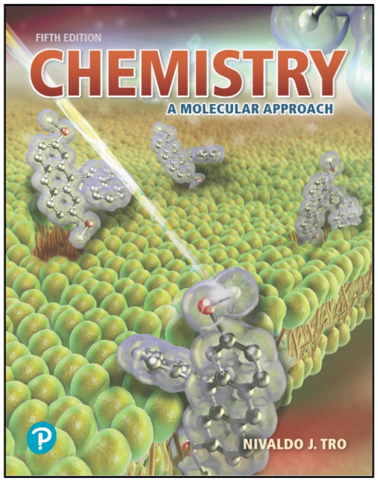
Chemistry: A Molecular Approach

Fifth Edition



Chapter 4

Chemical Reactions and Chemical Quantities

Pearson

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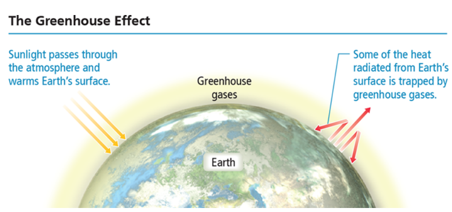
The Greenhouse Effect

- The **greenhouse gases** in the atmosphere
  - allow sunlight to enter the atmosphere.
  - warm Earth's surface.
  - prevent some of the heat generated by the sunlight from escaping.



  - The balance between incoming and outgoing energy from the sun determines Earth's average temperature.

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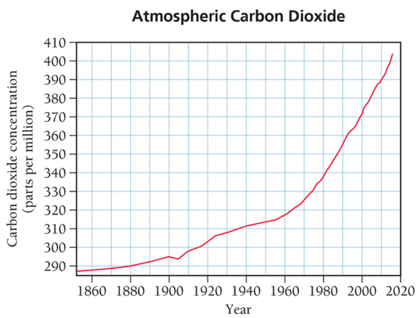
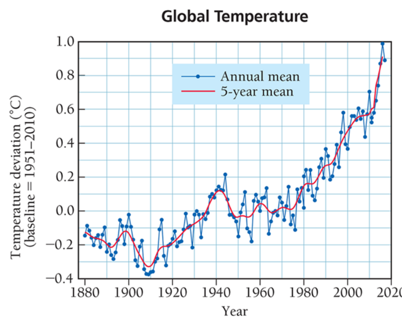
Global Warming (1 of 2)

- Scientists have measured an average 0.7°C rise in atmospheric temperature since 1860.
- During the same period, atmospheric CO_2 levels have risen 38%.
- Are the two trends causal?





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Global Warming (2 of 2)

- One source of CO_2 is the combustion reactions of fossil fuels we use to get energy.
- Another source of CO_2 is volcanic action.
  - How can we judge whether global warming is natural or due to our use of fossil fuels?

→

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Chemical Reaction

- **Chemical reaction**: a process in which one or more substances are converted into one or more different ones
  - Involve chemical changes in matter resulting in new chemical substances
- **Combustion reaction:** a particular type of chemical reaction in which a substance combines with oxygen to form one or more oxygen-containing compounds
  - Also emit heat

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Chemical Equations

- Shorthand way of describing a reaction

Reactants $\rightarrow$ **Products**

- Provide information about the reaction
  - Formulas of reactants and products
  - States of reactants and products
  - Relative numbers of reactant and product molecules that are required can be used to determine weights of reactants used and products that can be made

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States of Reactants and Products in Chemical Equations

Abbreviation	State
(*g*)	Gas
(*l*)	Liquid
(*s*)	Solid
(*aq*)	Aqueous (water solution)

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Combustion of Methane (1 of 4)

- Methane gas burns to produce carbon dioxide gas and gaseous water.
  - Whenever something burns it combines with $O_2(g)$.

$$CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

- If you look closely, you should immediately spot a problem.

$$CH_4(g) + \underbrace{O_2(g)}_{\textbf{2 O atoms}} \longrightarrow \underbrace{CO_2(g) + H_2O(g)}_{\text{2 O atoms + 1 O atom = } \textbf{3 O atoms}}$$

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Combustion of Methane (2 of 4)

- Notice also that the left side has four hydrogen atoms while the right side has only two.

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

4 H atoms

2 H atoms

- To correct these problems, we must **balance** the equation by changing the coefficients, not the subscripts.

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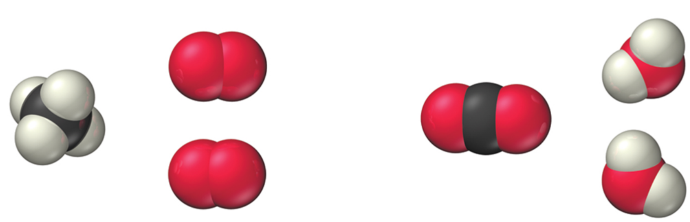
Combustion of Methane: Balanced

- To show that the reaction obeys the Law of Conservation of Mass, the equation must be **balanced**.
  - We adjust the numbers of molecules so there are equal numbers of atoms of each element on both sides of the arrow.

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$



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Conceptual Connection 4.1 (1 of 2)

How many oxygen atoms are on the right-hand side of the following chemical equation?

$$4\ FeCO_3(s) + O_2(g) \rightarrow 2\ Fe_2O_3(s) + 4\ CO_2(g)$$

a. 4 b. 5 c. 6 d. 14

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Conceptual Connection 4.1 (2 of 2)

How many oxygen atoms are on the right-hand side of the following chemical equation?

$$4\ FeCO_3(s) + O_2(g) \rightarrow 2\ Fe_2O_3(s) + 4\ CO_2(g)$$

a. 4 b. 5 c. 6 **d. 14**

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Conceptual Connection 4.2 (1 of 2)

Which quantity or quantities must always be the same on both sides of a chemical equation?

a. the number of atoms of each kind

b. the number of molecules of each kind

c. the number of moles of each kind of molecule

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Conceptual Connection 4.2 (2 of 2)

Which quantity or quantities must always be the same on both sides of a chemical equation?

a. the number of atoms of each kind

b. the number of molecules of each kind

c. the number of moles of each kind of molecule

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Reaction Stoichiometry: How Much Carbon Dioxide?

- The balanced chemical equations for fossil-fuel combustion reactions provide the exact relationships between the amount of fossil fuel burned and the amount of carbon dioxide emitted.

$$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$$

  - 16 CO_2 molecules are produced for every 2 molecules of octane burned.

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Quantities in Chemical Reactions

- The amount of every substance used and made in a chemical reaction is related to the amounts of all the other substances in the reaction.
  - Law of conservation of mass
  - Balancing equations by balancing atoms
- The study of the numerical relationship between chemical quantities in a chemical reaction is called **stoichiometry**.

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Reaction Stoichiometry

- The coefficients in a chemical reaction specify the relative amounts in moles of each of the substances involved in the reaction.

$$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$$

  – 2 molecules of C_8H_{18} react with 25 molecules of O_2 to form 16 molecules of CO_2 and 18 molecules of H_2O.
  – 2 moles of C_8H_{18} react with 25 moles of O_2 to form 16 moles of CO_2 and 18 moles of H_2O.

2 mol C_8H_{18} : 25 mol O_2 : 16 mol CO_2 : 18 mol H_2O

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Making Pizza

- The number of pizzas you can make depends on the amount of the ingredients you use.

1 crust + 5 oz tomato sauce + 2 cups cheese → 1 pizza

This relationship can be expressed mathematically.

1 crust : 5 oz sauce : 2 cups cheese : 1 pizza

- We can compare the amount of pizza that can be made from 10 cups of cheese:

2 cups cheese : 1 pizza, then,

$$10\ \cancel{\text{cups cheese}} \times \frac{1\ \text{pizza}}{2\ \cancel{\text{cups cheese}}} = 5\ \text{pizzas}$$

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Making Molecules: Mole-to-Mole Conversions

- We use the ratio from the balanced chemical equation in the same way that we used the ratio from the pizza recipe.

The ratio of the coefficients acts as a conversion factor between the amount in moles of the reactants and products.

$$\mathbf{2\ C_8H_{18}(\mathit{l}) + 25\ O_2(\mathit{g}) \rightarrow 16\ CO_2(\mathit{g}) + 18\ H_2O(\mathit{g})}$$

stoichiometric ratio: 2 moles C_8H_{18}: 16 moles CO_2

The ratio acts as a conversion factor between the amount in moles of the reactant, C_8H_{18}, and the amount in moles of the product, CO_2.

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Mole-to-Mole Conversions

- Suppose we burn 22.0 moles of C_8H_{18}; how many moles of CO_2 form?

$$\mathbf{2\ C_8H_{18}(\mathit{l}) + 25\ O_2(\mathit{g}) \rightarrow 16\ CO_2(\mathit{g}) + 18\ H_2O(\mathit{g})}$$

stoichiometric ratio: 2 moles C_8H_{18}: 16 moles CO_2

$$22.0\ \cancel{\text{mol } C_8H_{18}} \times \frac{16\text{ mol } CO_2}{2\ \cancel{\text{mol } C_8H_{18}}} = 176\text{ mol } CO_2$$

- The combustion of 22.0 moles of C_8H_{18} adds 176 moles of CO_2 to the atmosphere.

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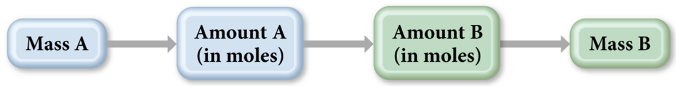
Making Molecules: Mass-to-Mass Conversions

- The world burned the equivalent of 3.7×10^{15} g of gasoline (octane) in 2013. We can estimate the mass of CO_2 produced based on the flowchart below.



- We use molar mass as a conversion factor between the mass given and amount in moles.
- We use coefficients as the conversion factor between the reactant, C_8H_{18}, and the amount in moles of the product, CO_2, and then molar mass as the conversion factor to get the mass of CO_2 produced.

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Conceptual Connection 4.3 (1 of 2)

Use the balanced equation for the combustion of octane to determine how many moles of H_2O are produced by the combustion of 22.0 moles of C_8H_{18}.

$$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$$

a. 18 moles H_2O

b. 22 moles H_2O

c. 176 moles H_2O

d. 198 moles H_2O

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Conceptual Connection 4.3 (2 of 2)

Use the balanced equation for the combustion of octane to determine how many moles of H_2O are produced by the combustion of 22.0 moles of C_8H_{18}.

$$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$$

a. 18 moles H_2O

b. 22 moles H_2O

c. 176 moles H_2O

d. 198 moles H_2O

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Mass-to-Mass Conversions

- If we burn 3.7×10^{15} g C_8H_{18}, how many grams of CO_2 form?

$$\mathbf{2\ C_8H_{18}(\mathit{l}) + 25\ O_2(\mathit{g}) \rightarrow 16\ CO_2(\mathit{g}) + 18\ H_2O(\mathit{g})}$$

molar mass: C_8H_{18} 114.22 g/mol, CO_2 44.01 g/mol

stoichiometric ratio: 2 moles C_8H_{18}: 16 moles CO_2

$$22.0\ \cancel{\text{mol } C_8H_{18}} \times \frac{16 \text{ mol } CO_2}{2\ \cancel{\text{mol } C_8H_{18}}} = 176 \text{ mol } CO_2$$

- The combustion 3.7×10^{15} g C_8H_{18} adds 1.1×10^{16}g CO_2 to the atmosphere.

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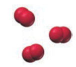
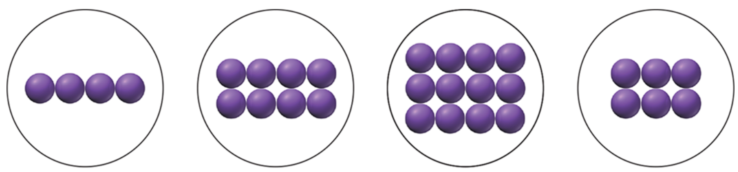
Conceptual Connection 4.4 (1 of 2)

Under certain conditions, sodium reacts with oxygen to form sodium oxide according to the reaction:

$$4\ Na(s) + O_2(g) \rightarrow 2\ Na_2O(s)$$



Which of the following images best represents the amount of sodium required to completely react with all of the oxygen in the above diagram according to the equation?



(a) (b) (c) (d)

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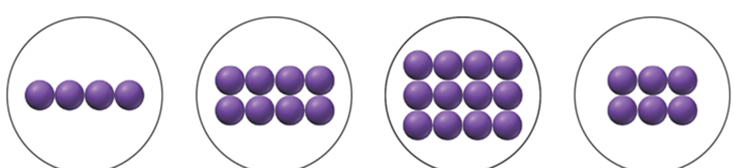
Conceptual Connection 4.4 (2 of 2)

Under certain conditions, sodium reacts with oxygen to form sodium oxide according to the reaction:

$$4\ Na(s) + O_2(g) \rightarrow 2\ Na_2O(s)$$

Which of the following images best represents the amount of sodium required to completely react with all of the oxygen in the above diagram according to the equation?



(a) (b) (c) (d)

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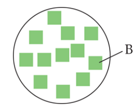
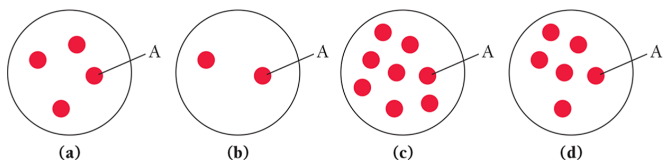
Conceptual Connection 4.5 (1 of 2)



Consider the generic chemical equation $A + 3B \rightarrow 2C$.

Let circles represent molecules of A and squares represent molecules of B. The diagram shown above represents the amount of B available for reaction. Which diagram in the answer options accurately represents the amount of A necessary to completely react with B?



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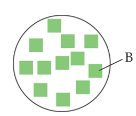
Conceptual Connection 4.5 (2 of 2)



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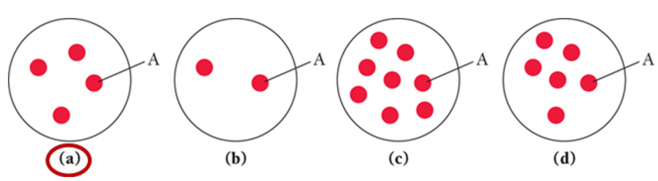

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Limiting Reactant, Theoretical Yield, Percent Yield

- Recall our pizza recipe:

$$\textbf{1 crust} + \textbf{5 oz tomato sauce} + \textbf{2 cups cheese} \rightarrow \textbf{1 pizza}$$

- If we have 4 crusts, 10 cups of cheese, and 15 oz tomato sauce, how many pizzas can we make?

We have enough crusts to make $4 \ \cancel{\text{crusts}} \times \dfrac{1 \text{ pizza}}{1 \ \cancel{\text{crust}}} = 4 \text{ pizzas}$

We have enough cheese to make $10 \ \cancel{\text{cups cheese}} \times \dfrac{1 \text{ pizza}}{2 \ \cancel{\text{cups cheese}}} = 5 \text{ pizzas}$

We have enough tomato sauce to make

$$15 \ \cancel{\text{ounces tomato sauce}} \times \frac{1 \text{ pizza}}{5 \ \cancel{\text{ounces tomato sauce}}} = 3 \text{ pizzas}$$

Limiting reactant

Smallest number of pizzas

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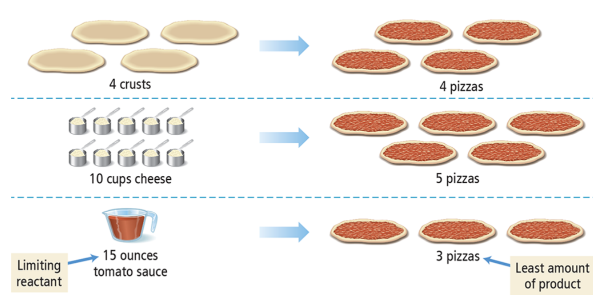
Limiting Reactant

- We have enough crusts for four pizzas, enough cheese for five pizzas, but enough tomato sauce for only three pizzas.
  - We can make only three pizzas. The tomato sauce **limits** how many pizzas we can make.



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Theoretical Yield

- Tomato sauce is the **limiting reactant**, the reactant that makes **the least amount of product.**
  – The limiting reactant is also known as the **limiting reagent**.
- The maximum number of pizzas we can make depends on this ingredient. In chemical reactions, we call this the **theoretical yield**.
  – This is the amount of product that can be made in a chemical reaction based on the amount of limiting reactant.
  – The ingredient that makes the least amount of pizza determines how many pizzas you can make (**theoretical yield**).

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Percent Yield

Assume that while making pizzas, we burn a pizza, drop one on the floor, or other uncontrollable events happen so that we make only two pizzas. The actual amount of product made in a chemical reaction is called the **actual yield**.

We can determine the efficiency of making pizzas by calculating the percentage of the maximum number of pizzas we actually make. In chemical reactions, we call this the **percent yield**.

$$\%\text{ yield} = \frac{\overset{\text{Actual yield}}{2\ \cancel{\text{pizzas}}}}{\underset{\text{Theoretical yield}}{3\ \cancel{\text{pizzas}}}} \times 100\% = 67\%$$

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In a Chemical Reaction

- For reactions with multiple reactants, it is likely that one of the reactants will be completely used before the others.
- When this reactant is used up, the reaction stops and no more product is made.
- The reactant that limits the amount of product is called the **limiting reactant**.
  - It is sometimes called the limiting reagent.
  - The limiting reactant gets completely consumed.
- Reactants not completely consumed are called **excess reactants**.
- The amount of product that can be made from the limiting reactant is called the **theoretical yield**.

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Summarizing Limiting Reactant and Yield (1 of 2)

- The **limiting reactant** (or **limiting reagent**) is the reactant that is completely consumed in a chemical reaction and limits the amount of product.
- The **reactant in excess** is any reactant that occurs in a quantity greater than is required to completely react with the limiting reactant.

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Summarizing Limiting Reactant and Yield (2 of 2)

- The **theoretical yield** is the amount of product that can be made in a chemical reaction based on the amount of limiting reactant.
- The **actual yield** is the amount of product actually produced by a chemical reaction.
- The **percent yield** is calculated as follows:

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

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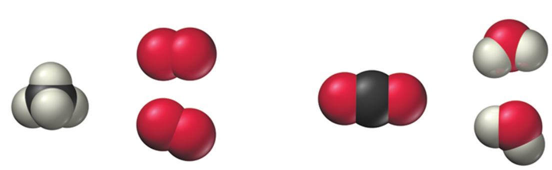
Calculating Limiting Reactant, Theoretical Yield, and Percent Yield

- Recall our balanced equation for the combustion of methane:

$$CH_4(g) + \mathbf{2}\,O_2(g) \rightarrow CO_2(g) + \mathbf{2}\,H_2O(g)$$

  – Our balanced equation for the combustion of methane implies that every one molecule of CH_4 reacts with two molecules of O_2.

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(l)$$



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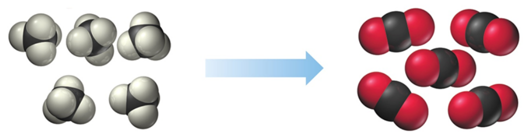
Combustion of Methane (3 of 4)

- If we have five molecules of CH_4 and eight molecules of O_2, which is the limiting reactant?

$$CH_4(g) + \mathbf{2}\, O_2(g) \rightarrow CO_2(g) + \mathbf{2}\, H_2O(g)$$

  - First we calculate the number of CO_2 molecules that can be made from five CH_4 molecules.

$$5\ \cancel{CH_4} \times \frac{1\ CO_2}{1\ \cancel{CH_4}} = 5\ CO_2$$



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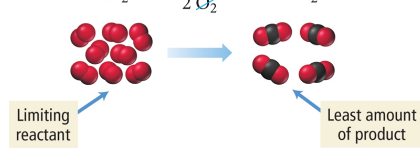
Combustion of Methane (4 of 4)

- Then we calculate the number of CO_2 molecules that can be made from eight O_2 molecules.

$$8\ \cancel{O_2} \times \frac{1\ CO_2}{2\ \cancel{O_2}} = 4\ CO_2$$



  - We have enough CH_4 to make five CO_2 molecules and four CO_2 molecules.
  - Therefore, O_2 is the limiting reactant, and four CO_2 molecules is the theoretical yield.
  - CH_4 is in excess.

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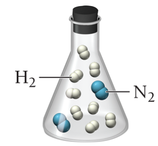
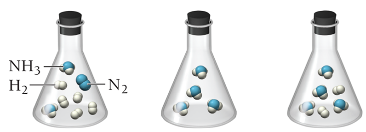
Conceptual Connection 4.6 (1 of 2)

Nitrogen and hydrogen gas react to form ammonia according to the reaction:

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



Which of the following images best represents the mixture in the flask above after the reactants have reacted as completely as possible? What is the limiting reactant? Which reactant is in excess?



(a) (b) (c)

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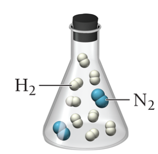
Conceptual Connection 4.6 (2 of 2)

Nitrogen and hydrogen gas react to form ammonia according to the reaction:

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



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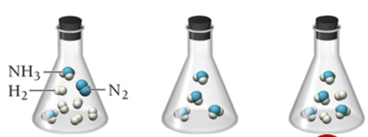

(a) (b) (c)

Limiting reactant: N_2
Excess reactant: H_2

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Calculating Limiting Reactant, Theoretical Yield, and Percent Yield from Reactant Masses (1 of 3)

- When working in the lab, we normally measure reactant quantities in grams.
- To find the limiting reactant and theoretical yield, we must first convert grams to moles.

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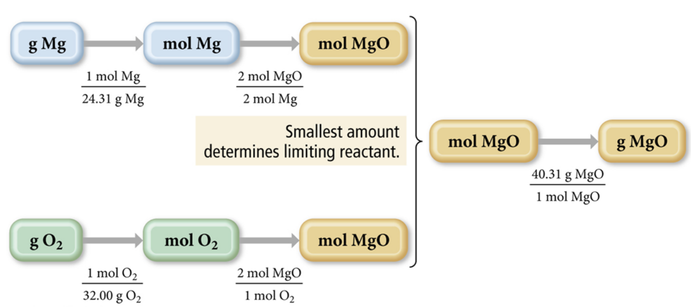
Calculating Limiting Reactant, Theoretical Yield, and Percent Yield from Reactant Masses (2 of 3)

- A reactant mixture contains 42.5 g Mg and 33.8 g O_2. What is the limiting reactant and theoretical yield?

$$2\ Mg(s) + O_2(g) \rightarrow \mathbf{2}\ MgO(s)$$



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Calculating Limiting Reactant, Theoretical Yield, and Percent Yield from Reactant Masses (3 of 3)

$$\boxed{42.5\ \cancel{\text{g Mg}}} \times \frac{1\ \cancel{\text{mol Mg}}}{24.31\ \cancel{\text{g Mg}}} \times \frac{2\ \text{mol MgO}}{2\ \cancel{\text{mol Mg}}} = \boxed{1.7\underline{4}83\ \text{mol MgO}}$$

Limiting reactant

Least amount of product

$$33.8\ \cancel{\text{g O}_2} \times \frac{1\ \cancel{\text{mol O}_2}}{32.00\ \cancel{\text{g O}_2}} \times \frac{2\ \text{mol MgO}}{1\ \cancel{\text{mol O}_2}} = 2.1\underline{1}25\ \text{mol MgO}$$

$$1.7\underline{4}83\ \cancel{\text{mol MgO}} \times \frac{40.31\ \text{g MgO}}{1\ \cancel{\text{mol MgO}}} = 70.5\ \text{g MgO}$$

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Conceptual Connection 4.7 (1 of 2)

Nitrogen dioxide reacts with water to form nitric acid and nitrogen monoxide according to the equation:

$$3\ NO_2(g) + H_2O(l) \rightarrow 2\ HNO_3(l) + NO(g)$$

Suppose that 5 mol NO_2 and 1 mol H_2O combine and react completely. How many moles of the reactant in excess are present after the reaction has completed?

a. 1 mol NO_2

b. 1 mol H_2O

c. 2 mol NO_2

d. 2 mol H_2O

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Conceptual Connection 4.7 (2 of 2)

Nitrogen dioxide reacts with water to form nitric acid and nitrogen monoxide according to the equation:

$$3\ NO_2(g) + H_2O(l) \rightarrow 2\ HNO_3(l) + NO(g)$$

Suppose that 5 mol NO_2 and 1 mol H_2O combine and react completely. How many moles of the reactant in excess are present after the reaction has completed?

a. 1 mol NO_2

b. 1 mol H_2O

c. 2 mol NO_2

d. 2 mol H_2O

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Combustion Reactions

- **Combustion reactions** are characterized by the reaction of a substance with O_2 to form one or more oxygen-containing compounds, often including water.
  - Combustion reactions also emit heat.
- For example, as you saw earlier in this chapter, natural gas (CH_4) reacts with oxygen to form carbon dioxide and water:

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$$

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Combustion

- Ethanol, the alcohol in alcoholic beverages, also reacts with oxygen in a combustion reaction to form carbon dioxide and water.

$$C_2H_5OH(l) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(g)$$

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Alkali Metal Reactions

- Reactions of alkali metals with nonmetals are vigorous.
  - Sodium and chlorine form sodium chloride.

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

- Alkali metals also react with water to form the dissolved alkali metal ion, the hydroxide ion, and hydrogen gas.

$$2\ M(s) + 2\ H_2O(l) \rightarrow 2\ M^+(aq) + 2\ OH^-(aq) + H_2(g)$$

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Reactions of the Alkali Metals with Water

Lithium Sodium Potassium

The reactions become progressively more vigorous as we move down the group.

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Halogen Reactions

- The halogens all react with many metals to form metal halides.

$$2\ Fe(s) + 3\ Cl_2(g) \rightarrow 2\ FeCl_3(s)$$

- The halogens also react with hydrogen to form hydrogen halides.

$$H_2(g) + X_2 \rightarrow 2\ HX(g)$$

- The halogens also react with each other to form interhalogen compounds.

$$Br_2(l) + F_2(g) \rightarrow 2\ BrF(g)$$

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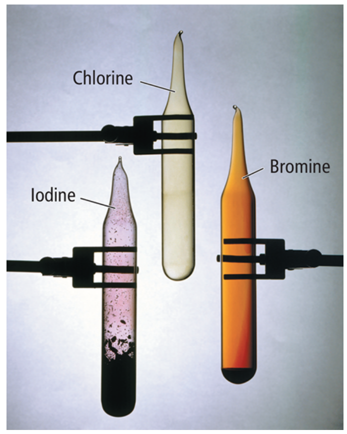
Three Halogens



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