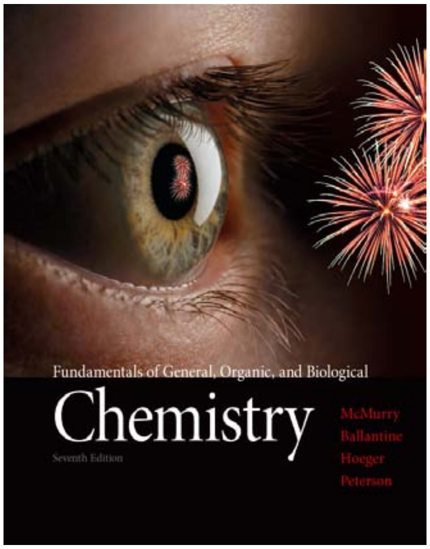
Chapter 6 Lecture



Fundamentals of General, Organic, and Biological Chemistry

7th Edition

McMurry, Ballantine, Hoeger, Peterson

Chapter Six

Chemical Reactions:
Mole and Mass Relationships

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Outline

6.1 The Mole and Avogadro's Number

6.2 Gram–Mole Conversions

6.3 Mole Relationships and Chemical Equations

6.4 Mass Relationships and Chemical Equations

6.5 Limiting Reagent and Percent Yield

Goals

1. **What is the mole, and why is it useful in chemistry?**

   Be able to explain the meaning and uses of the mole and Avogadro's number.

2. **How are molar quantities and mass quantities related?**

   Be able to convert between molar and mass quantities of an element or compound.

3. **What are the limiting reagent, theoretical yield, and percent yield of a reaction?**

   Be able to take the amount of product actually formed in a reaction, calculate the amount that could form theoretically, and express the results as a percent yield.

6.1 The Mole and Avogadro's Number

- Balanced chemical equations indicate what is happening at the molecular level during a reaction.
- To obtain the correct ratio of reactant molecules, the reactants must be weighed.
- To determine how many molecules of a given substance are in a certain mass, it is helpful to define a quantity called *molecular weight*.

6.1 The Mole and Avogadro's Number

- Atomic weight is the average mass of an element's *atoms*.
- **Molecular weight (MW)** is the average mass of a substance's *molecules*.
- A substance's molecular weight (or **formula weight** for an ionic compound) is the sum of the atomic weights for all the atoms in the molecule or formula unit.

6.1 The Mole and Avogadro's Number

- Molecular weights can be used to provide mass ratios for reactants.
- Samples of different substances always contain the same number of molecules or formula units whenever their mass ratio is the same as their molecular or formula weight ratio.

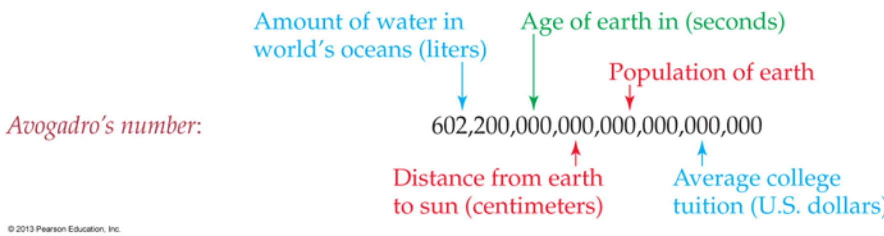
6.1 The Mole and Avogadro's Number

- The mass/number relationship for molecules can be used to measure amounts in grams equal to molecular weights.
- This will always provide a 1-to-1 ratio of reactant molecules.

6.1 The Mole and Avogadro's Number

- A **mole** is the amount of a substance whose mass in grams is numerically equal to its molecular or formula weight.
- One mole of any substance contains 6.022 × 10^{23} formula units.
- This value is called **Avogadro's number** (abbreviated N_A) after the Italian scientist who first recognized the importance of the mass/number relationship in molecules.

6.1 The Mole and Avogadro's Number



6.2 Gram–Mole Conversions

- **Molar mass** is the mass, in grams, of 1 mol of a substance, numerically equal to molecular weight.
- Molar mass serves as a conversion factor between numbers of moles and mass.

Molar mass used as conversion factor

$$27\ \cancel{g\ H_2O} \times \frac{1\ mol\ H_2O}{18.0\ \cancel{g\ H_2O}} = 1.5\ mol\ H_2O$$

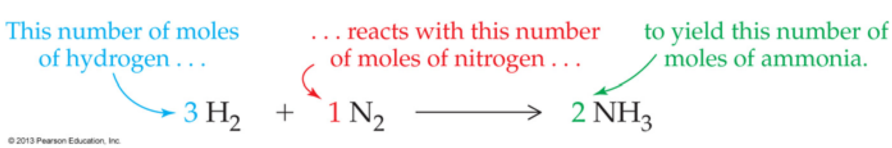
6.2 Gram–Mole Conversions

Did Ben Franklin Have Avogadro's Number?
A Ballpark Calculation

- His measurement of the extent to which oil spreads on water makes possible a simple estimate of molecular size and Avogadro's number.
  - Dividing the original volume of oil by the area it covers provides an estimate of molecular size.
  - Volume is equal to the surface area times the length of one side of one molecule (height of the layer, assuming it is only one molecule thick, and that the molecules are cubes). This formula provides an estimate of molecular size.
  - The number of molecules can be determined by dividing the total surface area by the surface area of one molecule.
  - Using an average density and molecular weight for oil, a rough estimate of the number of moles can be made.
  - The number of molecules divided by the number of moles will approximate Avogadro's number.

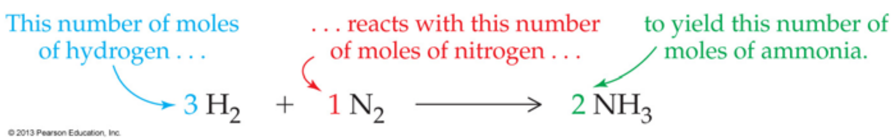
6.3 Mole Relationships and Chemical Equations

- In chemical reactions, the unit to specify the relationship between reactants and products is the mole.
- Coefficients in a chemical equation tell how many *molecules*, and thus how many *moles*, of each reactant are needed and of each product are formed.



6.3 Mole Relationships and Chemical Equations

- Coefficients can be put in the form of *mole ratios*, which act as conversion factors when setting up factor-label calculations.



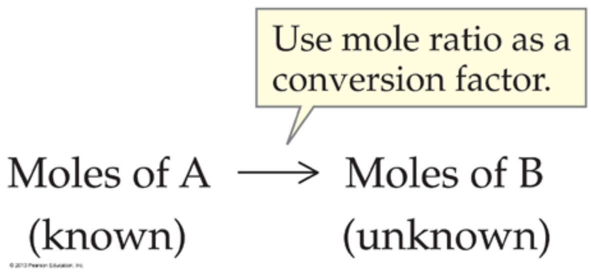
- Three moles of hydrogen produce 2 moles of ammonia.
- One mole of nitrogen produces 2 moles of ammonia.

6.4 Mass Relationships and Chemical Equations

- Coefficients in a balanced chemical equation represent molecule-to-molecule or mole-to-mole relationships between reactants and products.
- Actual amounts of substances used in the laboratory are weighed out in grams.
- Three types of conversions are needed when doing chemical arithmetic.

6.4 Mass Relationships and Chemical Equations

- **Mole to mole conversions** are carried out using *mole ratios* as conversion factors.



6.4 Mass Relationships and Chemical Equations

WORKED EXAMPLE 6.5

Rusting involves the reaction of iron with oxygen to form iron(III) oxide.

$$4\ Fe\ (s) + 3\ O_2\ (g) \rightarrow 2\ Fe_2O_3\ (s)$$

(a) What are the mole ratios of the product to each reactant and of the reactants to each other?

(b) How many moles of iron(III) oxide are formed by the complete oxidation of 6.2 mol of iron?

6.4 Mass Relationships and Chemical Equations

WORKED EXAMPLE 6.5 (Continued)

Analysis and Solution

(a) The coefficients of a balanced equation represent the mole ratios.

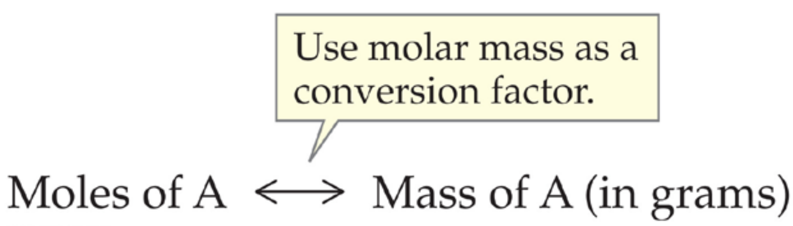
$$\frac{2\ \text{mol}\ Fe_2O_3}{4\ \text{mol Fe}} \quad \frac{2\ \text{mol}\ Fe_2O_3}{3\ \text{mol}\ O_2} \quad \frac{4\ \text{mol Fe}}{3\ \text{mol}\ O_2}$$

(b) To find how many moles of iron oxide are formed, write down the known information—6.2 mol of iron—and select the mole ratio that allows the quantities to cancel, leaving the desired quantity.

$$6.2\ \text{mol Fe} \times \frac{2\ \text{mol}\ Fe_2O_3}{4\ \text{mol Fe}} = 3.1\ \text{mol}\ Fe_2O_3$$

6.4 Mass Relationships and Chemical Equations

- **Mole-to-mass and mass-to-mole conversions** are carried out using *molar mass* as a conversion factor.



6.4 Mass Relationships and Chemical Equations

WORKED EXAMPLE 6.3
Molar Mass: Mole to Gram Conversion

The nonprescription pain relievers Advil and Nuprin contain ibuprofen, whose molecular weight is 206.3 amu. If all the tablets in a bottle of pain reliever together contain 0.082 mol of ibuprofen, what is the number of grams of ibuprofen in the bottle?

Analysis—We are given a number of moles and asked to find the mass. Molar mass is the conversion factor between the two.

Ballpark Estimate—Since 1 mol of ibuprofen has a mass of about 200 g, 0.08 mol has a mass of about 0.08 × 200 = 16 g.

6.4 Mass Relationships and Chemical Equations

WORKED EXAMPLE 6.3
Molar Mass: Mole to Gram Conversion (Continued)

SOLUTION

STEP 1: **Identify known information.**	0.082 mol ibuprofen in bottle
STEP 2: **Identify answer and units.**	mass ibuprofen in bottle = ?? g
STEP 3: **Identify conversion factor.** We use the molecular weight of ibuprofen to convert from moles to grams.	1 mol ibuprofen = 206.3 g $\frac{206.3 \text{ g ibuprofen}}{1 \text{ mol ibuprofen}}$
STEP 4: **Solve.** Set up an equation using the known information and conversion factor so that unwanted units cancel.	$0.082 \text{ mol } C_{13}H_{18}O_2 \times \frac{206.3 \text{ g ibuprofen}}{1 \text{ mol ibuprofen}} = 17 \text{ g } C_{13}H_{18}O_2$

6.4 Mass Relationships and Chemical Equations

WORKED EXAMPLE 6.4 Molar Mass: Gram to Mole Conversion

The maximum dose of sodium hydrogen phosphate (Na_2HPO_4, MW = 142.0 g/mol) that should be taken in one day for use as a laxative is 3.8 g. How many moles of sodium hydrogen phosphate, how many moles of Na^+ ions, and how many total moles of ions are in this dose?

Analysis—Molar mass is the conversion factor between mass and number of moles. The chemical formula shows that each formula unit contains 2 Na^+ ions and 1 HPO_4^{2-} ion.

Ballpark Estimate—The maximum dose is about two orders of magnitude smaller than the molecular weight (about 4 g compared to 142 g). The number of moles of sodium hydrogen phosphate in 3.8 g should be about two orders of magnitude less than one mole. The number of moles of Na_2HPO_4 and total moles of ions, then, should be on the order of 10^{-2}.

6.4 Mass Relationships and Chemical Equations

WORKED EXAMPLE 6.4 Molar Mass: Gram to Mole Conversion

STEP 1:Identify known information. We are given the mass and molecular weight of Na_2HPO_4.

3.8 g Na_2HPO_4
MW = 142.0 g/mol

STEP 2:Identify answer and units. We need to find the number of moles of Na_2HPO_4, and the total number of moles of ions.

Moles of Na_2HPO_4 = ?? mol
Moles of Na^+ ions = ?? mol
Total moles of ions = ?? Mo

STEP 3:Identify conversion factor. We can use the molecular weight of Na_2HPO_4 to convert from grams to moles.

$$\frac{1 \text{ mol } Na_2HPO_4}{142.0 \text{ g } Na_2HPO_4}$$

6.4 Mass Relationships and Chemical Equations

WORKED EXAMPLE 6.4 Molar Mass: Gram to Mole Conversion

STEP 4: Solve.

We use the known information and conversion factor to obtain moles of Na_2HPO_4; since 1 mol of Na_2HPO_4 contains 2 mol of Na^+ ions and 1 mol of HPO^{2-} ions, we multiply these values by the number of moles in the sample.

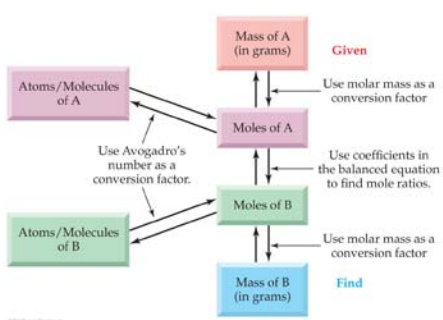
$$3.8\ \text{g}\ Na_2HPO_4 \times \frac{1\ \text{mol}\ Na_2HPO_4}{140.0\ \text{g}\ Na_2HPO_4} = 0.027\ \text{mol}\ Na_2HPO_4$$

$$\frac{2\ \text{mol}\ Na^+}{\text{mol}\ Na_2HPO_4} \times 0.027\ \text{mol}\ Na_2HPO_4 = 0.054\ \text{mol}\ Na^+$$

$$\frac{3\ \text{mol ions}}{\text{mol}\ Na_2HPO_4} \times 0.027\ \text{mol}\ Na_2HPO_4 = 0.081\ \text{mol ions}$$

6.4 Mass Relationships and Chemical Equations

- **Mass to mass conversions** cannot be carried out directly. If you know the mass of A and need to find the mass of B, first convert the mass of A into moles of A, then carry out a mole to mole conversion to find moles of B, then convert moles of B into the mass of B.



6.4 Mass Relationships and Chemical Equations

There are four steps for determining mass relationships among reactants and products.

STEP 1: Write the balanced chemical equation.

STEP 2: Choose molar masses and mole ratios to convert the known information into the needed information.

STEP 3: Set up the factor-label expression and calculate the answer.

STEP 4: Check the answer against the ballpark estimate you made before you began your calculations.

6.5 Limiting Reagent and Percent Yield

- Only rarely are all reactants converted to products.
- When running a chemical reaction, we don't always have the exact amounts of reagents to allow all of them to react completely.
- The **limiting reagent** is the reactant that runs out first.

6.5 Limiting Reagent and Percent Yield

- **Theoretical yield** is the amount of product formed assuming complete reaction of the limiting reagent.
- Chemical reactions do not always yield the exact amount predicted. A majority of reactant molecules behave as written but other processes, called *side reactions*, also occur.
- **Actual yield** is the amount of product actually formed in a reaction.

6.5 Limiting Reagent and Percent Yield

- **Percent yield** is the percent of the theoretical yield actually obtained from a chemical reaction.

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

- Actual yield is found by weighing the amount of product obtained.
- Theoretical yield is found by using the amount of limiting reagent in a mass-to-mass calculation.

6.4 Mass Relationships and Chemical Equations

WORKED EXAMPLE 6.8 Percent Yield

The combustion of acetylene gas (C_2H_2) produces carbon dioxide and water as indicated in the following reaction.

$$2\ C_2H_2\ (g) + 5\ O_2\ (g) \rightarrow 4\ CO_2\ (g) + 2\ H_2O\ (g)$$

When 26.0 g of acetylene is burned in sufficient oxygen for complete reaction, the theoretical yield of CO_2 is 88.0 g. Calculate the percent yield for this reaction if the actual yield is only 72.4 g CO_2.

Analysis—The percent yield is calculated by dividing the actual yield by the theoretical yield and multiplying by 100.

Ballpark Estimate—The theoretical yield (88.0 g) is close to 100 g. The actual yield (72.4 g) is about 15 g less than the theoretical yield. The actual yield is thus, about 15% less than the theoretical yield, so the percent yield is about 85%.

6.4 Mass Relationships and Chemical Equations

WORKED EXAMPLE 6.8 Percent Yield (Continued)

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

$$\frac{72.4\ \text{g}\ CO_2}{88.0\ \text{g}\ CO_2} \times 100 = 82.3\%$$

Ballpark Check—The calculated percent yield agrees very well with our estimate of 85%.

6.4 Mass Relationships and Chemical Equations

WORKED EXAMPLE 6.9 Mass to Mole Conversions: Limiting Reagent and Theoretical Yield

The element boron is produced commercially by the reaction of boric oxide with magnesium at high temperature.

$$B_2O_3\,(l) + 3\,Mg\,(s) \rightarrow 2\,B\,(s) + 3\,MgO\,(s)$$

What is the theoretical yield of boron when 2350 g of boric oxide is reacted with 3580 g of magnesium? The molar masses of boric oxide and magnesium are 69.6 g/mol and 24.3 g/mol, respectively.

Analysis—To calculate theoretical yield, we first have to identify the limiting reagent. The theoretical yield in grams is then calculated from the amount of limiting reagent used in the reaction.

6.4 Mass Relationships and Chemical Equations

WORKED EXAMPLE 6.9 Mass to Mole Conversions: Limiting Reagent and Theoretical Yield (Continued)

SOLUTION

STEP 1: **Identify known information.** We have the masses and molar masses of the reagents.

2350 g B_2O_3, molar mass 69.6 g/mol
3580 g Mg, molar mass 24.3 g/mol

STEP 2: **Identify answer and units.** We are solving for the theoretical yield of boron.

Theoretical mass of B = ?? g

STEP 3: **Identify conversion factors.** We can use the molar masses to convert from masses to moles of reactants (B_2O_3, Mg). From moles of reactants, we can use mole ratios from the balanced chemical equation to find the number of moles of B produced, assuming complete conversion of a given reactant. B_2O_3 is the limiting reagent, since complete conversion of this reagent yields less product (67.6 mol B formed) than does complete conversion of Mg (98.0 mol B formed).

$$(2350\ \text{g } B_2O_3) \times \frac{1\ \text{mol } B_2O_3}{69.6\ \text{g } B_2O_3} = 33.8\ \text{mol } B_2O_3$$

$$(3580\ \text{g Mg}) \times \frac{1\ \text{mol Mg}}{24.3\ \text{g Mg}} = 147\ \text{mol Mg}$$

$$33.8\ \text{mol } B_2O_3 \times \frac{2\ \text{mol B}}{1\ \text{mol } B_2O_3} = 67.6\ \text{mol B}^*$$

$$147\ \text{mol Mg} \times \frac{2\ \text{mol B}}{3\ \text{mol Mg}} = 98.0\ \text{mol B}$$

(*B_2O_3 is the limiting reagent because it yields fewer moles of B!)

STEP 4: **Solve.** Once the limiting reagent has been identified (B_2O_3), the theoretical amount of B that should be formed can be calculated using a mole to mass conversion.

$$67.6\ \text{mol B} \times \frac{10.8\ \text{g B}}{1\ \text{mol B}} = 730\ \text{g B}$$

6.4 Mass Relationships and Chemical Equations

WORKED EXAMPLE 6.10

Mass to Mole Conversion: Percent Yield

The reaction of ethylene with water to give ethyl alcohol (CH_3CH_2OH) occurs in 78.5% actual yield. How many grams of ethyl alcohol are formed by reaction of 25.0 g of ethylene? (For ethylene, MW = 28.0 amu for ethyl alcohol, MW = 46.0 amu.)

$$H_2C{=}CH_2 + H_2O \rightarrow CH_3CH_2OH$$

Analysis—Treat this as a typical mass relationship problem to find the amount of ethyl alcohol that can theoretically be formed from 25.0 g of ethylene, and then multiply the answer by 78.5% to find the amount actually formed.

Ballpark Estimate—The 25.0 g of ethylene is a bit less than 1 mol; since the percent yield is about 78%, a bit less than 0.78 mol of ethyl alcohol will form—perhaps about 3/4 mol, or ¾ × 46 g = 34 g

6.4 Mass Relationships and Chemical Equations

WORKED EXAMPLE 6.10 (Continued)

Solution: The theoretical yield of ethyl alcohol is

$$25.0\text{ g ethylene} \times \frac{1\text{ mol ethylene}}{28\text{ g ethylene}} \times$$

$$\frac{1\text{ mol ethyl alc.}}{1\text{ mol ethylene}} \times \frac{46.0\text{ g ethyl alc.}}{1\text{ mol ethyl alc.}} = 41.1\text{ g ethyl alc.}$$

So, the actual yield is as follows.

$$41.1\text{ g ethyl alc.} \times 0.785 = 32.3\text{ g ethyl alcohol}$$

Ballpark Check—The calculated result (32.3 g) is close to our estimate (34 g).

6.5 Limiting Reagent and Percent Yield

Anemia – A Limiting Reagent Problem?

- Anemia is the most commonly diagnosed blood disorder.
- The most common cause is insufficient dietary intake or absorption of iron.
- Hemoglobin (abbreviated Hb), the iron-containing protein found in red blood cells, is responsible for oxygen transport throughout the body.
- Low iron levels in the body result in decreased production and incorporation of Hb in red blood cells.
- In the United States, nearly 20% of women of child-bearing age suffer from anemia compared to only 2% of adult men.
- One way to ensure sufficient iron intake is a well-balanced diet. Vitamin C increases the absorption of iron by the body.
- The simplest way to increase dietary iron may be to use cast iron cookware. The iron content of many foods increases when cooked in an iron pot.

Chapter Summary

1. What is the mole, and why is it useful in chemistry?

- A *mole* refers to *Avogadro's number* 6.022×10^{23} formula units of a substance.
- One mole of any substance has a mass (a *molar mass*) equal to the molecular or formula weight of the substance in grams.
- Because equal numbers of moles contain equal numbers of formula units, molar masses act as conversion factors between numbers of molecules and masses in grams.

Chapter Summary, *Continued*

2. How are molar quantities and mass quantities related?

- The coefficients in a balanced chemical equation represent the numbers of moles of reactants and products in a reaction.
- The ratios of coefficients act as *mole ratios* that relate amounts of reactants and/or products.
- By using molar masses and mole ratios in factor-label calculations, unknown masses or molar amounts can be found from known masses or molar amounts.

Chapter Summary, *Continued*

3. What are the limiting reagent, theoretical yield, and percent yield of a reaction?

- The *limiting reagent* is the reactant that runs out first.
- The *theoretical yield* is the amount of product that would be formed based on the amount of the limiting reagent.
- The *actual yield* of a reaction is the amount of product obtained.
- The *percent yield* is the amount of product obtained divided by the amount theoretically possible and multiplied by 100%.