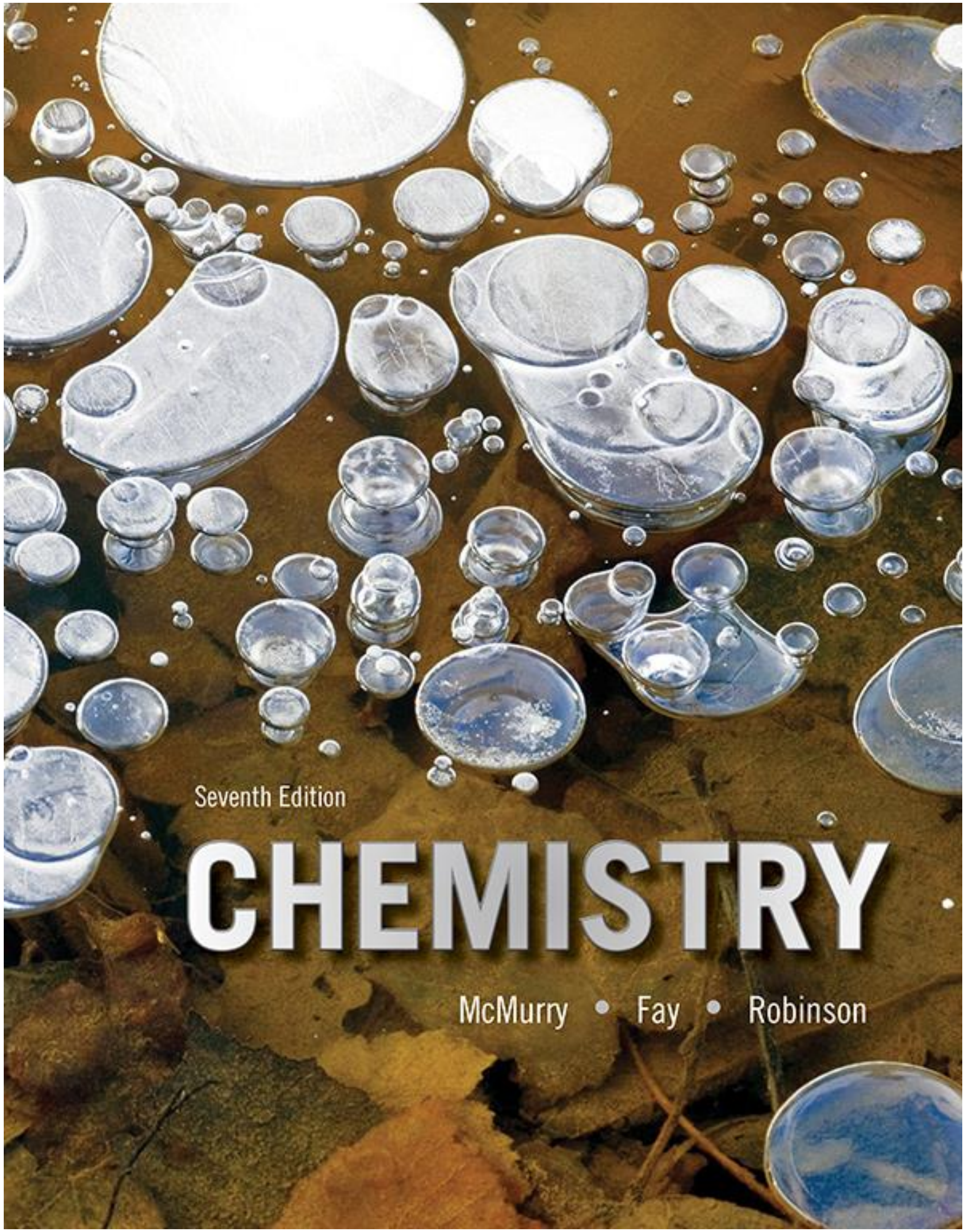

Lecture Presentation

Chapter 3

Mass Relationships in Chemical Reactions

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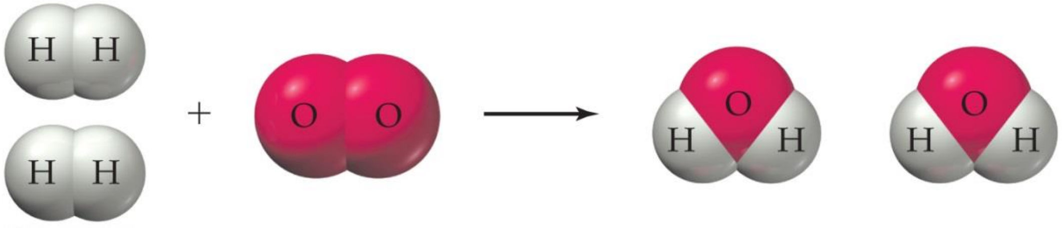
Instructor's Resource Materials (Download only) for *Chemistry,* 7e
John E. McMurry, Robert C. Fay, Jill Robinson

Representing Chemistry on Different Levels

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$

Microscopic: 2 molecules of hydrogen gas react with 1 molecule of oxygen gas to yield 2 molecules of liquid water.



Representing Chemistry on Different Levels

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$

Microscopic: 2 *molecules* of hydrogen gas react with 1 *molecule* of oxygen gas to yield 2 *molecules* of liquid water.

Macroscopic: 2 *moles* of hydrogen gas react with 1 *mole* of oxygen gas to yield 2 *moles* of liquid water.

Balancing Chemical Equations

A balanced chemical equation shows that the **law of conservation of mass** is adhered to.

In a **balanced** chemical equation, the numbers and kinds of atoms on both sides of the reaction arrow are identical.

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

left side:	right side:
2 Na	2 Na
2 Cl	2 Cl

Balancing Chemical Equations

$$Hg(NO_3)_2(aq) + 2KI(aq) \longrightarrow HgI_2(s) + 2KNO_3(aq)$$

left side:	right side:
1 Hg	1 Hg
2 N	2 I
6 O	2 K
2 K	2 N
2 I	6 O

Balancing Chemical Equations

1. Write the unbalanced equation using the correct chemical formula for each reactant and product.

$$H_2(g) + O_2(g) \longrightarrow H_2O(l)$$

2. Find suitable coefficients—the numbers placed before formulas to indicate how many formula units of each substance are required to balance the equation.

$$\color{red}{2}H_2(g) + O_2(g) \longrightarrow \color{red}{2}H_2O(l)$$

Balancing Chemical Equations

3. Reduce the coefficients to their smallest whole-number values, if necessary, by dividing them all by a common divisor.

$$4H_2(g) + 2O_2(g) \longrightarrow 4H_2O(l)$$

$$\downarrow$$

divide all by 2

$$\downarrow$$

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$

Balancing Chemical Equations

4. Check your answer by making sure that the numbers and kinds of atoms are the same on both sides of the equation.

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$

left side:	right side:
4 H	4 H
2 O	2 O

Chemical Arithmetic: Stoichiometry

Molecular Mass: Sum of atomic masses of all atoms in a molecule

Formula Mass: Sum of atomic masses of all atoms in a formula unit of any compound, molecular or ionic

HCl: 1.0 amu + 35.5 amu = 36.5 amu

C_2H_4: 2(12.0 amu) + 4(1.0 amu) = 28.0 amu

Chemical Arithmetic: Stoichiometry

HCl: 1 mole = 36.5 g

6.022×10^{23} molecules = 36.5 g

C_2H_4: 1 mole = 28.0 g

6.022×10^{23} molecules = 28.0 g

Chemical Arithmetic: Stoichiometry

Stoichiometry: The chemical arithmetic needed for mole-mass conversions

$$a\mathrm{A} + b\mathrm{B} \longrightarrow c\mathrm{C} + d\mathrm{D}$$

Grams of A → Moles of A → Moles of B → Grams of B

Grams of A → Moles of A	Moles of A → Moles of B	Moles of B → Grams of B
Molar mass of A	Mole ratio between A and B (coefficients)	Molar mass of B

Chemical Arithmetic: Stoichiometry

$$2NaOH(aq) + Cl_2(g) \longrightarrow NaOCl(aq) + NaCl(aq) + H_2O(l)$$

How many grams of NaOH are needed to react with 25.0 g Cl_2?

Grams of Cl_2 → Moles of Cl_2 → Moles of NaOH → Grams of NaOH

Molar mass of Cl_2

Mole ratio between Cl_2 and NaOH

Molar mass of NaOH

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Chemical Arithmetic: Stoichiometry

$$2NaOH(aq) + Cl_2(g) \longrightarrow NaOCl(aq) + NaCl(aq) + H_2O(l)$$

How many grams of NaOH are needed to react with 25.0 g Cl_2?

$$25.0 \text{ g } Cl_2 \times \frac{1 \text{ mol } Cl_2}{70.9 \text{ g } Cl_2} \times \frac{2 \text{ mol NaOH}}{1 \text{ mol } Cl_2} \times \frac{40.0 \text{ g NaOH}}{1 \text{ mol NaOH}}$$

$$= 28.2 \text{ g NaOH}$$

Yields of Chemical Reactions

Actual Yield: The amount actually formed in a reaction

Theoretical Yield: The amount predicted by calculations

$$\textbf{Percent Yield} = \frac{\text{Actual yield of product}}{\text{Theoretical yield of product}} \times 100\%$$

Reactions with Limiting Amounts of Reactants

Limiting Reactant: The reactant that is present in limiting amount. The extent to which a chemical reaction takes place depends on the limiting reactant.

Excess Reactant: Any of the other reactants still present after determination of the limiting reactant

Reactions with Limiting Amounts of Reactants

At a high temperature, ethylene oxide reacts with water to form ethylene glycol, which is an automobile antifreeze and a starting material in the preparation of polyester polymers:

$$C_2H_4O(aq) \quad + \quad H_2O(l) \longrightarrow C_2H_6O_2(l)$$

Because water is so cheap and abundant, it is used in excess when compared to ethylene oxide. This ensures that all of the relatively expensive ethylene oxide is entirely consumed.

Reactions with Limiting Amounts of Reactants

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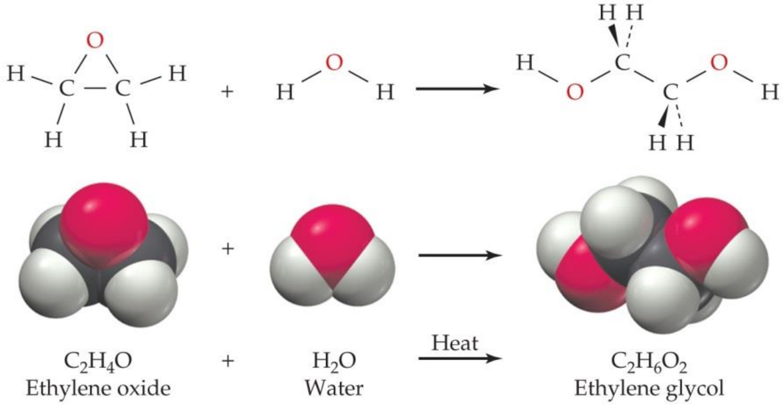
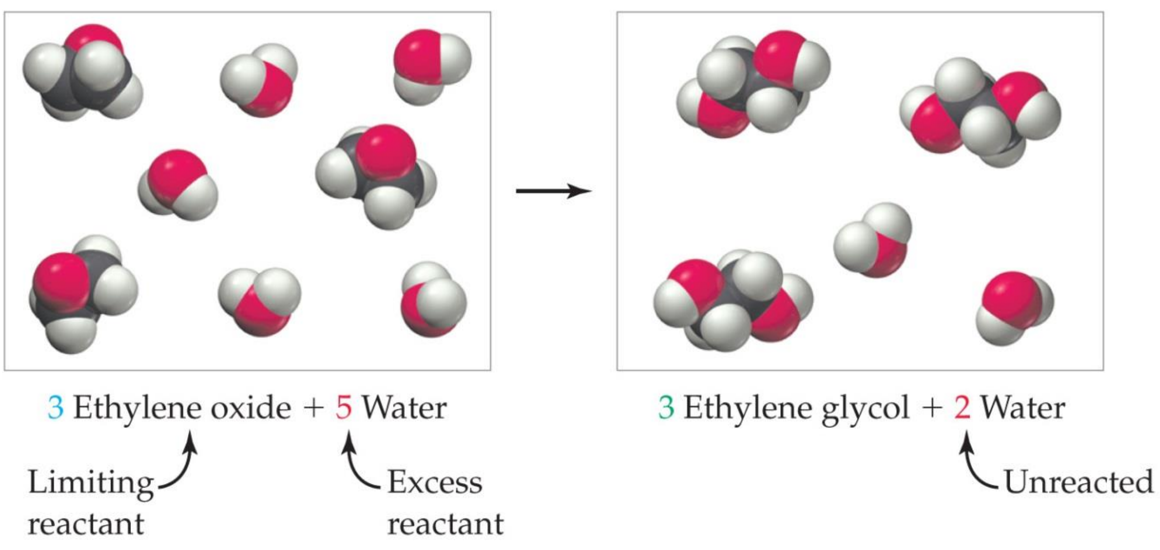


If 3 mol of ethylene oxide react with 5 mol of water, which reactant is limiting and which reactant is present in excess?

Reactions with Limiting Amounts of Reactants

At a high temperature, ethylene oxide reacts with water to form ethylene glycol, which is an automobile antifreeze and a starting material in the preparation of polyester polymers:

$$C_2H_4O(aq) \quad + \quad H_2O(l) \longrightarrow C_2H_6O_2(l)$$



Reactions with Limiting Amounts of Reactants

Lithium oxide is used aboard the space shuttle to remove water from the air supply according to the equation

$$Li_2O(s) + H_2O(g) \longrightarrow 2LiOH(s)$$

If 80.0 g of water is to be removed and 65.0 g of Li_2O is available, which reactant is limiting? How many grams of excess reactant remain? How many grams of LiOH are produced?

Reactions with Limiting Amounts of Reactants

$$Li_2O(s) + H_2O(g) \longrightarrow 2LiOH(s)$$

Which reactant is limiting?

Amount of H_2O that will react with 65.0 g Li_2O:

$$65.0\ \text{g}\ Li_2O \times \frac{1\ \text{mol}\ Li_2O}{29.9\ \text{g}\ Li_2O} \times \frac{1\ \text{mol}\ H_2O}{1\ \text{mol}\ Li_2O} = \boxed{2.17\ \text{mol}\ H_2O}$$

Amount of H_2O given:

$$80.0\ \text{g}\ H_2O \times \frac{1\ \text{mol}\ H_2O}{18.0\ \text{g}\ H_2O} = 4.44\ \text{mol}\ H_2O$$

Li_2O is limiting.

Reactions with Limiting Amounts of Reactants

$$Li_2O(s) + H_2O(g) \longrightarrow 2LiOH(s)$$

How many grams of excess H_2O remain?

$$2.17 \text{ mol } H_2O \times \frac{18.0 \text{ g } H_2O}{1 \text{ mol } H_2O} = 39.1 \text{ g } H_2O \text{ (consumed)}$$

$$\underset{\text{initial}}{80.0 \text{ g } H_2O} - \underset{\text{consumed}}{39.1 \text{ g } H_2O} = \boxed{\underset{\textbf{remaining}}{40.9 \text{ g } H_2O}}$$

Reactions with Limiting Amounts of Reactants

$$Li_2O(s) + H_2O(g) \longrightarrow 2LiOH(s)$$

How many grams of LiOH are produced?

$$2.17 \text{ mol } H_2O \times \frac{2 \text{ mol LiOH}}{1 \text{ mol } H_2O} \times \frac{23.9 \text{ g LiOH}}{1 \text{ mol LiOH}} = \boxed{104 \text{ g LiOH}}$$

Percent Composition and Empirical Formulas

Percent Composition: Expressed by identifying the elements present and giving the mass percent of each

Empirical Formula: It tells the smallest whole-number ratios of the atoms in the compound.

Molecular Formula: It tells the actual numbers of atoms in a compound. It can be either the empirical formula or a multiple of it.

$$\text{Multiple} = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

Percent Composition and Empirical Formulas

A colorless liquid has a composition of 84.1% carbon and 15.9% hydrogen by mass. Determine the empirical formula. Also, assuming the molar mass of this compound is 114.2 g/mol, determine the molecular formula of this compound.

$$\underbrace{\text{Mass percents} \longrightarrow \text{Moles}}_{\text{Molar masses}} \longrightarrow \underbrace{\text{Mole ratios} \longrightarrow \text{Subscripts}}_{\text{Relative mole ratios}}$$

Percent Composition and Empirical Formulas

Assume 100.0 g of the substance:

Mole of carbon:

$$84.1\text{ g C} \times \frac{1\text{ mol C}}{12.0\text{ g C}} = 7.01\text{ mol C}$$

Mole of hydrogen:

$$15.9\text{ g H} \times \frac{1\text{ mol H}}{1.0\text{ g H}} = 15.9\text{ mol H}$$

Percent Composition and Empirical Formulas

Empirical formula:

$$C_{7.01}H_{15.9} \qquad \frac{C_{7.01}}{7.01}\frac{H_{15.9}}{7.01} = C_1H_{2.27}$$

smallest value for the ratio

$$C_1H_{2.27} \qquad C_{1\times 4}H_{2.27\times 4} = \boxed{C_4H_9}$$

need whole numbers

Molecular formula:

$$\text{multiple} = \frac{114.2}{57.0} = 2 \qquad C_{4\times 2}H_{9\times 2} = \boxed{C_8H_{18}}$$

Determining Empirical Formulas: Elemental Analysis

Combustion Analysis: A compound of unknown composition is burned with oxygen to produce the volatile combustion products CO_2 and H_2O, which are separated and have their amounts determined by an automated instrument.

$$\underbrace{\text{hydrocarbon}}_{\substack{\text{carbon} \\ \text{hydrogen}}} + O_2(g) \longrightarrow xCO_2(g) + yH_2O(g)$$

Determining Molecular Masses: Mass Spectrometry