel{mol}\ C \times 8\ \cancel{mol}\ C/mol\ C_8H_{10}N_4O_2 = 96.0856\ g/mol\ C_8H_{10}N_4O_2$

$1.00794\ g/\cancel{mol}\ H \times 10\ \cancel{mol}\ H/mol\ C_8H_{10}N_4O_2 = 10.0794\ g/mol\ C_8H_{10}N_4O_2$

$14.0067\ g/\cancel{mol}\ N \times 4\ \cancel{mol}\ N/mol\ C_8H_{10}N_4O_2 = 56.0268\ g/mol\ C_8H_{10}N_4O_2$

$15.9994\ g/\cancel{mol}\ O \times 2\ \cancel{mol}\ O/mol\ C_8H_{10}N_4O_2 = 31.9988\ g/mol\ C_8H_{10}N_4O_2$

Sum of elemental masses = molecular mass of Caffeine

96.0856 + 10.0794 + 56.0268 + 31.9988 =

194.1906 g/mol $C_8H_{10}N_4O_2$

Practice Problem

How many Sulfur atoms are in 25 g of Al_2S_3?

Al = 26.9815 g/mol S = 32.065 g/mol

Al_2S_3 = 26.98915 g/$\cancel{\text{mol}}$ Al x 2 $\cancel{\text{mol}}$ Al + 32.065 g/$\cancel{\text{mol}}$ S x 3 $\cancel{\text{mol}}$ S

= 150.158 g/mol Al_2S_3

25 $\cancel{\text{g}}$ Al_2S_3 / 150.158 $\cancel{\text{g}}$/mol Al_2S_3 = 0.166491 mol Al_2S_3

Compute moles of Sulfur atoms

0.166491 mol $\cancel{Al_2S_3}$ x 3 mol S/1mol $\cancel{Al_2S_3}$ = 0.499474 mol S atoms

Compute atoms of Sulfur

0.499474 mol $\cancel{\text{S}}$ atoms x 6.022 x 10^{23} S atoms/1mol $\cancel{\text{S}}$ atoms =

3.008 x 10^{23} atoms S

Percent Composition

- It is often necessary to determine the mass percentage of a component in a mixture or an element in a compound

$$\textbf{Mass \% A} = \frac{\textbf{Mass of A in Whole}}{\textbf{Mass of Whole}} \times \textbf{100}$$

- <u>Example calculation</u>: What are the mass percentages of C, H and O in $C_2H_4O_2$ (Acetic Acid)?

1 mol acetic acid = 60.052 g

% C = [2 mol C x (12.0107 g/mol C)] ÷ 60.052 g/mol x 100 = 40.00%

% H = [4 mol H x (1.00794 g/mol C)] ÷ 60.052 g/mol x 100 = 6.71%

% O = [2 mol O x (15.9994 g/mol C)] ÷ 60.052 g/mol x 100 = 53.29%

Practice Problem

What is the mass percentage of C in in l-Carvone, $C_{10}H_{14}O$, which is the principal component of spearmint?

C = 12.0107 g/mol H = 1.00794 g.mol O = 15.9994 g/mol

a. 30% b. 40% c. 60% d. 70% e. 80%

Ans: e

Molar Mass C = 12.0170 g/mol C x 10 mol C = 120.170 g C

H = 1.00794 g/mol H x 14 mol H = 14.1112 g H

O = 15.9994 g/mol O x 1 mol O = 15.9994 g O

Molar Mass $C_{10}H_{14}O$ = 150.218 g/mol

Mass % C = 120.170 / 150.218 x 100 = 79.9971 (80%)

Empirical & Molecular Formulas

- **Empirical formula** – formula of a substance written with the smallest whole number subscripts
- EF of Acetic Acid = $C_2H_5O_2$
- For small molecules, empirical formula is identical to the **molecular formula**: formula for a single molecule of substance
- For Succinic acid, its molecular formula is: $C_4H_6O_4$

  Its empirical formula is: $C_2H_3O_2$ (n = 2)
- Molecular weight = n x empirical formula weight (n = number of empirical formula units in the molecule)

Practice Problem

Of the following, the only <u>empirical</u> formula is

a. C_2H_4 b. C_5H_{12} c. N_2O_4 d. S_8 e. N_2H_4

Ans: b

Subscript (5) cannot be further divided into whole numbers

Molecular Formula from Elemental Analysis:

A moth repellant, para-dichlorobenzene, has the composition 49.1% C, 2.7% H and 48.2% Cl. Its molecular weight is determined from mass spectrometry (next slide). What is its molecular formula?

Assume a sample mass of 100 grams

49.1 g C x 1 mol C	/	12.0107	g C	= 4.0880 mol C
2.7 g H x 1 mol H	/	1.00794	g H	= 2.6787 mol H
48.2 g Cl x 1 mol Cl	/	35.453	g Cl	= 1.3595 mol Cl

Convert Mole values to "Whole" numbers (divide each value by smallest)

4.0880 / 1.3595 = 3.01 (3 mol C)

2.6787 / 1.3595 = 1.97 (2 mol H)

1.3595 / 1.3595 = 1.00 (1 mol Cl)

$\therefore$ Empirical Formula is: C_3H_2Cl

Con't on next slide

Molecular Formula from Elemental Analysis: An Example Calculation (Con't)

Empirical formula = C_3H_2Cl

Compute Empirical Formula Weight (EFW)

$$\text{EFW} = \underbrace{3 \times 12.01}_{\text{Carbon}} + \underbrace{2 \times 1.01}_{\text{Hydrogen}} + \underbrace{1 \times 35.45}_{\text{Chlorine}} = 73.51 \text{ amu}$$

EFW = 73.51 g/mol

Molecular weight (M^+ ion from mass spectrum) = 146 amu

$$n = 146/73.51 = 1.99 = 2$$

$\therefore$ Molecular Formula = $C_6H_4Cl_2$

Empirical Formula from Mole Fraction

- A sample of an unknown compound contains:

0.21 mol Zn   0.14 mol P   0.56 mol O

What is the Empirical Formula?

Ans: Express preliminary formula using mole fraction values

$$Zn_{0.21}P_{0.14}O_{0.56}$$

Divide Each Fraction Value by the Smallest Fraction value

0.21 / 0.14 = 1.5

0.14 / 0.14 = 1.0

0.56 / 0.14 = 4.0

$$Zn_{1.5}P_{1.0}P_{4.0}$$

Multiply through by the smallest integer that turns all subscripts into whole number integers

$$Zn_{(1.5\ x2)}P_{1.0x2)}P_{(4.0x2)} \quad = \quad Zn_3P_2O_8$$

The Chemical Equation

A chemical equation is the representation of the reactants and products in a chemical reaction in terms of chemical symbols and formulas

The subscripts represent the number of atoms of an element in the compound

The coefficients in front of the compound represents the number of moles of each compound required to balance the equation

The Chemical Equation

A balanced equation will have an equal number of atoms of each element on both sides of equation

$$\underbrace{N_2(g) \quad + \quad O_2(g)}_{\text{Reactants}} \quad \rightarrow \quad \underbrace{2\,NO(g)}_{\text{Products}}$$

(4 atoms on each side)

1 mole nitrogen + 1 mole oxygen yields 2 moles nitrogen monoxide

Phase representations in Chemical Equations

$\rightarrow$ = yields, or forms (g) = gas phase

(l) = liquid phase (s) = Solid phase

The Chemical Equation

- Example Problem

  Balance the following reaction

$$C_8H_{18} \quad + \quad O_2 \rightarrow \quad CO_2 \quad + \quad H_2O$$

$$1\ C_8H_{18} \quad + \quad O_2 \rightarrow \quad 8\ CO_2 \quad + \quad H_2O$$

$$1\ C_8H_{18} \quad + \quad O_2 \rightarrow \quad 8\ CO_2 \quad + \quad 9\ H_2O$$

$$1\ C_8H_{18} \quad + \quad {}^{25/2}\ O_2 \rightarrow \quad 8CO_2 \quad + \quad 9\ H_2O$$

$$2\ C_8H_{18} \quad + \quad 25\ O_2 \rightarrow \quad 16\ CO_2 \quad + \quad 18H_2O$$

Stoichiometry in Chemical Equations

When Dinitrogen Pentoxide, N_2O_5, a while solid, is heated, it decomposes to Nitrogen Dioxide and Oxygen

$2\ N_2O_5(s) \quad \rightarrow \quad 4\ NO_2(g) \quad + \quad O_2(g)$ Molar Ratio 2 : 4 : 1

If a sample of N_2O_5 produces 1.315 g of O_2, how many grams of NO_2 are formed? How many grams of N_2O_5 are consumed?

Strategy:

1. Compute actual no. of moles of oxygen produced

   $1.315\ \cancel{g\ O_2} \quad x \quad (1\ mol\ O_2/32.00\ \cancel{g\ O_2}) = 0.04109$ moles Oxygen (O_2)

2. Determine molar ratio of NO_2 & N_2O_5 relative to O_2

   (4:1 & 2:1)

2. Compute mass of NO_2 produced from molar ratio and actual moles O_2

   $0.04109\ mol\ \cancel{O_2}\ x\ (4\ mol\ \cancel{NO_2}/1\ mol\ \cancel{O_2})\ x\ (46.01\ g\ NO_2/1\ mol\ \cancel{NO_2}) = 7.563\ g\ NO_2$

4. Compute mass of N_2O_5 from molar ratio and actual moles O_2

   $0.04109\ mol\ \cancel{O_2}\ x\ (2\ mol\ \cancel{N_2O_5}\ /1\ mol\ \cancel{O_2})\ x\ (108.0\ g\ N_2O_5\ /mol\ \cancel{N_2O_5}) = 8.834g\ N_2O_5$

Note: Mass Ratio $8.876\ g\ N_2O_5 \quad \rightarrow \quad 7.563\ g\ NO_2 \quad + \quad 1.315\ g\ O_2$

Stoichiometry in Chemical Equations

How many grams of HCl are required to react with 5.00 grams Manganese Dioxide (MnO_2) according to the equation?

$$4\ HCl(aq) + MnO_2(s) \rightarrow 2\ H_2O(l) + MnCl_2(aq) + Cl_2(g)$$

Strategy: 1. Determine the Molar Ratio of HCL to MnO_2

2. Compute the no. moles MnO_2 actually used

3. Use actual moles MnO_2 & Molar ratio to compute mass HCL

Molar Ratio HCl : MnO_2 = 4 : 1

$5.00\ \cancel{g\ M}nO_2 \times (1\ mol\ MnO_2/86.9368\ \cancel{g\ M}nO_2) = 0.575\ mol\ MnO_2$

$0.575\ mol\ \cancel{M}nO_2 \times (4\ mol\ H\cancel{C}l/1\ mol\ \cancel{M}nO_2) \times (36.461\ g\ HCl/mol\ \cancel{H}Cl)$

$= 8.39\ g\ HCl$

Reactions that Occur in a Sequence

- In many situations, a product of one reaction becomes a reactant for the next
- For stoichiometric purposes, when the same (common) substance forms in one reaction and reacts (used up) in the next, it is eliminated in the overall reaction
- Steps in the addition of reactions:
  - Write the sequence of balanced equations
  - Adjust the equations arithmetically to cancel the common substance
  - Add the adjusted equations together to obtain the overall balanced equation

Reactions that Occur in a Sequence

- Ex. Write the two balanced equations

$$2Cu_2S(s) + 3O_2(g) \rightarrow 2Cu_2O(s) + 2SO_2(g)$$

$$Cu_2O(s) + C(s) \rightarrow 2Cu(s) + CO(g)$$

- Adjust Coefficients – Multiply 2nd equation by 2

$$2Cu_2S(s) + 3O_2(g) \rightarrow \cancel{2Cu_2O(s)} + 2SO_2(g)$$

$$\cancel{2Cu_2O(s)} + 2C(s) \rightarrow 4Cu(s) + 2CO(g)$$

$$2Cu_2S(s) + 3O_2(g) + 2C(s) \rightarrow 4Cu(s) + 2CO(g)) + 2SO_2(g)$$

The common compound in both reactions (Cu_2O) is eliminated

Biological systems have many examples of Multistep reaction sequences called "Metabolic Pathways"

Limiting Reactants and Yields

Limiting Reagent & Theoretical Yield

- The "Limiting Reagent" is that reactant whose mass (on a molar equivalent basis) actually consumed in the reaction is less than the amount of the other reactant, i.e., the reactant in excess

- From the Stoichiometric balanced equation determine the molar ratio among the reactants and products, i.e., how many moles of reagent A react with how many moles of reagent B to yield how many moles of product C, D, etc.

Limiting Reactants and Yields

- If the ratio of moles of A to moles of B <u>actually used</u> is <u>greater</u> than the Stoichiometric molar ratio of A to B, then

reagent A is in "Excess"

reagent B is "Limiting"

- If, however, the actual molar ratio of A to B used is <u>less</u> than the Stoichiometric molar ratio, then B is in excess and A is "Limiting"
- The moles of product(s) (theoretical yield) is determined by the moles of "limiting Reagent" on a molar equivalent basis

Limiting Reactants and Yields

Example 1

$$A \quad + \quad B \quad \rightarrow \quad C \qquad \text{Molar Ratio A:B} = 1$$

Moles actually used: A = 0.345 B = 0.698

Ratio of moles actually used (A/B)

$$0.345/0.698 = 0.498$$

$$0.498 < 1.0$$

∴ On a molar equivalent basis (1/1) there is not enough reagent A (0.345 mol) to react with reagent B (0.498 mol); therefore reagent B is in excess & reagent A is Limiting

∴ Since 1 mol "A" produces 1 mol "C"

Theoretical Yield of "C" = 0.345 moles

Limiting Reactants and Yields

Example 2

$$A \quad + \quad B \rightarrow C$$

Stoichiometric Molar ratio A:B = 1 : 1 = 1.0

Moles actually used: A = 0.20 B = 0.12

Ratio of Moles actually used (A/B):

0.20 / 0.12 = 1.67

The ratio of A:B is greater than 1.00

A is in excess and B is limiting

Only 0.12 moles of the 0.2 moles of "A" would be required to react with the 0.12 moles of B

The reaction would have a theoretical yield of:

0.12 moles of "C" (Molar Ratio of B:C = 1:1)

Limiting Reactants and Yields

Example 3

$$A \quad + \quad 2B \quad \rightarrow \quad C$$

Stoichiometric Molar ratio A:B = 1 : 2 = 0.5

Moles actually used: A = 0.0069; B = 0.023

Ratio of Moles actually used (A/B):

0.0069 / 0.023 = 0.30 < 0.5

$\therefore$ "A" is limiting

Only 0.0069 • 2 = 0.0138 moles of the 0.023 moles of B are required to react with the 0.0069 moles of A

Since 0.0138 < 0.023 "B" is in excess, "A" is limiting

The reaction would have a theoretical yield of:

0.0069 moles of C (Molar Ratio of A:C = 1:1)

Limiting Reactants and Yields

Theoretical Yield & Percent Yield

- The Theoretical Yield, in grams, is computed from the number of moles of the "Limiting Reagent", the Stoichiometric Molar Ratio, and the Molecular Weight of the product

Yield = mol (Lim) x Mol Ratio Prod/Lim x Mol Wgt Product

- The Percent Yield of a product obtained in a "Synthesis" experiment is computed from the amount of product actually obtained in the experiment and the Theoretical Yield

% Yield = Actual Yield / Theoretical Yield x 100

Note: The yield values can be expressed in either grams or moles

Example Yield Calculation

Methyl Salicylate (MSA) is prepared by heating Salicylic Acid (SA), $C_7H_6O_3$, with Methanol (ME), CH_3OH

$$C_7H_6O_3 \quad + \quad CH_3OH \quad \rightarrow \quad C_8H_8O_3 \quad + \quad H_2O$$

1.50 g of Salicylic acid (SA) is reacted with 11.20 g of Methanol (ME). The yield of Methyl Salicylate is 1.27 g. What is the limiting reactant? What is the percent yield of Methyl Salicylate (MSA)?

Molar Ratio: 1 mole SA reacts with 1 mole ME to produce 1 mole MSA

Moles SA: 1.50 g SA x (1 mol SA/138.12 g SA) = 0.0109 mol SA

Moles ME: 11.20 g ME x (1 mol ME/32.04 g ME) = 0.350 mol ME

0.0109 mol SA x (1mol SA/1 mol ME) $<$ 0.350 mol ME

$\therefore$ Salicylic acid (SA) is limiting; Methanol (ME) is in "Excess"

<u>Theoretical Yield</u> = 0.0109 mol SA x (1 mol MSA/1 mol SA) x (152.131 g MSA/1 mol MSA) = 1.66 g MSA

Example Yield Calculation

Hydrogen (H_2) is a possible clean fuel because it reacts with Oxygen (O) to form non-polluting water (H_2O)

$$2\ H_2(g) \quad + \quad O_2(g) \quad \rightarrow \quad 2\ H_2O(g)$$

If the yield of this reaction is 87% what mass of Oxygen is required to produce 105 kg of Water?

Molar Ratio: 2 mol H_2 reacts with 1 mol O_2 to form 2 mol Water (H_2O)

Moles H_2O: 105 kg H_2O x (1000g/1 kg) x (1 mol H_2O/18.01 g/mol H_2O)

= 5,830 mol H_2O

Moles O_2 = (1 mol O_2/2 mol H_2O) x 5,830 mol H_2O = 2,915 mol O_2

Mass O_2 = 2,915 mol O_2 x (32.0 g O_2 / 1 mol O_2)x (1 kg/1000 g)

= 93.2 kg O_2 required to produce 105 kg H_2O (100%)

At 87% efficiency: 93.2 kg x 100%/87% = 107 kg O_2 required

Sample Problem

In the study of the following reaction:

$$2N_2H_4(l) \;+\; N_2O_4(l) \;\rightarrow\; 3N_2(g) \;+\; 4H_2O(g)$$

the yield of N_2 was less than expected

It was then discovered that a 2nd side reaction also occurs:

$$N_2H_4(l) \;+\; 2N_2O_4(l) \;\rightarrow\; 6NO(g) \;+\; 2H_2O)g)$$

In one experiment, 10.0 g of NO formed when 100.0 g of each reactant was used

What is the highest percent yield of N_2 that can be expected?

Answer on next Slide

Sample Problem

Ans:

If 100.0 g of Dinitrogen Tetroxide (N_2O_4) reacts with 100.0 g of Hydrazine (N_2H_4), what is the theoretical yield of Nitrogen if no side reaction takes place?

First, we need to identify the limiting reactant!

The limiting reactant is used to calculate the theoretical yield

Determine the amount of limiting reactant required to produce 10.0 grams of NO

Reduce the amount of limiting reactant by the amount used to produce NO

The reduced amount of limiting reactant is then used to calculate an "actual yield"

The "actual" and theoretical yields will give the maximum percent yield

Con't on next Slide

Sample Problem (con't)

Solution (con't):

Determining the limiting reagent:

$$N_2 \text{ from } N_2O_4 = (100.0\text{ g } N_2O_4)\left(\frac{1\text{ mol } N_2O_4}{92.02\text{ g } N_2O_4}\right)\left(\frac{3\text{ mol } N_2}{1\text{ mol } N_2O_4}\right) = 3.26016\text{ mol } N_2$$

$$N_2 \text{ from } N_2H_4 = (100.0\text{ g } N_2H_4)\left(\frac{1\text{ mol } N_2H_4}{32.05\text{ g } N_2H_4}\right)\left(\frac{3\text{ mol } N_2}{2\text{ mol } N_2H_4}\right) = 4.68019\text{ mol } N_2$$

$$\text{Molar Ratio} = \frac{3.26016\text{ mol } N_2}{4.68019\text{ mol } N_2} = 0.696779 < 1.0$$

$\therefore$ N_2O_4 is the limiting reagent - Produces 3.26016 mol N_2 vs.

4.68019 mol N_2 from N_2H_2

$$\text{Theoretical Yield of } N_2 = (100.0\text{ g } N_2O_4)\left(\frac{1\text{ mol } N_2O_4}{92.02\text{ g } N_2O_4}\right)\left(\frac{3\text{ mol } N_2}{1\text{ mol } N_2O_4}\right)\left(\frac{28.02\text{ g } N_2}{1\text{ mol } N_2}\right) = 91.3497\text{ g } N_2$$

Con't on next Slide

Sample Problem (con't)

Soln (con't)

How much limiting reagent (N_2O_4) is used to produce 10.0 g NO?

$$\text{Grams } N_2O_4 \text{ used} = (10.0\text{ g NO})\left(\frac{1\text{ mol NO}}{30.01\text{ g NO}}\right)\left(\frac{2\text{ mol } N_2O_4}{6\text{ mol NO}}\right)\left(\frac{92.02\text{ g } N_2O_4}{1\text{ mol } N_2O_4}\right) = 10.221\text{ g } N_2O_4$$

Determine the actual yield

$$\text{Initial amount of } N_2O_4 \text{ available} = 100.0\text{ g}$$

$$\text{Mass of } N_2O_4 \text{ remaining to produce Nitrogen } (N_2)$$

$$100.0\text{ g} - 10.221\text{ g} = 89.779\text{ g } N_2O_4$$

$$\text{Actual yield of } N_2 = (89.779\text{ g } N_2O_4\left(\frac{1\text{ mol } N_2O_4}{92.02\text{ g } N_2O_4}\right)\left(\frac{3\text{ mol } N_2}{1\text{ mol } N_2O_4}\right)\left(\frac{28.02\text{ g } N_2}{1\text{ mol } N_2}\right) = 82.01285\text{ g } N_2$$

$$\text{Theoretical Yield} = \left(\frac{\text{Actual Yield}}{\text{Theoretical Yield}}\right)(100) = \left(\frac{82.01285}{91.3497}\right)(100) = 89.8\%$$

Solution Stoichiometry

- **Solute** – A substance dissolved in another substance
- **Solvent** – The substance in which the "Solute" is dissolved
- **Concentration** – The amount of solute dissolved in a given amount of solvent
- **Molarity (M)** – Expresses the concentration of a solution in units of moles solute per liter of solution

$$\textbf{Molarity} = \frac{\textbf{moles of solute}}{\textbf{liters of solution}} \qquad \textbf{M} = \frac{\textbf{mol solute}}{\textbf{L soln}}$$

- **Molality (m)** – Expresses the number of moles dissolved in 1000g (1KG) of solvent.

$$\textbf{Molality} = \frac{\textbf{Moles Solute}}{\textbf{Kilogram Solvent}} = \textbf{m} = \frac{\textbf{mol solute}}{\textbf{Kg solvent}}$$

Solution Volume vs. Solvent Volume

- The Volume term in the denominator of the molarity expression is the solution volume not the volume of the solvent
- 1 mole of solute dissolved in 1 Liter of a solvent does not produce a 1 molar (M) solution.
- The Mass term in the denominator of the molality expression is the Mass of solvent

Solution Stoichiometry

- (Mole – Mass) Conversions involving Solutions

Calculating the Mass of a substance given the Volume and Molarity

Ex. How many grams of Sodium Hydrogen Phosphate (Na_2HPO_4) are in 1.75 L of a 0.460 M solution?

Moles Na_2HPO_4 = 1.75 $\cancel{\text{L}}$ x 0.460 mol Na_2HPO_4/1 $\cancel{\text{L}}$ soln

= 0.805 mol Na_2HPO_4

Mass Na_2HPO_4 = 0.805 $\cancel{\text{mol}}$ x 141.96 g Na_2HPO_4/$\cancel{\text{mol Na}_2\text{HPO}_4}$

= 114. g

Practice Problem

- Calculate the volume of a 3.30 M Sucrose solution containing 135 g of solute.

(FW Sucrose – 342.30 g/mol)

Ans:

<u>moles solute</u>

135 g sucrose x 1 mol sucrose / 342.30 g sucrose = 0.3944 mol

<u>Vol soln</u>

0.3944 mol sucrose x 1.00 L solution/3.30 mol sucrose = 0.120 L

Dilution

- The amount of solute in a solution is the same after the solution is diluted with additional solvent
- Dilution problems utilize the following relationship between the molarity (M) and volume (V)

$$\mathbf{M_{dil} \times V_{dil} \;=\; number\ of\ moles \;=\; M_{conc} \times V_{conc}}$$

Practice Problem

- Calculate the Molarity of the solution prepared by diluting 37.00 mL of 0.250 M Potassium Chloride (KCl) to 150.00 mL.

Ans: Dilution problem ($M_1V_1 = M_2V_2$)

$M_1 = 0.250$ M KCl $\qquad V_1 = 37.00$ mL

$M_2 = ?$ $\qquad V_2 = 150.00$ mL

$M_1V_1 = M_2V_2 \quad \therefore \quad M_2 = M_1V_1 / V_2$

$M_2 = (0.250\text{ M}) \times 37.00\text{ mL}) / 150.0\text{ mL}$

$= 0.0617$ M

Practice Problem

- How many liters (L) of stomach acid (0.10 M HCl) react with (neutralize) 0.10 grams (g) of Magnesium Hydroxide (antacid)

$$\mathbf{Mg(OH)_2(s) \;+\; 2HCl(aq) \;\rightarrow\; MgCl_2(aq) \;+\; 2H_2O(l)}$$

- Convert mass (g) of $Mg(OH)_2$ to moles

$$\text{mol } Mg(OH)_2 = 0.1 \text{ g } Mg(OH)_2 \times \frac{1 \text{ mol } Mg(OH)_2}{58.33 \text{ g } Mg(OH)_2} = 1.7 \times 10^{-3} \text{ mol}$$

- Convert from moles of $Mg(OH)_2$ to moles of HCl

$$\text{mol HCl} = 1.7 \times 10^{-3} \text{ mol } Mg(OH)_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol } Mg(OH)_2} = 3.4 \times 10^{-3} \text{ mol HCl}$$

- Convert moles HCl to volume (L)

$$\text{Vol (L) HCL} = 3.4 \times 10^{-3} \text{ mol HCl} \times \frac{1 \text{L}}{0.10 \text{ mol HCl}} = 3.4 \times 10^{-2} \text{ L}$$

Equation summary

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}} \qquad M = \frac{\text{mol solute}}{\text{L soln}}$$

$$\text{Molality} = \frac{\text{Moles Solute}}{\text{Kilogram Solvent}} \qquad m = \frac{\text{mol solute}}{\text{Kg solvent}}$$

$$M_{dil} \times V_{dil} = \text{number of moles} = M_{conc} \times V_{conc}$$