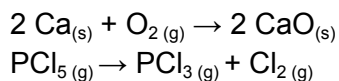


Influences on Reaction Rates

This worksheet will cover various common influences on reaction rates. Understanding the factors influencing reaction rates is important for predicting and controlling chemical processes. Concentration, temperature, the presence of a catalyst, and the nature of the reactants all affect the reaction rate in different ways. Some of these relationships can be represented mathematically, while others require a conceptual understanding.

1. Write the rate law for the equation $2 \text{H}_2\text{O}_{(g)} + \text{O}_{2(g)} \rightarrow 2 \text{H}_2\text{O}_{(l)}$
2. A reaction has a rate constant k of 0.05s^{-1} at a temperature of 300 K and an activation energy (E_a) of 40kJmol^{-1} . What is the new rate constant if the temperature is increased by 50 degrees? (Assume R is equal to $8.314 \text{J mol}^{-1}\text{K}^{-1}$)
3. Given a rate law $\text{Rate} = k[\text{A}]^{1/2}[\text{B}]^3$, determine the change in the reaction rate if the concentration of B is doubled while the concentration of A is cut in half.
4. The following reactions are conducted at the same time.



What is reactant nature in the context of these reactions, and how does it influence reaction rate? Predict which of the following reactions above will have the fastest initial reaction rate. (Hint: Use collision theory to justify your answer)

5. The reaction $3 \text{N}_{2(g)} + 3 \text{O}_{2(g)} \rightarrow 2 \text{NO}_{3(g)}$ is catalyzed by solid iron (Fe). How is the reaction rate affected with the presence of a catalyst? Draw a reaction coordinate diagram to justify your answer, and clearly label the transition state with and without a catalyst.
6. Challenge question: If the rate constant for a reaction is $0.02 \text{ M}^{-1}\text{s}^{-1}$ at 25°C and $0.10 \text{ M}^{-1}\text{s}^{-1}$ at 45°C calculate the activation energy for the reaction.

Answer Key

1. Write the rate law for the equation $2 \text{H}_2\text{O}_{(g)} + \text{O}_{2(g)} \rightarrow 2 \text{H}_2\text{O}_{(l)}$

Intro to Rate Laws:

The effect of concentration is represented by the “rate law”. For the equation $a\text{A} + b\text{B} \rightarrow c\text{C} + d\text{D}$, the rate law is written as “Rate = $k[\text{A}]^a[\text{B}]^b$ ”. This equation allows scientists to mathematically determine the influence of concentration on reaction rates. Rate constants, denoted by k , are introduced in rate laws to quantify the proportionality between the rate and the concentration. The rate constant, k , is constant for a given reaction at a given temperature.

Strategy - Using the balanced equation, coefficients, and skeleton rate law, “Rate = $k[\text{A}]^a[\text{B}]^b$ ”, write an expression.

$$\text{Rate} = k[\text{H}_2\text{O}]^2[\text{O}_2]^1$$

2. A reaction has a rate constant k of 0.05 s^{-1} at a temperature of 300 K and an activation energy (E_a) of 40 kJmol^{-1} . What is the new rate constant if the temperature is increased by 50 degrees? (Assume R is equal to $8.314 \text{ J mol}^{-1}\text{K}^{-1}$)

Intro to the Arrhenius Equation:

The influence of temperature is represented by the Arrhenius equation, $k = Ae^{\left(\frac{-E_a}{RT}\right)}$, where:

k is the rate constant

A is the pre-exponential (frequency) factor

E_a is the activation energy (the energy a reaction needs to break bonds in the reactants)

T is the temperature (in Kelvin)

R is the gas constant, which is equal to $8.314 \text{ Jmol}^{-1}\text{K}^{-1}$

Increasing the temperature of a reaction increases the rate constant, leading to a faster rate. Conceptually, this can be understood by using “collision theory”. Collision theory states that molecules must collide with the right orientation and enough speed in order to react. Increasing the temperature increases the speed at which the molecules collide, increasing the number of effective collisions.

Strategy - First, solve for A. Rearrange the Arrhenius Equation to get

$$A = k \div e^{\left(\frac{-E_a}{RT}\right)} = 0.05 \text{ s}^{-1} \div e^{\frac{-40,000\text{Jmol}^{-1}}{300\text{K} \cdot 8.314\text{J mol}^{-1}\text{K}^{-1}}} = 5.42 \times 10^{-9}$$

Now, use this value for A to find k at the new temperature.

$$k = (5.42 \times 10^{-9}) \times e^{\frac{-40,000\text{Jmol}^{-1}}{350\text{K} \cdot 8.314\text{J mol}^{-1}\text{K}^{-1}}} = 5.81 \times 10^{-15} \text{ s}^{-1}$$

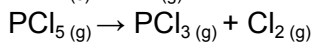
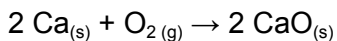
3. Given a rate law Rate = $k[\text{A}]^{1/2}[\text{B}]^3$, determine the change in the reaction rate if the concentration of B is doubled while the concentration of A is cut in half.

Strategy - The rate constant k will remain the same, so there is no need to include it when determining the change. Consider the factors in context with the rate law provided:

$$\text{Change} = [1/2]^{1/2} \times [2]^3 = 1/4 \times 8 = 2$$

Therefore, the rate will increase by a factor of 2.

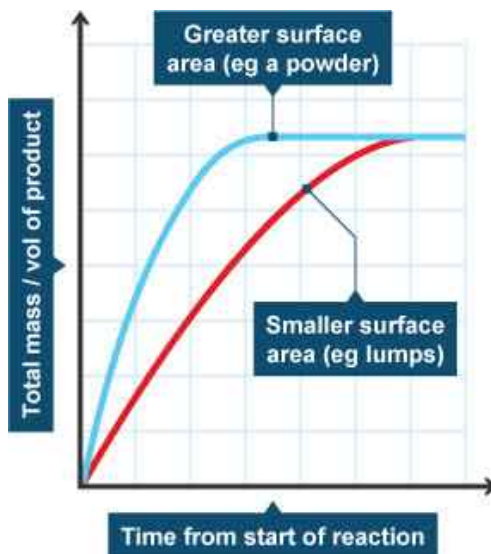
4. The following reactions are conducted at the same time.



What is reactant nature in the context of these reactions, and how does it influence reaction rate? Predict which of the following reactions above will have the fastest initial reaction rate. (Hint: Use collision theory to justify your answer)

Effect of Reactant Nature on Reaction Rate:

Reactant nature references the surface area of a reactant. A solid block of a substance would react slower than a powder, or smaller chunks, because the powder has more surface area. This relates to collision theory. Increasing the surface area increases the number of effective collisions, therefore increasing the reaction rate. This can be visually represented as such:



Credit: Pathwayz.org, "RATES OF REACTION - FACTORS"

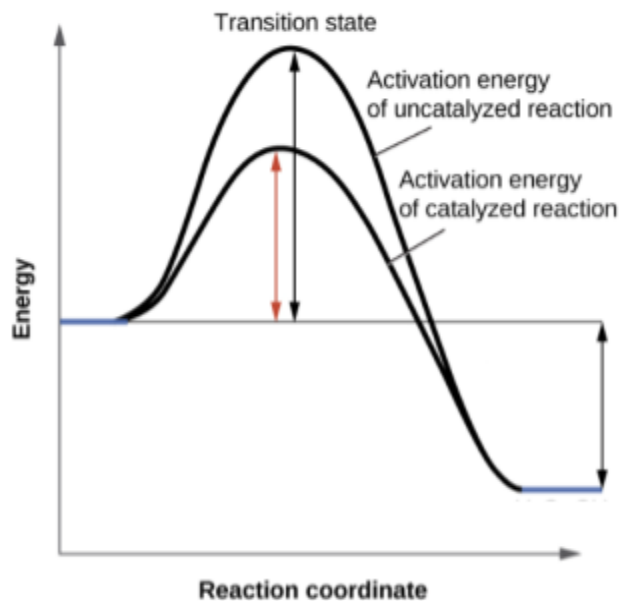
Strategy - Notice the phases of the reactants. The gaseous PCl_5 will react much quicker than the solid Ca . This, again, relates to collision theory. Gas molecules are already moving at higher speeds, increasing the chance for effective collisions. **Therefore, the 2nd reaction is predicted to have a faster initial rate.**

5. The reaction $3 \text{N}_{2(g)} + 3 \text{O}_{2(g)} \rightarrow 2 \text{NO}_{3(g)}$ is catalyzed by solid iron (Fe). How is the reaction rate affected with the presence of a catalyst? Draw a reaction coordinate

diagram to justify your answer, and clearly label the transition state with and without a catalyst.

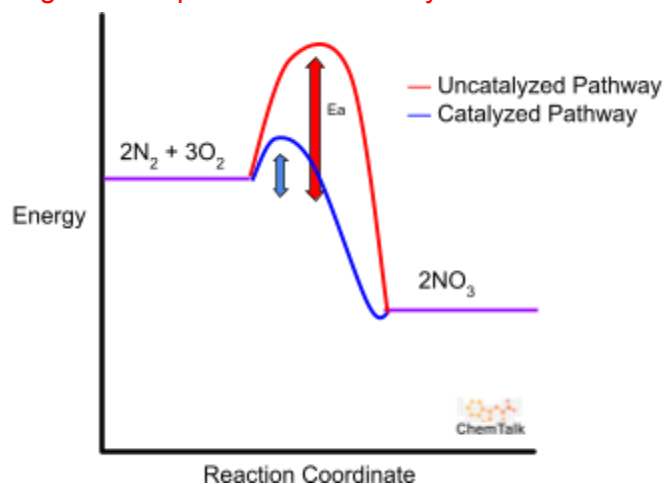
Effect of Catalysts on Reaction Rates:

The influence of catalysts is difficult to represent empirically, but can be conceptually understood. Catalysts function by providing an alternate reaction mechanism, lowering the overall activation energy of a reaction. This leads back to the arrhenius equation, where a lower E_a creates a faster reaction rate.



Credit: Pressbooks University of Hawaii, "Catalysis"

Strategy - The reaction rate will **increase** in the presence of a catalyst. When drawing your reaction coordinate, remember to label the products and reactants given in the problem. Note that the starting and ending points for the reaction coordinate remain the same for the catalyzed and uncatalyzed pathway. This is because the reactants and products are unchanged in the presence of a catalyst.



6. Challenge question: If the rate constant for a reaction is $0.02 \text{ M}^{-1}\text{s}^{-1}$ at 25°C and $0.10 \text{ M}^{-1}\text{s}^{-1}$ at 45°C calculate the activation energy for the reaction.

Strategy: This problem requires using algebra skills to rearrange the Arrhenius Equation. Set up two equations using the Arrhenius equation for the two temperature points:

$$k_1 = Ae^{\left(\frac{-Ea}{RT_1}\right)}$$

$$k_2 = Ae^{\left(\frac{-Ea}{RT_2}\right)}$$

Take the natural log of both equations

$$\ln(k_1) = -\frac{Ea}{R}\left(\frac{1}{T_1}\right) + \ln(A)$$

$$\ln(k_2) = -\frac{Ea}{R}\left(\frac{1}{T_2}\right) + \ln(A)$$

Subtract the two equations by each other to eliminate the pre-exponential factor (A)

$$\ln(k_2) - \ln(k_1) = -\frac{Ea}{R}\left(\frac{1}{T_2}\right) + \ln(A) - \left\{-\frac{Ea}{R}\left(\frac{1}{T_1}\right) + \ln(A)\right\}$$

Finally, combine the two equations

$$\ln\left(\frac{k_2}{k_1}\right) = -\frac{Ea}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

Plug in the numbers given in the problem, and solve for the activation energy. Remember to convert the temperatures given into Kelvin.

$$\ln\left(\frac{0.10 \text{ M}^{-1}\text{s}^{-1}}{0.02 \text{ M}^{-1}\text{s}^{-1}}\right) = -\frac{Ea}{8.314 \text{ J mol}^{-1}\text{K}^{-1}}\left(\frac{1}{298 \text{ K}} - \frac{1}{318 \text{ K}}\right)$$

Simplify, and isolate Ea

$$\ln(1/5) = -\frac{Ea}{8.314 \text{ J mol}^{-1}\text{K}^{-1}}\left(\frac{1}{298 \text{ K}} - \frac{1}{318 \text{ K}}\right)$$

$$Ea = -\frac{\ln(1/5) \times 8.314 \text{ J mol}^{-1}\text{K}^{-1}}{\frac{1}{298 \text{ K}} - \frac{1}{318 \text{ K}}} = 63401 \text{ J/mol} = 63.4 \text{ kJ/mol}$$