Background for HW 3

IPOL 8512

Species-Area Curve

- A **species-area curve** is a relationship between the area of a habitat, or of part of a habitat, and the number of species found within that area.
- If S is the number of species, c is a constant that equals the number of species that would exist if the habitat area was confined to one area unit, A is the habitat area, and z is the slope of the species area relationship on a log S vs. log A graph,

$$S = cA^z$$

Species-Area Curve

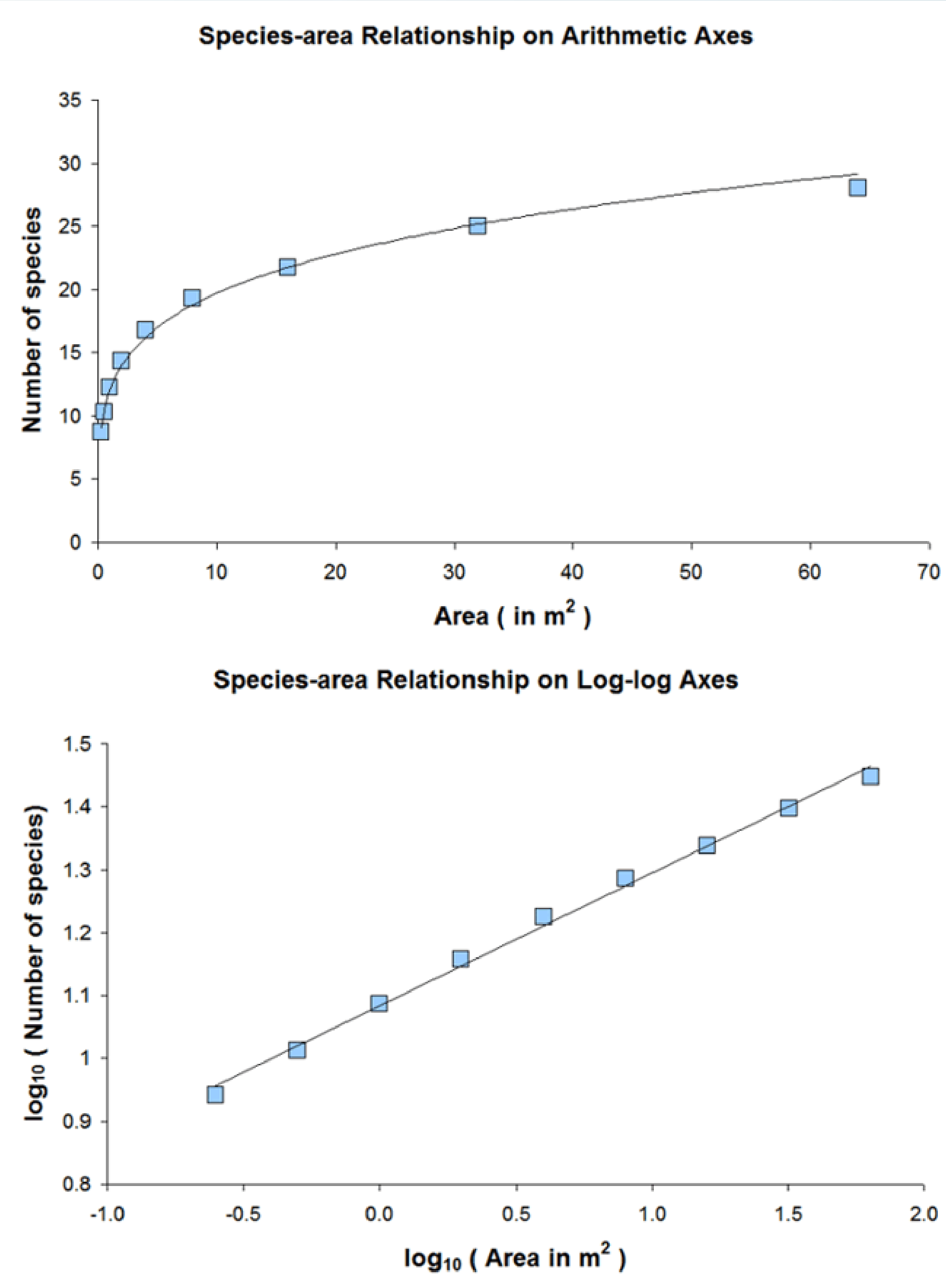
S = number of species

c = a constant that equals the number of species that would exist if the habitat area was confined to one area unit.

A = habitat area

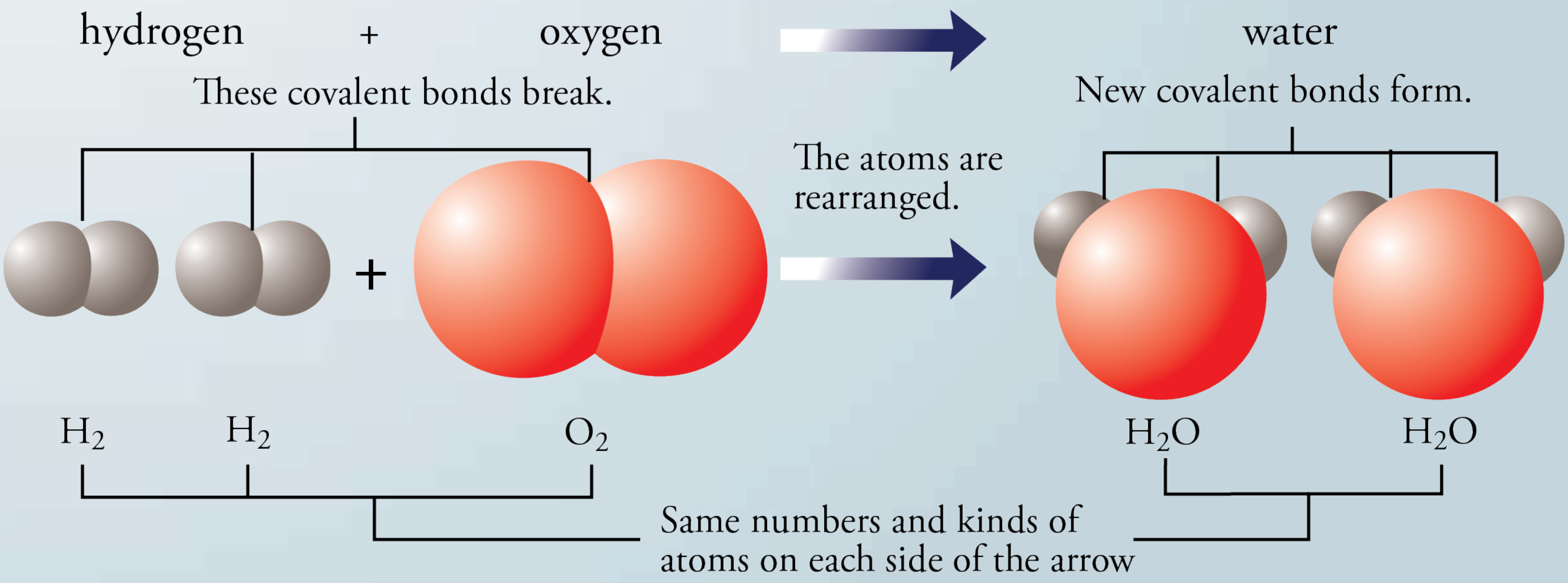
z = slope of the species area relationship on a log S vs. log A graph

$$S = cA^z$$



Chemical Reaction

- A ***chemical change*** or ***chemical reaction*** is a process in which one or more pure substances are converted into one or more different pure substances.



Chemical Equations (1)

- Chemical equations show the formulas for the substances that take part in the reaction.
  - The formulas on the left side of the arrow represent the ***reactants***, the substances that change in the reaction. The formulas on the right side of the arrow represent the ***products***, the substances that are formed in the reaction. If there are more than one reactant or more than one product, they are separated by plus signs. The arrow separating the reactants from the products can be read as *"goes to"* or *"yields" or "produces."*
- The physical states of the reactants and products are provided in the equation.
  - A *(g)* following a formula tells us the substance is a gas. Solids are described with *(s)*. Liquids are described with *(l)*. When a substance is dissolved in water, it is described with *(aq)* for "aqueous," which means "mixed with water."

Chemical Equations (2)

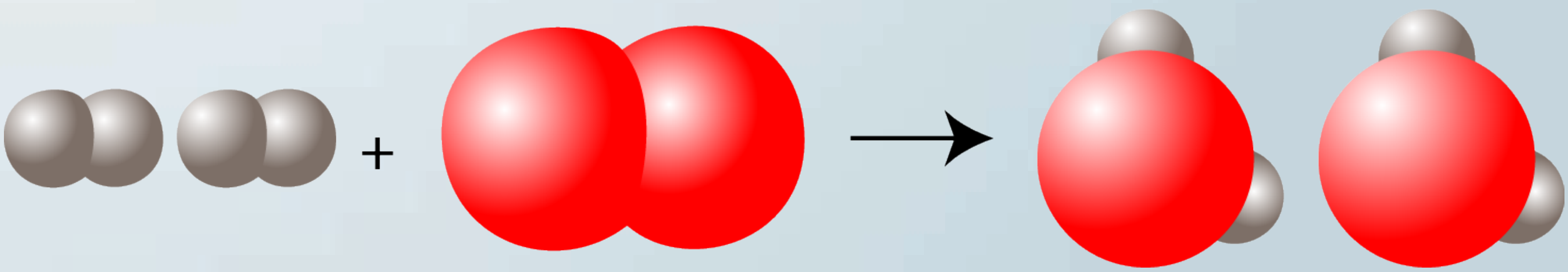
- The relative numbers of particles of each reactant and product are indicated by numbers placed in front of the formulas.
  - These numbers are called ***coefficients***. An equation containing correct coefficients is called a balanced equation.
  - If a formula in a balanced equation has no stated coefficient, its coefficient is understood to be 1.
- If special conditions are necessary for a reaction to take place, they are often specified above the arrow.
  - Some examples of special conditions are electric current, high temperature, high pressure, or light.

Chemical Equation Example

Coefficient   Coefficient

$$2H_2(g) + O_2(g) \xrightarrow{\text{"goes to"}} 2H_2O(l)$$

Physical states   Physical state



Special Conditions

A special condition

$$2H_2O(l) \xrightarrow{\text{Electric current}} 2H_2(g) + O_2(g)$$

This indicates the continuous addition of heat.

$$2KClO_3(s) \xrightarrow{\Delta} 2KCl(s) + 3O_2(g)$$

Balancing Chemical Equations

- Consider the first element listed in the first formula in the equation.
  - If this element is mentioned in two or more formulas on the same side of the arrow, skip it until after the other elements are balanced.
  - If this element is mentioned in one formula on each side of the arrow, balance it by placing coefficients in front of one or both of these formulas.
- Moving from left to right, repeat the process for each element.
- When you place a number in front of a formula that contains an element you tried to balance previously, recheck that element and put its atoms back in balance.

Balancing Equations – Strategies (1)

- **Strategy 1:** Often, an element can be balanced by using the subscript for this element on the left side of the arrow as the coefficient in front of the formula containing this element on the right side of the arrow and vice versa (using the subscript of this element on the right side of the arrow as the coefficient in front of the formula containing this element on the left side).
- **Strategy 2:** The pure nonmetallic elements (H_2, O_2, N_2, F_2, Cl_2, Br_2, I_2, S_8, Se_8, and P_4) can be temporarily balanced with a fractional coefficient (1/2, 3/2, 5/2, etc.). If you do use a fraction during the balancing process, you can eliminate it later by multiplying each coefficient in the equation by the fraction’s denominator.
- **Strategy 3:** If polyatomic ions do not change in the reaction, and therefore appear in the same form on both sides of the chemical equation, they can be balanced as if they were single atoms.
- **Strategy 4:** If you find an element difficult to balance, leave it for later.

Combustion Reactions

- A ***combustion reaction*** is a rapid oxidation that is accompanied by heat and usually light. They usually have oxygen, O_2, as one of the reactants.
  - When any substance that contains carbon is combusted (or burned) completely, the carbon forms carbon dioxide.
  - When a substance that contains hydrogen is burned completely, the hydrogen forms water.

$$2C_6H_{14}(g) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(l)$$

Combustion Reactions

- The complete combustion of a substance, like ethanol, C_2H_5OH, that contains carbon, hydrogen, and oxygen also yields carbon dioxide and water.

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

- When any substance that contains sulfur burns completely, the sulfur forms sulfur dioxide.

$$CH_3SH(g) + 3O_2(g) \rightarrow CO_2(g) + 2H_2O(l) + SO_2(g)$$

Byproducts of Combustion

- Incomplete combustion of hydrocarbons yields carbon monoxide and pure carbon (soot or ash).

- Combustion in air (which contains about 78% nitrogen) yields NO(g) and NO_2(g).

$$N_2(g) + O_2(g) \rightarrow 2NO(g)$$

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

Endergonic and Exergonic Changes

- **Endergonic**

more stable + energy $\rightarrow$ less stable system

lesser capacity to do work + energy $\rightarrow$ greater capacity to do work

lower PE + energy $\rightarrow$ higher PE

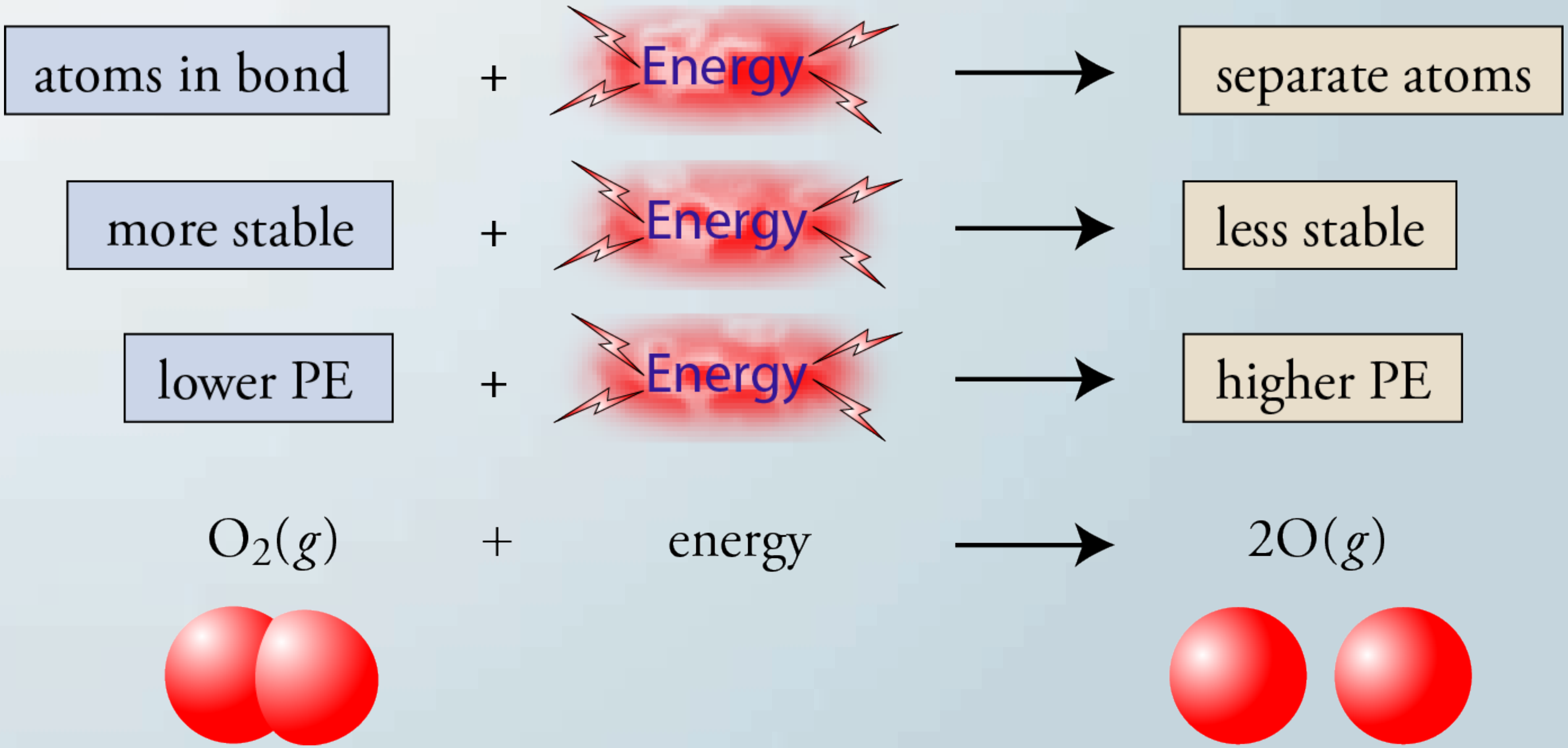
- **Exergonic**

less stable system $\rightarrow$ more stable + energy

greater capacity to do work $\rightarrow$ lesser capacity to do work + energy

higher PE $\rightarrow$ lower PE + energy

Bond Breaking and Potential Energy



Bond Making and Potential Energy

separate atoms ⟶ atoms in bond + Energy

less stable ⟶ more stable + Energy

higher PE ⟶ lower PE + Energy

$O(g) + O_2(g) \longrightarrow O_3(g) +$ energy

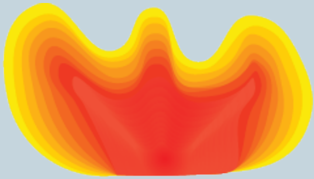
Exergonic (Exothermic) Reaction

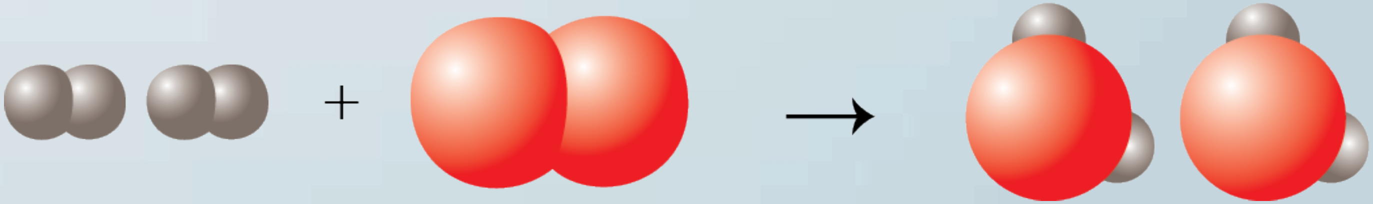
weaker bonds → stronger bonds + energy

less stable → more stable + energy

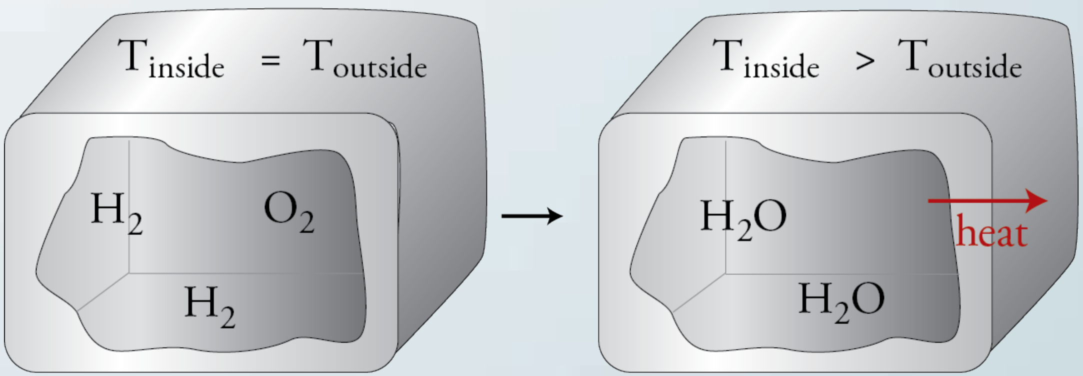
higher PE → lower PE + energy

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





Exothermic Reaction



Stronger bonds → More stable
↓
Energy released ← Lower PE
↓
Increases KE_{ave} of product particles
↓
Increased T → $T_{inside} > T_{outside}$
↓
Heat transferred to surroundings
↓
Exothermic

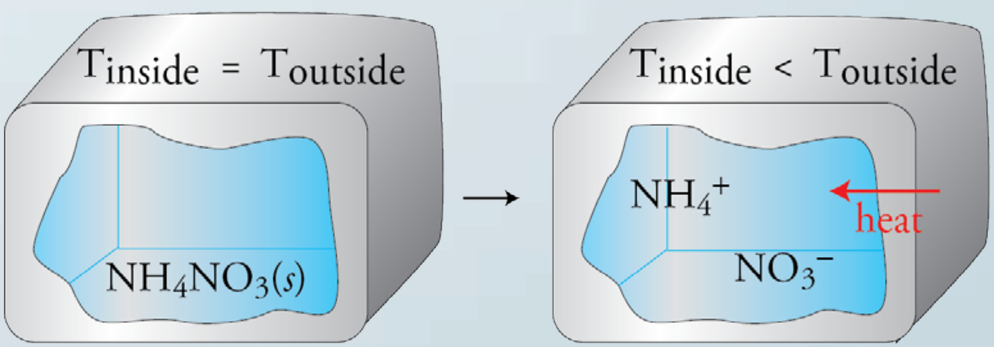
Endothermic Reaction

stronger bonds + energy $\rightarrow$ weaker bonds

more stable + energy $\rightarrow$ less stable

lower PE + energy $\rightarrow$ higher PE

$$\mathbf{NH_4NO_3}(\boldsymbol{s}) + \textbf{energy} \rightarrow \mathbf{NH_4^+}(\boldsymbol{aq}) + \mathbf{NO_3^-}(\boldsymbol{aq})$$



Weaker bonds $\rightarrow$ Less stable

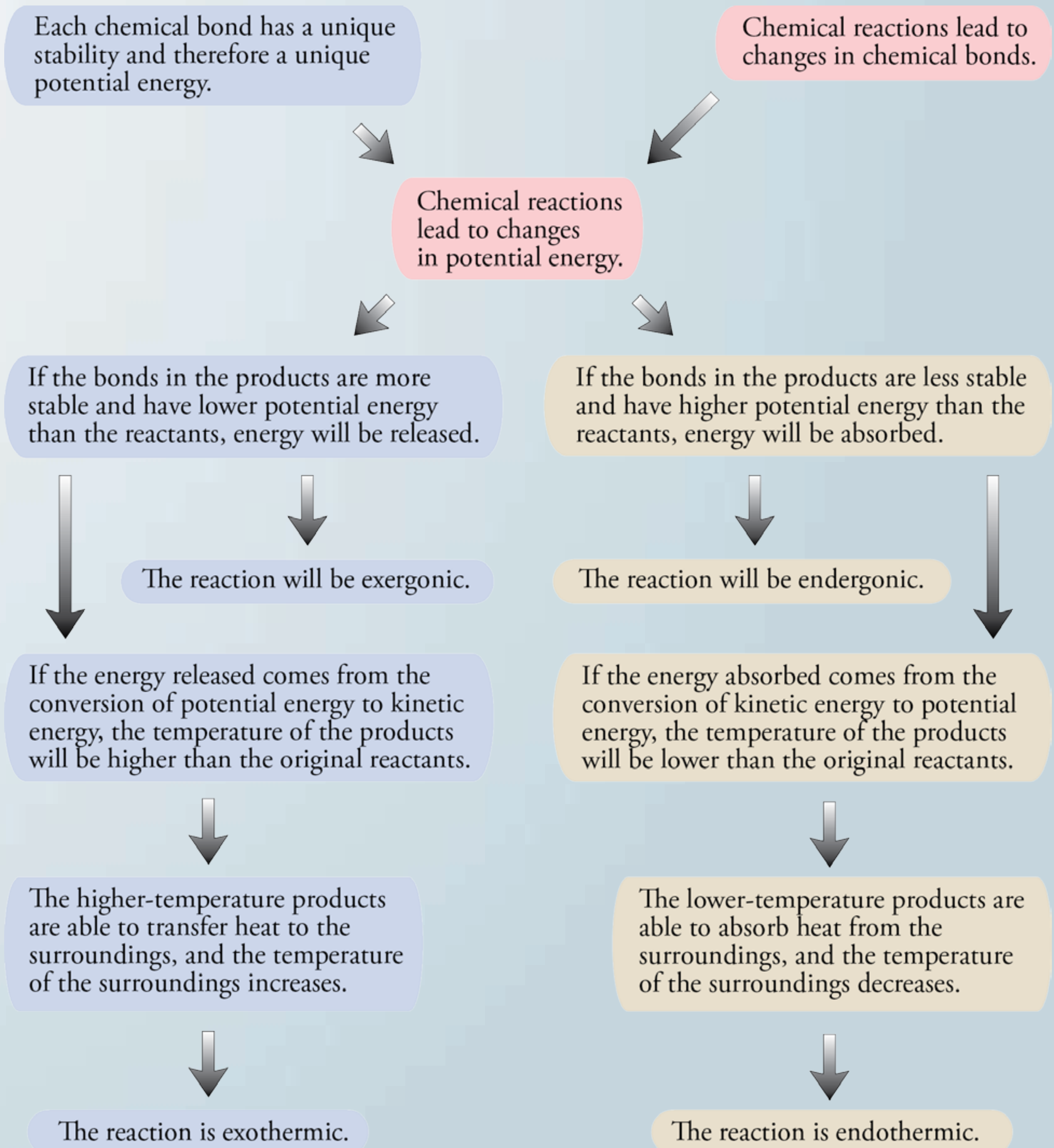
Energy absorbed $\leftarrow$ Higher PE

Decreases KE_{ave} of product particles

Decreased T $\rightarrow T_{inside} < T_{outside}$

Heat transferred to system $\rightarrow$ Endothermic

Energy and Chemical Reactions



Sample Calculations

- What is the minimum mass of water that must be added to 2.50×10^4 kg P_4O_{10} to form phosphoric acid in the following reaction?

$$P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$$

- The coefficients in the balanced equation provide us with a conversion factor that converts from units of P_4O_{10} to units of H_2O.

$$\left(\frac{1 \text{ mol } P_4O_{10}}{6 \text{ mol } H_2O}\right) \qquad \left(\frac{1 \text{ mol } P_4O_{10}}{4 \text{ mol } H_3PO_4}\right) \qquad \left(\frac{6 \text{ mol } H_2O}{4 \text{ mol } H_3PO_4}\right)$$

Example

What is the minimum mass of water that must be added to 2.50×10^4 kg P_4O_{10} to form phosphoric acid in the following reaction?

$$P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$$

$$? \text{ kg } H_2O = 2.50 \times 10^4 \cancel{\text{kg } P_4O_{10}} \left(\frac{10^3 \cancel{\text{g}}}{1 \cancel{\text{kg}}}\right)\left(\frac{1 \cancel{\text{mol } P_4O_{10}}}{283.889 \cancel{\text{g } P_4O_{10}}}\right)\left(\frac{6 \cancel{\text{mol } H_2O}}{1 \cancel{\text{mol } P_4O_{10}}}\right)\left(\frac{18.0153 \cancel{\text{g}} \; H_2O}{1 \cancel{\text{mol } H_2O}}\right)\left(\frac{1 \text{ kg}}{10^3 \cancel{\text{g}}}\right)$$

$$= 9.52 \times 10^3 \text{ kg } H_2O$$

$$? \text{ kg } H_2O = 2.50 \times 10^4 \cancel{\text{kg } P_4O_{10}} \left(\frac{6 \times 18.0153 \text{ kg } H_2O}{1 \times 283.889 \cancel{\text{kg } P_4O_{10}}}\right)$$

$$= 9.52 \times 10^3 \text{ kg } H_2O$$

General Steps

Measurable Property 1	Mass 1
↓	↓
Number of Particles 1	Moles 1
↓	↓
Number of Particles 2	Moles 2
↓	↓
Measurable Property 2	Mass 2

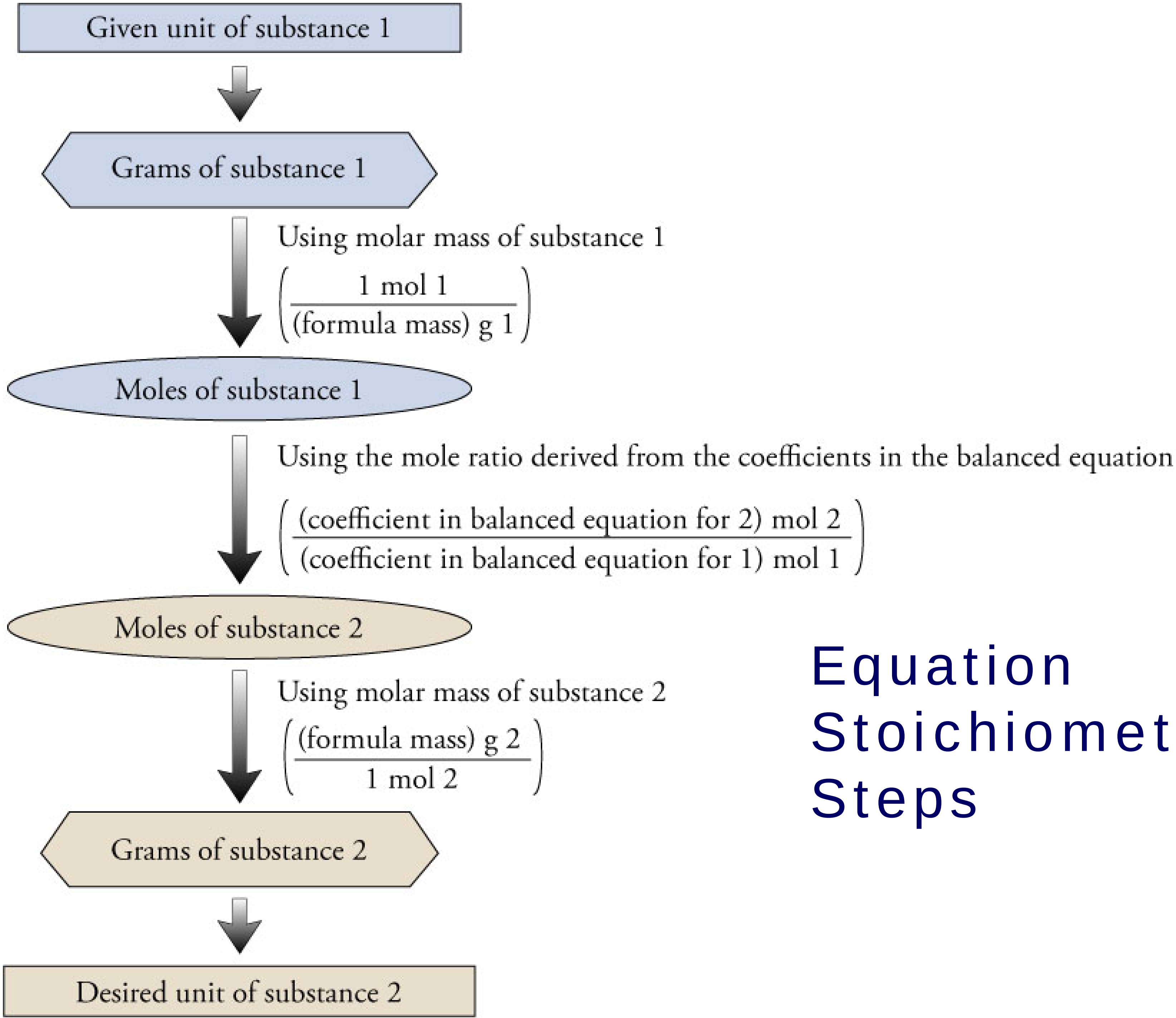
Equation Stoichiometry

- **Tip-off -** The calculation calls for you to convert from amount of one substance to amount of another, both of which are involved in a chemical reaction.

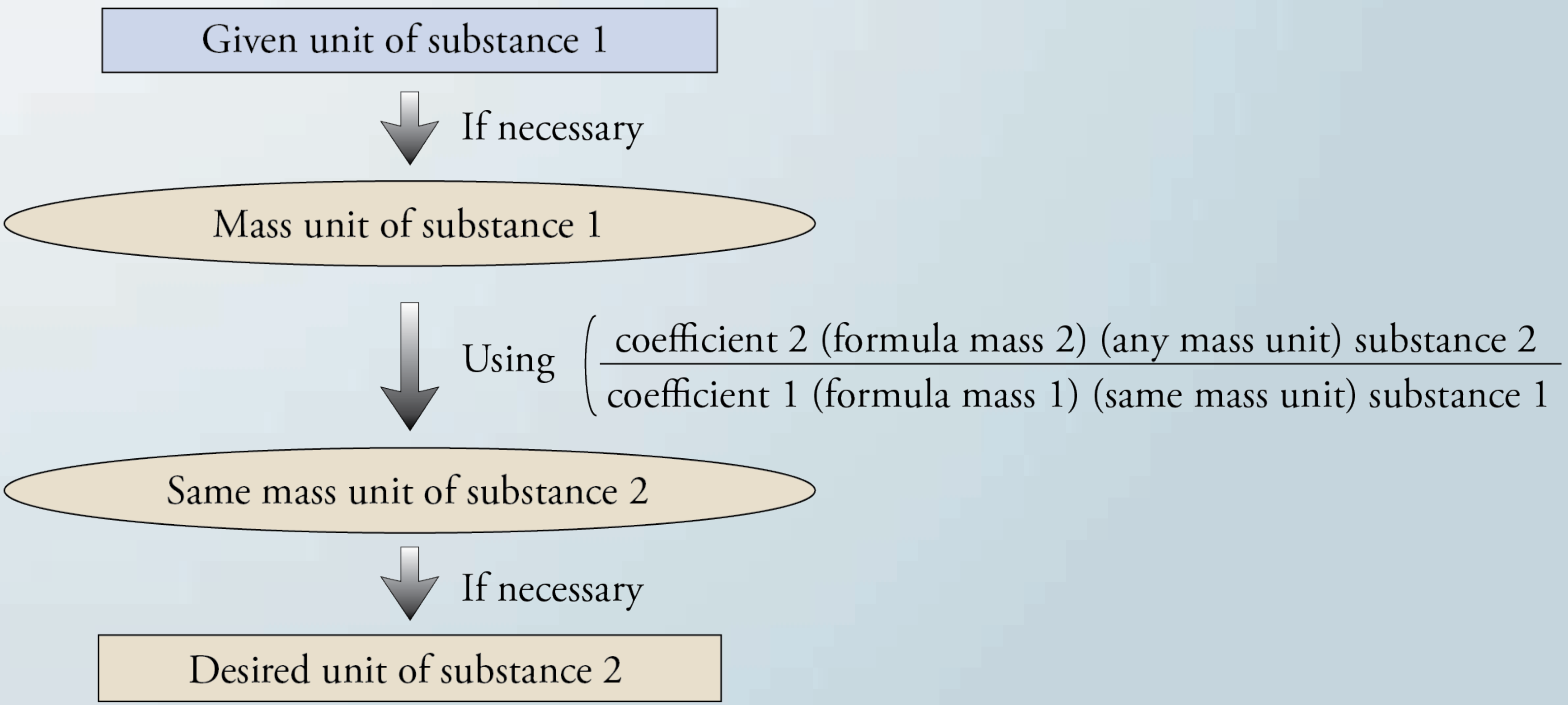
- **General Steps**

1. If you are not given it, write and balance the chemical equation for the reaction.
2. Start your unit analysis in the usual way.
3. If you are not given grams of substance 1, convert from the unit that you are given to grams. This may take one or more conversion factors.
4. Convert from grams of substance 1 to moles of substance 1.
5. Convert from moles of substance 1 to moles of substance 2 using the coefficients from the balanced equation to create the molar ratio used as a conversion factor.
6. Convert from moles of substance 2 to grams of substance 2, using its molar mass.
7. If necessary, convert from grams of 2 to the desired unit for 2. This may take one or more conversion factors.

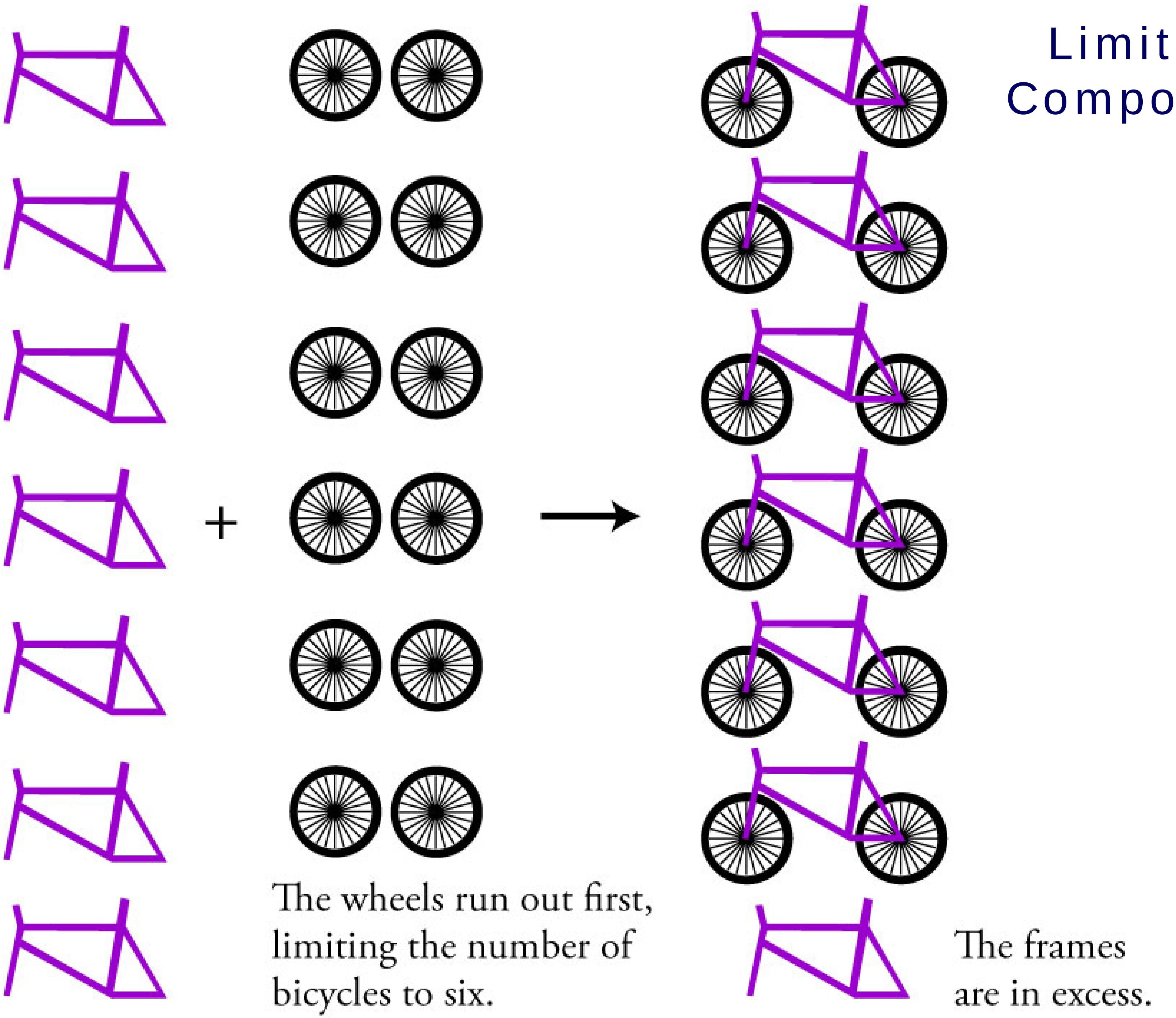
Equation Stoichiometry Steps



Equation Stoichiometry Shortcut



Limiting Component



Limiting Component (2)

$$\left(\frac{1\text{ bicycle}}{1\text{ frame}}\right) \quad \text{and} \quad \left(\frac{1\text{ bicycle}}{2\text{ wheels}}\right)$$

$$?\text{ bicycles} = 7\ \cancel{\text{frames}}\left(\frac{1\text{ bicycle}}{1\ \cancel{\text{frame}}}\right) = 7\text{ bicycles}$$

$$?\text{ bicycles} = 12\ \cancel{\text{wheels}}\left(\frac{1\text{ bicycle}}{2\ \cancel{\text{wheels}}}\right) = 6\text{ bicycles}$$

Limiting Reactant Problems

- The reactant that runs out first in a chemical reaction limits the amount of product that can form. This reactant is called the ***limiting reactant***.
- **Tip-off -** You are given two amounts of reactants in a chemical reaction, and you are asked to calculate the maximum amount of a product that can form from the combination of the reactants.
- **General Steps**

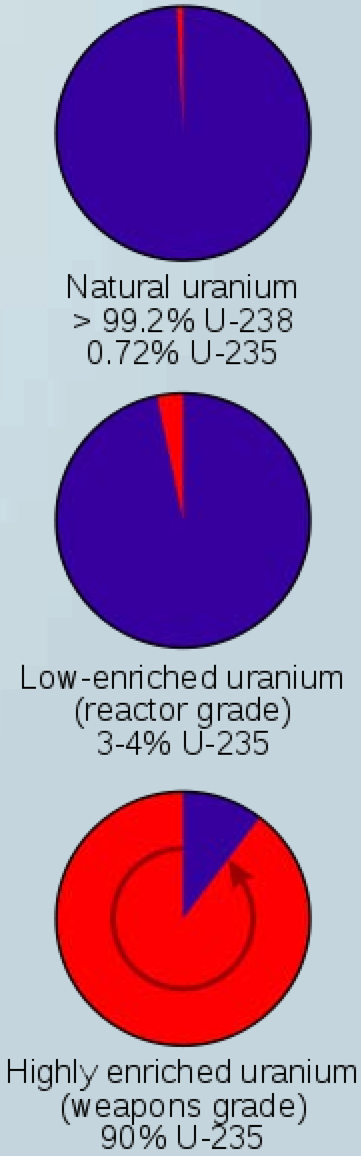
1. Do two separate calculations of the maximum amount of product that can form from each reactant.
2. The smaller of the two values calculated in the step above is your answer. It is the maximum amount of product that can be formed from the given amounts of reactants.

Example Problem

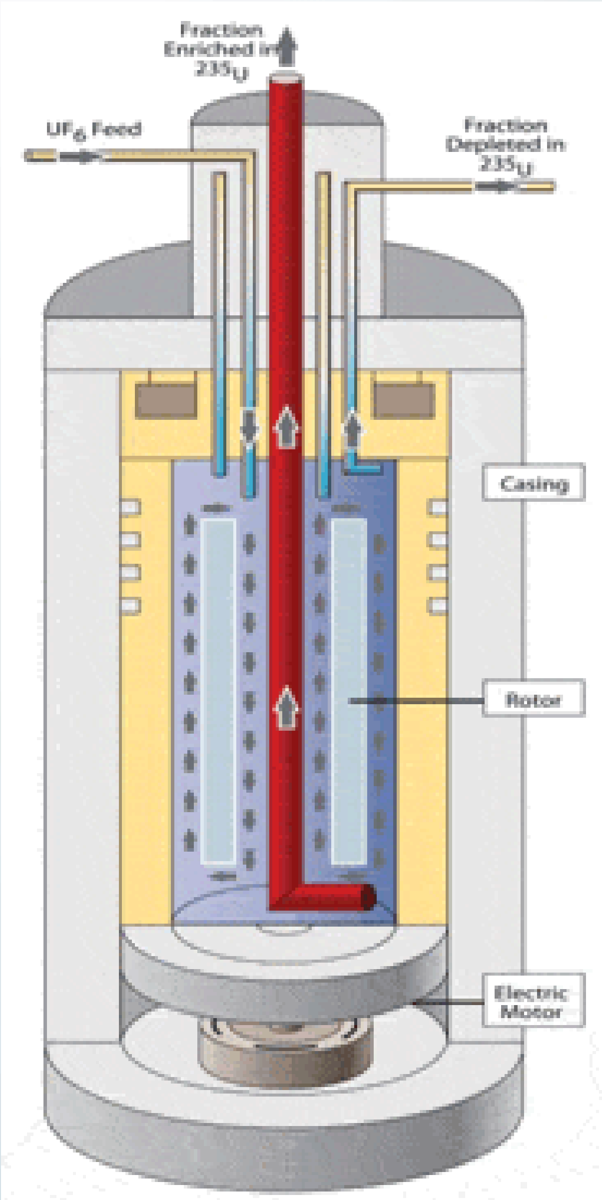
- The uranium(IV) oxide, UO_2, which is used as fuel in nuclear power plants, has a higher percentage of the fissionable isotope uranium-235 than is present in the UO_2 found in nature. To make fuel-grade UO_2, chemists first convert uranium oxides to uranium hexafluoride, UF_6, whose concentration of uranium-235 can be increased by a process called gas diffusion. The enriched UF_6 is then converted back to UO_2 in a series of reactions, beginning with

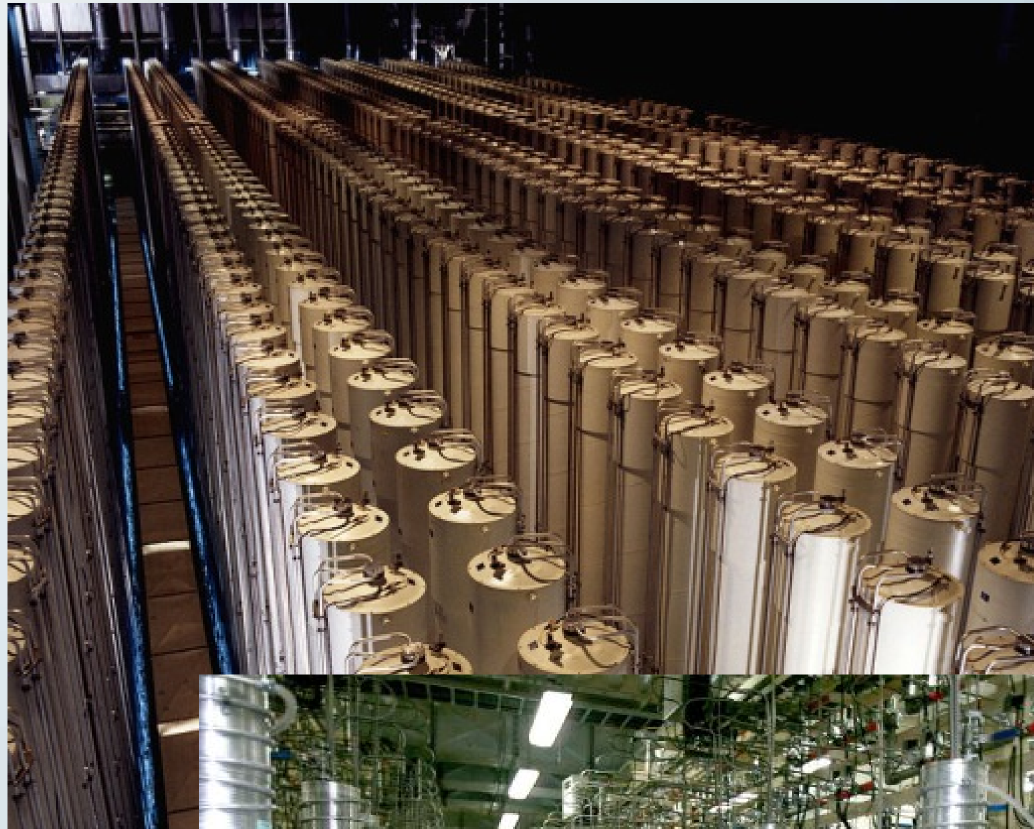
$$UF_6 + 2H_2O \rightarrow UO_2F_2 + 4HF$$

- How many megagrams of UO_2F_2 can be formed from the reaction of 24.543 Mg UF_6 with 8.0 Mg of water?



Gas Centrifuges





How many megagrams of UO_2F_2 can be formed from the reaction of 24.543 Mg UF_6 with 8.0 Mg of water?

$$UF_6 + 2H_2O \rightarrow UO_2F_2 + 4HF$$

$$? \text{ Mg } UO_2F_2 = 24.543 \text{ Mg } UF_6 \left(\frac{10^6 \text{ g}}{1 \text{ Mg}}\right)\left(\frac{1 \text{ mol } UF_6}{352.019 \text{ g } UF_6}\right)\left(\frac{1 \text{ mol } UO_2F_2}{1 \text{ mol } UF_6}\right)\left(\frac{308.0245 \text{ g } UO_2F_2}{1 \text{ mol } UO_2F_2}\right)\left(\frac{1 \text{ Mg}}{10^6 \text{ g}}\right)$$

or $$? \text{ Mg } UO_2F_2 = 24.543 \text{ Mg } UF_6 \left(\frac{1 \times 308.0245 \text{ Mg } UO_2F_2}{1 \times 352.019 \text{ Mg } UF_6}\right) = \mathbf{21.476 \text{ Mg } UO_2F_2}$$

$$? \text{ Mg } UO_2F_2 = 8.0 \text{ Mg } H_2O \left(\frac{10^6 \text{ g}}{1 \text{ Mg}}\right)\left(\frac{1 \text{ mol } H_2O}{18.0153 \text{ g } H_2O}\right)\left(\frac{1 \text{ mol } UO_2F_2}{2 \text{ mol } H_2O}\right)\left(\frac{308.0245 \text{ g } UO_2F_2}{1 \text{ mol } UO_2F_2}\right)\left(\frac{1 \text{ Mg}}{10^6 \text{ g}}\right)$$

or $$? \text{ Mg } UO_2F_2 = 8.0 \text{ Mg } H_2O \left(\frac{1 \times 308.0245 \text{ Mg } UO_2F_2}{2 \times 18.0153 \text{ Mg } H_2O}\right) = 68 \text{ Mg } UO_2F_2$$