

The Advanced Placement Examination in Chemistry

Part II - Free Response Questions & Answers 1970 to 2005

Kinetics

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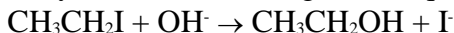
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1971

Ethyl iodide reacts with a solution of sodium hydroxide to give ethyl alcohol according to the equation.



The reaction is first order with respect to both ethyl iodide and hydroxide ion, and the overall-rate expression for the reaction is as follows:

$$\text{rate} = k[\text{CH}_3\text{CH}_2\text{I}][\text{OH}^-]$$

What would you do in the laboratory to obtain data to confirm the order in the rate expression for either of the reactants.

Answer:

The molar concentration of the hydroxide ion $[\text{OH}^-]$ can be determined by conducting the above reaction with a pH meter monitoring it. $[\text{OH}^-] = 1 \times 10^{-14} / -\log \text{pH}$. By measuring $[\text{OH}^-]$ over time and plotting $\ln[\text{OH}^-]$ vs time, if a straight line results, the reaction is 1st order with respect to $[\text{OH}^-]$.

1972



The following data about the reaction above were obtained from three experiments:

Experiment	[A]	[B]	Initial Rate of Formation of C (mole·liter ⁻¹ ·min ⁻¹)
1	0.60	0.15	6.3×10^{-3}
2	0.20	0.60	2.8×10^{-3}
3	0.20	0.15	7.0×10^{-4}

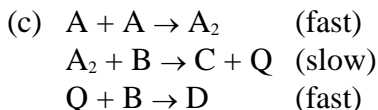
- What is the rate equation for the reaction?
- What is the numerical value of the rate constant k ? What are its dimensions?
- Propose a reaction mechanism for this reaction.

Answer:

$$(a) \text{ rate} = k [\text{A}]^2 [\text{B}]^1$$

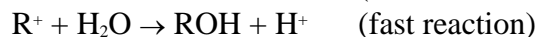
$$k = \frac{\text{rate}}{[\text{A}]^2 [\text{B}]} = \frac{6.3 \times 10^{-3} \text{ mol } \text{L}^{-1} \text{ min}^{-1}}{(0.60 \text{ mol } \text{L}^{-1})^2 (0.15 \text{ mol } \text{L}^{-1})}$$

$$(b) = 0.12 \text{ L}^2 \text{ mol}^{-2} \text{ min}^{-1}$$



1973 D

Some alkyl halides, such as $(\text{CH}_3)_3\text{CCl}$, $(\text{CH}_3)_3\text{CBr}$, and $(\text{CH}_3)_3\text{CI}$, represented by RX are believed to react with water according to the following sequence of reactions to produce alcohols:



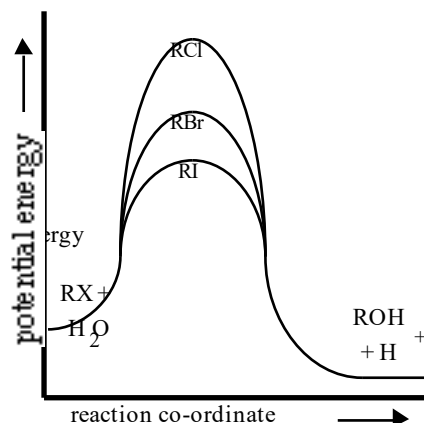
- For the hydrolysis of RX , write a rate expression consistent with the reaction sequence above.
- When the alkyl halides RCl , RBr , and RI are added to water under the same experimental conditions, the rates are in the order $\text{RI} > \text{RBr} > \text{RCl}$.

Construct properly labeled potential energy diagrams that are consistent with the information on the rates of hydrolysis of the three alkyl halides. Assume that the reactions are exothermic.

Answer:

$$(a) \text{ rate} = k [\text{RX}]$$

(b)



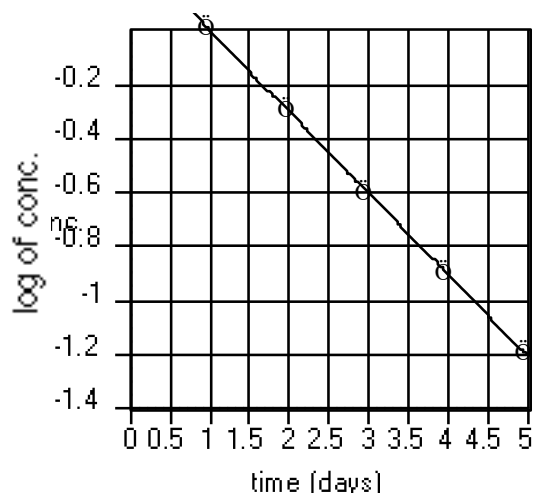
1974 D

A measure of the rate of a reaction is its half life. One method of determining the half life of a first order reaction is to plot certain appropriate data. Sketch a graph that illustrates the application of such a method. Label each axis with its name and appropriate units, and show how the half life can be obtained from the graph.

Answer:

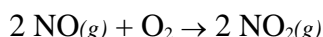
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The slope of the line is k , $t_{1/2} = 0.693/k$

1975 B



A rate expression for the reaction above is:

$$-d[\text{O}_2]/dt = k[\text{NO}]^2[\text{O}_2]$$

	ΔH_f° kcal/mole	S° cal/(mole)(K)	ΔG_f° kcal/mole
$\text{NO}(g)$	21.60	50.34	20.72
$\text{O}_2(g)$	0	49.00	0
$\text{NO}_2(g)$	8.09	57.47	12.39

- (a) For the reaction above, find the rate constant at 25°C if the initial rate, as defined by the equation above, is 28 moles per liter-second when the concentration of nitric oxide is 0.20 mole per liter and the concentration of oxygen is 0.10 mole per liter.
- (b) Calculate the equilibrium constant for the reaction at 25°C.

Answer:

$$k = \frac{\text{rate}}{[\text{NO}]^2[\text{O}_2]} = \frac{28 \text{ mol L}^{-1} \text{sec}^{-1}}{(0.20 \text{ mol L}^{-1})^2 (0.10 \text{ mol L}^{-1})}$$

(a)

$$= 7000 \text{ L}^2 \text{mol}^{-2} \text{sec}^{-1}$$

$$(b) \Delta G^\circ = 2\Delta G^\circ_{\text{NO}_2} - 2\Delta G^\circ_{\text{NO}} = (2)(12.39) - (2)(20.72)$$

$$= -16.66 \text{ kcal} = -16660 \text{ cal} = -69710 \text{ J}$$

$$K = e^{-\Delta G/RT} = e^{-[-69710/(8.314 \times 298)]} = 1.65 \times 10^{12}$$

1976 D

Changing the temperature and no other conditions changes the rates of most chemical reactions. Two factors are commonly cited as accounting for the increased rate of chemical reaction as the temperature is

increased. State briefly and discuss the two factors. Which of the two is more important?

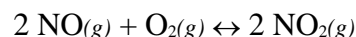
Answer:

Energy factor - enough energy in the collision for the formation of an activated complex, where bonds are breaking and new ones forming. When temperature is increased, a greater number of molecular collisions possess enough energy to activate the reaction (activation energy).

Frequency of collisions increases - an increase in temperature makes particles move faster and collide more frequently, increasing the possibility of a reaction between them.

Energy factor is more important.

1977 B



For the reaction above, the rate constant at 380°C for the forward reaction is $2.6 \times 10^3 \text{ liter}^2/\text{mole}^2\text{-sec}$ and this reaction is first order in O_2 and second order in NO . The rate constant for the reverse reaction at 380°C is $4.1 \text{ liter/mole-sec}$ and this reaction is second order in NO_2 .

- (a) Write the equilibrium expression for the reaction as indicated by the equation above and calculate the numerical value for the equilibrium constant at 380°C.
- (b) What is the rate of the production of NO_2 at 380°C if the concentration of NO is 0.0060 mole/liter and the concentration of O_2 is 0.29 mole/liter?
- (c) The system above is studied at another temperature. A 0.20 mole sample of NO_2 is placed in a 5.0 liter container and allowed to come to equilibrium. When equilibrium is reached, 15% of the original NO_2 has decomposed to NO and O_2 . Calculate the value for the equilibrium constant at the second temperature.

Answer:

$$K = \frac{[\text{NO}_2]^2}{[\text{NO}]^2[\text{O}_2]}$$

(a)

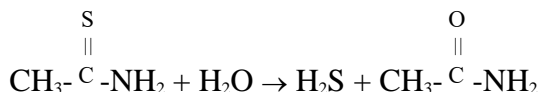
$$k_{\text{for.}}[\text{NO}]^2[\text{O}_2] = k_{\text{rev.}}[\text{NO}_2]^2; K = k_{\text{for.}} / k_{\text{rev.}}$$

$$K = 2.6 \times 10^3 / 4.1 = 6.3 \times 10^2 \text{ M}^{-1}$$

- (b) $R = k[\text{NO}]^2[\text{O}_2] = (2.6 \times 10^3)(0.0060)^2(0.29) = 0.027 \text{ M/sec}$ acceptable but $R = d[\text{NO}_2]/dt = 0.054 \text{ M/sec}$ better.

- (c) $[\text{NO}_2] = (0.20 - 15\%) / 5\text{L} = 0.034 \text{ M}$
 $[\text{O}_2] = (15\% \text{ of } 0.20 / 2) / 5\text{L} = 0.0030 \text{ M}$
 $[\text{NO}] = (15\% \text{ of } 0.20) / 5\text{L} = 0.0060 \text{ M}$
 $K = (0.034)^2 / [(0.0060)^2(0.0030)] = 1.1 \times 10^4 \text{ M}^{-1}$

1979 D



The hydrolysis of thioacetamide is used to generate H_2S as shown by the equation above. The rate of the reaction is given by the rate law as follows:

$$\text{rate} = k[\text{H}^+][\text{CH}_3-\overset{\text{S}}{\underset{\parallel}{\text{C}}}-\text{NH}_2]$$

Consider a solution which is 0.10 molar in H^+ and 0.10 molar in thioacetamide at 25° for each of the changes listed in (1), (2) and (3) below, state whether

- the rate of reaction increases, decreases or remains the same.
- the numerical value of k increases, decreases or remains the same.
 - Sodium acetate is added to the solution.
 - The solution is heated to 75°C .
 - Water is added to the solution.

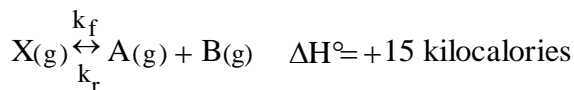
Give a brief explanation for each of your answers.

Answer:

- Rate decreases. $[\text{H}^+]$ decreases because H^+ reacts with CH_3COO^- , a good base.
 - k remains constant. k is a function of temperature only or is concentration independent.
- Rate increases. The number - or fraction - of effective collisions increases.
 - k increases. This change is predicted by the Arrhenius equation, $k = A_e^{-E_a/RT}$. [The explanation in (a) was also accepted.]
- Rate decreases. Adding water decreases the concentration and results in a decrease in the frequency of collisions.
 - k remains constant. Explanation is the same as in (1b).

1980 D

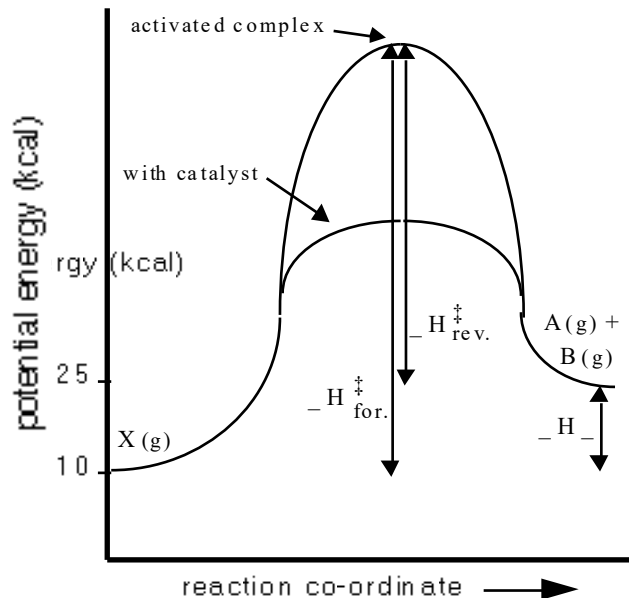
The decomposition of compound X is an elementary process that proceeds as follows:



The forward reaction is slow at room temperature but becomes rapid when a catalyst is added.

- Draw a diagram of potential energy vs reaction coordinate for the uncatalyzed reaction. On this diagram label:
 - the axes
 - the energies of the reactants and the products
 - the energy of the activated complex
 - all significant energy differences
- On the same diagram indicate the change or changes that result from the addition of the catalyst. Explain the role of the catalyst in changing the rate of the reaction.
- If the temperature is increased, will the ratio k_f/k_r increase, remain the same, or decrease? Justify your answer with a one or two sentence explanation. [k_f and k_r are the specific rate constants for the forward and the reverse reactions, respectively.]

Answer:



- [See above diagram]. The catalyst changes the mechanism and/or increases the number of molecules with sufficient energy to react.
- A temperature increases the ratio k_f/k_r . Any one of the following applies:
 - $k_f/k_r = K$ and LeChatelier's principle applies.
 - Boltzmann distribution graph.

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- (3) ΔS changes value for ΔG for the equation
- (4) Forward reaction is endothermic so k_f will be increased more than k_r by temperature increase.
- (5) Plot of $\ln k$ vs. $1/T$ has slope proportional to E_a ; because $E_{a(\text{for.})}$ is greater than $E_{a(\text{rev.})}$, k_f will increase more than k_r .

1983 C

Graphical methods are frequently used to analyze data and obtain desired quantities.



The following data give the value of the rate constant at various temperatures for the gas phase reaction above.

T (K)	k (litre/mol sec)
647	8.58×10^{-5}
666	2.19×10^{-4}
683	5.11×10^{-4}
700	1.17×10^{-3}
716	2.50×10^{-3}

Describe, without doing any calculations, how a graphical method can be used to obtain the activation energy for this reaction.



The following data give the partial pressure of A as a function of time and were obtained at 100°C for the reaction above.

P_A (mm Hg)	t (sec)
348	0
247	600
185	1200
105	2400
58	3600

Describe, without doing any calculations, how graphs can be used to determine whether this reaction is first or second order in A and how these graphs are used to determine the rate constant.

Answer:

- (a) Plot $\ln k$ vs. $1/T$; $E_{\text{act}} = -R$ (slope)

OR

Plot $\log k$ vs. $1/T$; $E_{\text{act}} = -2.303 R$ (slope)

- (b) Plot $\ln P_A$ or $\log P_A$ vs. time

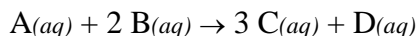
Plot $1/P_A$ vs. time

If the former is linear, the reaction is 1st order. If the latter is linear, the reaction is 2nd order. If the reaction is 1st order, slope = $-k_1$ or $-k_1/2.303$. If 2nd order, slope = k_2 .

1984 B

For a hypothetical chemical reaction that has the stoichiometry $2 \text{X} + \text{Y} \rightarrow \text{Z}$, the following initial rate data

1981 B



For the reaction above, carried out in solution of 30°C , the following kinetic data were obtained:

Experiment	Initial Conc. of Reactants (mole liter ⁻¹)		Initial Rate of Reaction (mole liter ⁻¹ hr ⁻¹)
	A _o	B _o	
1	0.240	0.480	8.00
2	0.240	0.120	2.00
3	0.360	0.240	9.00
4	0.120	0.120	0.500
5	0.240	0.0600	1.00
6	0.0140	1.35	?

- (a) Write the rate-law expression for this reaction.
- (b) Calculate the value of the specific rate constant k at 30°C and specify its units.
- (c) Calculate the value of the initial rate of this reaction at 30°C for the initial concentrations shown in experiment 6.
- (d) Assume that the reaction goes to completion. Under the conditions specified for experiment 2, what would be the final molar concentration of C?

Answer:

- (a) Rate = $k [\text{A}]^2 [\text{B}]$

$$k = \frac{\text{rate}}{[\text{A}]^2 [\text{B}]} = \frac{8.00 \text{ mol L}^{-1} \text{ hr}^{-1}}{(0.240 \text{ mol L}^{-1})^2 (0.480 \text{ mol L}^{-1})}$$

(b) $= 289 \text{ L}^2 \text{ mol}^{-2} \text{ hr}^{-1}$

- (c) Rate = $k [\text{A}]^2 [\text{B}]$
 $= (289 \text{ L}^2 \text{ mol}^{-2} \text{ hr}^{-1})(0.0140 \text{ mol L}^{-1})^2 (1.35 \text{ mol L}^{-1})$
 $= 0.0766 \text{ mol L}^{-1} \text{ hr}^{-1}$

- (d) According to the equation: 2 mol B reacts with 1 mol A, therefore, B is the limiting reagent, while only 0.006 mole/L of A reacts.

$$0.120 \text{ mol/L B} \propto \frac{3 \text{ mol/L C}}{2 \text{ mol/L B}} = 0.180 \text{ mol/L C}$$

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were obtained. All measurements were made at the same temperature.

Initial Rate of Formation of Z, (mol L ⁻¹ sec ⁻¹)	Initial [X] ₀ , (mol L ⁻¹)	Initial [Y] ₀ , (mol L ⁻¹)
7.0 × 10 ⁻⁴	0.20	0.10
1.4 × 10 ⁻³	0.40	0.20
2.8 × 10 ⁻³	0.40	0.40
4.2 × 10 ⁻³	0.60	0.60

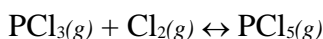
- Give the rate law for this reaction from the data above.
- Calculate the specific rate constant for this reaction and specify its units.
- How long must the reaction proceed to produce a concentration of Z equal to 0.20 molar, if the initial reaction concentrations are [X]₀ = 0.80 molar, [Y]₀ = 0.60 molar and [Z]₀ = 0 molar?
- Select from the mechanisms below the one most consistent with the observed data, and explain your choice. In these mechanisms M and N are reaction intermediates.

- $X + Y \rightarrow M$ (slow)
 $X + M \rightarrow Z$ (fast)
- $X + X \leftrightarrow M$ (fast)
 $Y + M \rightarrow Z$ (slow)
- $Y \rightarrow M$ (slow)
 $M + X \rightarrow N$ (fast)
 $N + X \rightarrow Z$ (fast)

Answer:

- Rate = $k [X]^2 [Y]$
- $k = \text{rate}/[Y] = (0.00070 \text{ mol L}^{-1} \text{ sec}^{-1}) / (0.10 \text{ mol L}^{-1}) = 0.0070 \text{ sec}^{-1}$
- $\ln c_0/c = kt; \ln 0.60/0.040 = (0.0070)t$
 $t = 58 \text{ sec.}$
- Mechanism 3 is correct. The rate law shows that the slow reaction must involve one Y, consistent with mechanism 3.
 Mechanisms 1 and 2 would involve both [X] and [Y] in the rate law, not consistent with the rate law.

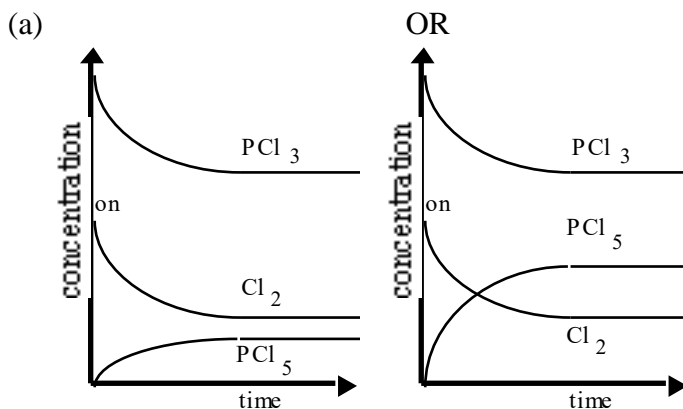
1985 D



In the equation above, the forward reaction is first order in both PCl₃ and Cl₂ and the reverse reaction is first order in PCl₅.

- Suppose that 2 moles of PCl₃ and 1 mole of Cl₂ are mixed in a closed container at constant temperature. Draw a graph that shows how the concentrations of PCl₃, Cl₂, and PCl₅ change with time until after equilibrium has been firmly established.
- Give the initial rate law for the forward reaction.
- Provide a molecular explanation for the dependence of the rate of the forward reaction on the concentrations of the reactants.
- Provide a molecular explanation for the dependence of the rate of the forward reaction on temperature.

Answer:



- Rate = $k [\text{PCl}_3][\text{Cl}_2]$
- Reaction requires effective collisions between molecules of PCl₃ and Cl₂. As concentrations of these molecules increase, the number of effective collisions must increase and the rate of reaction increases.
- The fraction of colliding molecules with the required activation energy increases as the temperature rises.

1986 D

The overall order of a reaction may not be predictable from the stoichiometry of the reaction.

- Explain how this statement can be true.
- $2 \text{XY} \rightarrow \text{X}_2 + \text{Y}_2$
 - For the hypothetical reaction above, give a rate law that shows that the reaction is first order in the reactant XY.

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- Give the units for the specific rate constant for this rate law.
- Propose a mechanism that is consistent with both the rate law and the stoichiometry.
- What is the numerical value for the initial rate of disappearance of $C_2O_4^{2-}$ for Experiment 1?
- Calculate the initial oxalate ion concentration for Experiment 4.

Answer:

- (a) Order of reaction determined by the slowest step in the mechanism. **OR**
Order of reaction determined by exponents in the rate law. **OR**

Providing a counterexample where the coefficients in equation and exponents in rate law are different.

- (b) 1. Rate = $k [XY]$ or equivalent
2. $k = 1/\text{time}$ or units consistent with student's rate equation
3. Mechanism proposed should show:
a. steps adding up to the overall reaction
b. one step starting with XY
c. rate-determining step involving XY
example:
 $XY \rightarrow X + Y$ (slow)
 $XY + X \rightarrow X_2 + Y$ (fast)
 $Y + Y \rightarrow Y_2$ (fast)

1987 B



The equation for the reaction between mercuric chloride and oxalate ion in hot aqueous solution is shown above. The reaction rate may be determined by measuring the initial rate of formation of chloride ion, at constant temperature, for various initial concentrations of mercuric chloride and oxalate as shown in the following table

Experiment	Initial $[\text{HgCl}_2]$	Initial $[\text{C}_2\text{O}_4^{2-}]$	Initial Rate of Formation of Cl^- ($\text{mol L}^{-1} \text{min}^{-1}$)
(1)	0.0836 M	0.202M	0.52×10^{-4}
(2)	0.0836 M	0.404M	2.08×10^{-4}
(3)	0.0418 M	0.404M	1.06×10^{-4}
(4)	0.0316 M	?	1.27×10^{-4}

- (a) According to the data shown, what is the rate law for the reaction above?
- (b) On the basis of the rate law determined in part (a), calculate the specific rate constant. Specify the units.

Answer:

(a) Rate = $k [\text{HgCl}_2][\text{C}_2\text{O}_4^{2-}]^2$

$$k = \frac{\text{Rate}}{[\text{HgCl}_2][\text{C}_2\text{O}_4^{2-}]^2} = \frac{5.2 \times 10^{-5} \text{ M min}^{-1}}{(0.0836)(0.202)^2 \text{ M}^3}$$

(b) $= 1.52 \times 10^{-2} \text{ M}^{-2} \text{ min}^{-1}$

$$\frac{d[\text{C}_2\text{O}_4^{2-}]}{dt} = \frac{5.2 \times 10^{-5} \text{ M}}{\text{min}} \propto \frac{1 \text{ mol C}_2\text{O}_4^{2-}}{2 \text{ mol Cl}^-}$$

(c) $= 2.6 \times 10^{-5} \text{ M/min}$

$$[\text{C}_2\text{O}_4^{2-}] = \sqrt{\frac{\text{Rate}}{k[\text{HgCl}_2]}}$$

(d) $= \sqrt{\frac{1.27 \times 10^{-4} \text{ M min}^{-1}}{(1.52 \times 10^{-2} \text{ M}^{-2} \text{ min}^{-1})(0.0316 \text{ M})}}$

$$= 0.514 \text{ M}$$

1989 D



Account for the following observations regarding the exothermic reaction represented by the equation above.

- An increase in the pressure of the reactants causes an increase in the reaction rate.
- A small increase in temperature causes a large increase in the reaction rate.
- The presence of metallic nickel causes an increase in reaction rate.
- The presence of powdered nickel causes a larger increase in reaction rate than does the presence of a single piece of nickel of the same mass.

Answer:

- Effective concentrations are increased. So collision frequency is increased.
- Slight increase in collision frequency occurs. More molecules have enough energy that many more collisions have the necessary activation energy. Raises reaction rate a great deal.
- Catalytic nickel lowers the activation energy needed for a reaction. More often molecules have the

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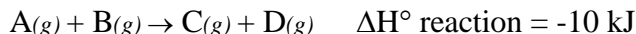
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needed energy when they collide. Reaction rate rises.

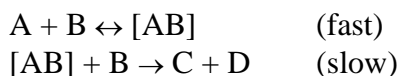
- (d) Greater surface area with powdered Ni. More catalytic sites means a greater rate.

1990 D

Consider the following general equation for a chemical reaction.



- (a) Describe the two factors that determine whether a collision between molecules of A and B results in a reaction.
- (b) How would a decrease in temperature affect the rate of the reaction shown above? Explain your answer.
- (c) Write the rate law expression that would result if the reaction proceeded by the mechanism shown below.

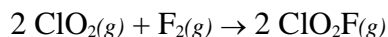


- (d) Explain why a catalyst increases the rate of a reaction but does not change the value of the equilibrium constant for that reaction.

Answer:

- (a) 1. The kinetic energy of the molecules - A certain minimum energy must be available for a reaction to occur. (activation energy)
2. The orientation of the molecules relative to one another - Even very energetic collisions may not lead to a reaction if the molecules are not oriented properly.
- (b) 1. A decrease in temperature would decrease the rate.
2. Fewer molecules would have the energy necessary for reaction (fewer effective collisions).
- (c) $\text{Rate} = k[A][B]^2$
- (d) 1. A catalyst increases the rate by providing an alternate pathway that has a lower activation energy.
2. The value of the equilibrium constant does not change, since a catalyst does not affect the energies (or concentrations) of the reactants and products.

1991 B



The following results were obtained when the reaction represented above was studied at 25°C.

Experiment	Initial $[\text{ClO}_2]$, (mol·L ⁻¹)	Initial $[\text{F}_2]$, (mol·L ⁻¹)	Initial Rate of Increase of $[\text{ClO}_2\text{F}]$, (mol·L ⁻¹ ·sec ⁻¹)
1	0.010	0.10	2.4×10^{-3}
2	0.010	0.40	9.6×10^{-3}
3	0.020	0.20	9.6×10^{-3}

- (a) Write the rate law expression for the reaction above.
- (b) Calculate the numerical value of the rate constant and specify the units.
- (c) In experiment 2, what is the initial rate of decrease of $[\text{F}_2]$?
- (d) Which of the following reaction mechanisms is consistent with the rate law developed in (a). Justify your choice.
- I. $\text{ClO}_2 + \text{F}_2 \leftrightarrow \text{ClO}_2\text{F}_2$ (fast)
 $\text{ClO}_2\text{F}_2 \rightarrow \text{ClO}_2\text{F} + \text{F}$ (slow)
 $\text{ClO}_2 + \text{F} \rightarrow \text{ClO}_2\text{F}$ (fast)
- II. $\text{F}_2 \rightarrow 2 \text{F}$ (slow)
 $2 (\text{ClO}_2 + \text{F} \rightarrow \text{ClO}_2\text{F})$ (fast)

Answer:

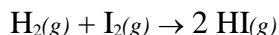
- (a) $[\text{F}_2]$ in expt. 2 is increased 4 times, the rate increases 4 times, \therefore 1st order in fluorine, $\text{rate} = k[\text{F}_2]^1$. In expt. 3, each reactant is doubled and the rate increases 4 times, \therefore 1st order in ClO_2 ,
 $\text{rate} = k[\text{ClO}_2]^1[\text{F}_2]^1$
- (b) $k = \frac{\text{initial rate}}{[\text{ClO}_2][\text{F}_2]}$
- $$k = \frac{2.4 \times 10^{-3} \text{ mol L}^{-1} \text{ sec}^{-1}}{(0.010)(0.10) \text{ mol}^2 \text{ L}^{-2}} = 2.4 \text{ L mol}^{-1} \text{ sec}^{-1}$$
- (c) $2 \text{ClO}_2 + \text{F}_2 \rightarrow 2 \text{ClO}_2\text{F}$
- $$-\frac{d[\text{F}_2]}{dt} = \frac{1}{2} \left(\frac{d[\text{ClO}_2\text{F}]}{dt} \right)$$
- $$= (9.6 \times 10^{-3})/2 = 4.8 \times 10^{-3} \text{ mol L}^{-1} \text{ sec}^{-1}$$
- (d) for step 1, rate forward = rate reverse,
 $k_f[\text{ClO}_2][\text{F}_2] = k_r[\text{ClO}_2\text{F}_2]$
 $[\text{ClO}_2\text{F}_2] = k_f/k_r[\text{ClO}_2][\text{F}_2]$
the overall rate is determined by the slowest step, step 2, \therefore $\text{rate} = k_2[\text{ClO}_2\text{F}_2]$

Kinetics

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$$\text{rate} = k_2 \frac{k_f}{k_r} [\text{ClO}_2][\text{F}_2] = k [\text{ClO}_2][\text{F}_2]$$

1992 D (Required)



For the exothermic reaction represented above, carried out at 298K, the rate law is as follows.

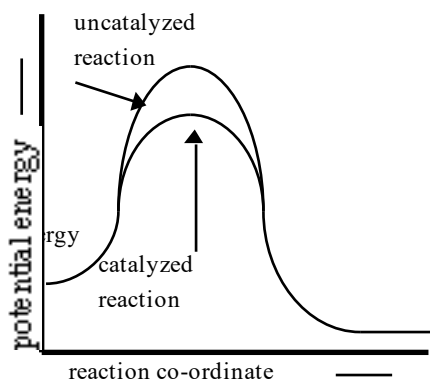
$$\text{Rate} = k[\text{H}_2][\text{I}_2]$$

Predict the effect of each of the following changes on the initial rate of the reaction and explain your prediction.

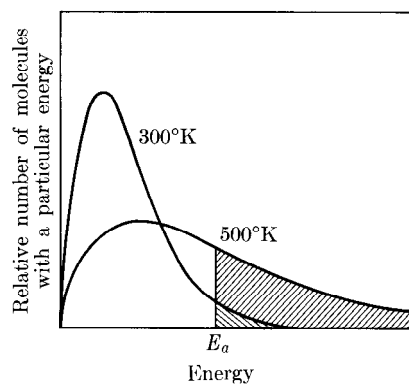
- Addition of hydrogen gas at constant temperature and volume
- Increase in volume of the reaction vessel at constant temperature
- Addition of catalyst. In your explanation, include a diagram of potential energy versus reaction coordinate.
- Increase in temperature. In your explanation, include a diagram showing the number of molecules as a function of energy.

Answer:

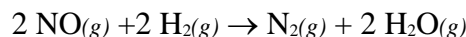
- Initial rate will increase. Relate increase in concentration of H_2 to an increase in collision rate or to the rate law.
- Initial rate will decrease. Decrease in the concentration of reactants.
- Initial rate will increase. Activation energy is lowered.



- Initial rate will increase. Maxwell-Boltzmann diagram



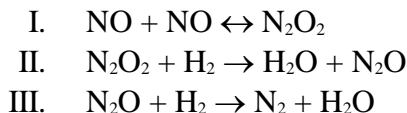
1994 B



Experiments were conducted to study the rate of the reaction represented by the equation above. Initial concentrations and rates of reaction are given in the table below.

Experiment	Initial Concentration (mol/L)		Initial Rate of Formation of N_2 (mol/L·min)
	[NO]	$[\text{H}_2]$	
1	0.0060	0.0010	1.8×10^{-4}
2	0.0060	0.0020	3.6×10^{-4}
3	0.0010	0.0060	0.30×10^{-4}
4	0.0020	0.0060	1.2×10^{-4}

- Determine the order for each of the reactants, NO and H_2 , from the data given and show your reasoning.
 - Write the overall rate law for the reaction.
- Calculate the value of the rate constant, k , for the reaction. Include units.
- For experiment 2, calculate the concentration of NO remaining when exactly one-half of the original amount of H_2 had been consumed.
- The following sequence of elementary steps is a proposed mechanism for the reaction.



Based on the data presented, which of the above is the rate-determining step? Show that the mechanism is consistent with

- the observed rate law for the reaction, and
- the overall stoichiometry of the reaction.

Answer:

- (a) (i) expt. 1 & 2 held $[\text{NO}]$ constant while $[\text{H}_2]$ doubled and the rate doubled, \therefore rate is 1st order with respect to $[\text{H}_2]$.

expt. 3 & 4 held $[\text{H}_2]$ constant while $[\text{NO}]$ doubled and the rate quadrupled, \therefore rate is 2nd order with respect to $[\text{NO}]$.

(ii) rate = $k [\text{H}_2] [\text{NO}]^2$ **OR**

(i) expt. 1, $1.8 \times 10^{-4} = k (0.0060)^m (0.0010)^n$

expt. 2, $3.6 \times 10^{-4} = k (0.0060)^m (0.0020)^n$

$$\frac{1.8 \times 10^{-4}}{0.0010^n} = \frac{3.6 \times 10^{-4}}{0.0020^n}; \quad 0.0020^n / 0.0010^n = 3.6 \times 10^{-4} / 1.8 \times 10^{-4}$$

$0.0020^n / 0.0010^n = 2$ where $n = 1$, $\therefore [\text{H}_2]$ is 1st order

expt. 3, $0.30 \times 10^{-4} = k (0.0010)^m (0.0060)^n$

expt. 4, $1.2 \times 10^{-4} = k (0.0020)^m (0.0060)^n$

$$\frac{0.30 \times 10^{-4}}{(0.0010)^m} = \frac{1.2 \times 10^{-4}}{(0.0020)^m}$$

$$\frac{(0.0020)^m}{(0.0010)^m} = \frac{1.2 \times 10^{-4}}{0.30 \times 10^{-4}}$$

$0.0020^m / 0.0010^m = 4$ where $m = 2$, $\therefore [\text{NO}]$ is 2nd order

- (b) using expt. 1,

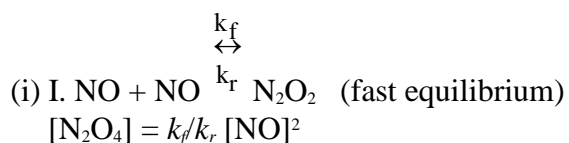
$$1.8 \times 10^{-4} \text{ mol/L} \cdot \text{min} = k \cdot 1.0 \times 10^{-3} \text{ mol/L} \times (6.0 \times 10^{-3} \text{ mol/L})^2$$

$$k = 5.0 \times 10^3 \text{ L}^2 \cdot \text{mol}^{-2} \cdot \text{min}^{-1}$$

- (c) $[\text{H}_2] = (0.0020 - 0.0010) \text{ M} = 0.0010 \text{ M}$

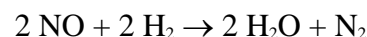
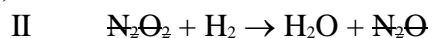
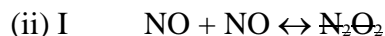
$$[\text{NO}] = (0.0060 - 0.0010) \text{ M} = 0.0050 \text{ M}$$

- (d) step II is the rate determining step

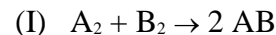


$$\text{rate} = \frac{k_f}{k_r} [\text{NO}]^2 \downarrow (k_3 [\text{H}_2]), k_f/k_r \times k_3 = k$$

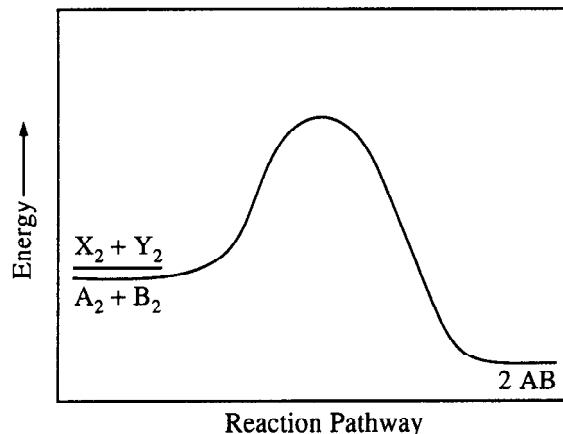
$$\text{rate} = k [\text{NO}]^2 [\text{H}_2]$$



1995 D



Two reactions are represented above. The potential-energy diagram for reaction I is shown below. The potential energy of the reactants in reaction II is also indicated on the diagram. Reaction II is endothermic, and the activation energy of reaction I is greater than that of reaction II.



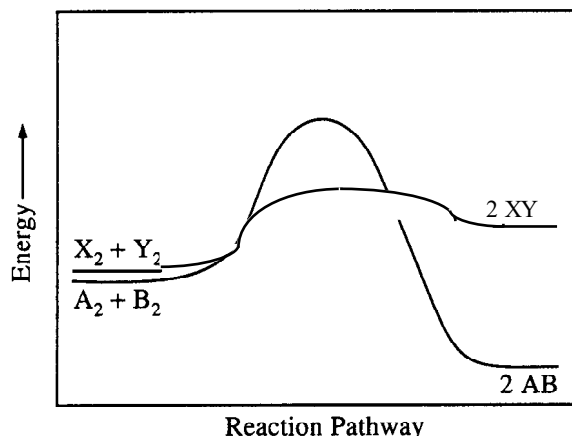
- (a) Complete the potential-energy diagram for reaction II on the graph above..
- (b) For reaction I, predict how each of the following is affected as the temperature is increased by 20°C. Explain the basis for each prediction.
- Rate of reaction
 - Heat of reaction
- (c) For reaction II, the form of the rate law is $\text{rate} = k[\text{X}_2]^m[\text{Y}_2]^n$. Briefly describe an experiment that can be conducted in order to determine the values of m and n in the rate law for the reaction.
- (d) From the information given, determine which reaction initially proceeds at the faster rate under the same conditions of concentration and temperature. Justify your answer.

Answer:

- (a)

Kinetics

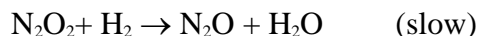
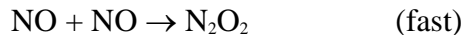
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factors will be the same for X_2 , Y_2 as for A_2 , B_2 , or that a similar mechanism is involved.

1996 D

The reaction between NO and H_2 is believed to occur in the following three-step process.



- Write a balanced equation for the overall reaction.
- Identify the intermediates in the reaction. Explain your reasoning.
- From the mechanism represented above, a student correctly deduces that the rate law for the reaction is $\text{rate} = k[NO]^2[H_2]$. The student then concludes that (1) the reaction is third-order and (2) the mechanism involves the simultaneous collision of two NO molecules and an H_2 molecule. Are conclusions (1) and (2) correct? Explain.
- Explain why an increase in temperature increases the rate constant, k , given the rate law in (c).

Answer:

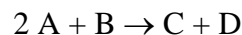
- $2 NO + 2 H_2 \rightarrow N_2 + 2 H_2O$
- N_2O_2 and N_2O ; they are part of the mechanism but are neither reactants nor products in the overall reaction.
- conclusion (1) is correct; the sum of the exponents in the rate law $(2 + 1) = 3$, the overall order.
conclusion (2) is incorrect; the three steps in the mechanism are all bimolecular collisions.
- an increase in temperature increases the rate and since there is no increase in concentrations then the rate constant, k , has to increase.

OR explanation using energy and frequency of collisions

OR explanation using Arrhenius equation

OR explanation using Maxwell-Boltzmann diagrams and activation energy.

1997 B



- Rate increases. At temperature increases, the molecules move faster and collide more frequently resulting in more possible reactions in the same time span as before. Also, and more importantly, they have more kinetic energy which results in a higher percentage of molecules that have sufficient activation energy when they collide, resulting in more effective collisions and reactions.
 - Heat of reaction is increased. The energy of the reactants is increased so the ΔH (difference between reactants and products) is larger.
- Conduct a series of experiments in which the $[Y_2]$ is kept constant and the $[X_2]$ is varied by a specific amount and measure the initial reaction rate. Repeat keeping $[X_2]$ constant and varying $[Y_2]$ as in the table below.

Expt. #	$[X_2]$	$[Y_2]$	Initial reaction rate
1	1	1	R_1
2	2	1	R_2
3	1	2	R_3

If $R_1 = R_2$ then $m = 0$, if $R_2 = 2R_1$ then $m = 1$, and if $R_2 = 4R_1$ then $m = 2$. Use similar logic to compare R_3 with R_1 and determine the value of n .

- Reaction II will initially be faster since it has the lower activation energy, a higher % of its molecules (since they are at the same temperature) will have sufficient energy to create the activated complex resulting in more effective collisions.

OR

It is not possible to determine which reaction has a faster rate without knowledge of other (pre-exponential) factors. It cannot be assumed these

Kinetics

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The following results were obtained when the reaction represented above was studied at 25°C.

Experiment	Initial [A]	Initial [B]	Initial Rate of Formation of C (mol L ⁻¹ min ⁻¹)
1	0.25	0.75	4.3×10 ⁻⁴
2	0.75	0.75	1.3×10 ⁻³
3	1.50	1.50	5.3×10 ⁻³
4	1.75	?	8.0×10 ⁻³

- Determine the order of the reaction with respect to A and to B. Justify your answer.
- Write the rate law for the reaction. Calculate the value of the rate constant, specifying units.
- Determine the initial rate of change of [A] in Experiment 3.
- Determine the initial value of [B] in Experiment 4.
- Identify which of the reaction mechanisms represented below is consistent with the rate law developed in part (b). Justify your choice.

- $A + B \rightarrow C + M$ *Fast*
 $M + A \rightarrow D$ *Slow*
- $B \leftrightarrow M$ *Fast equilibrium*
 $M + A \rightarrow C + X$ *Slow*
 $A + X \rightarrow D$ *Fast*
- $A + B \rightleftharpoons M$ *Fast equilibrium*
 $M + A \rightarrow C + X$ *Slow*
 $X \rightarrow D$ *Fast*

Answer:

- 1st order with respect to A, 1st order with respect to B.

Expt. 2: initial rate tripled compared to expt. 1

Expt. 2: [A] tripled compared to expt. 1.

Expt. 3: initial rate is 4x compared to expt. 2

Expt. 3: [B] is doubled compared to expt. 2 and [A] is double compared to expt. 2

OR

$$\text{initial rate}_1 = k_1[A_1]^m[B_1]^n$$

$$k_1[B_1]^n = \text{initial rate}_1/[A_1]^m$$

$$\text{initial rate}_2 = k_2[A_2]^m[B_2]^n$$

$$k_2[B_2]^n = \text{initial rate}_2/[A_2]^m \quad k_1[B_1]^n = k_2[B_2]^n$$

$$\text{initial rate}_1/[A_1]^m = \text{initial rate}_2/[A_2]^m$$

$$0.75^m/0.25^m = 1.3 \times 10^{-3}/4.3 \times 10^{-4} = 3.0/1 \text{ when } m = 1$$

$$k_2 = \text{initial rate}_2/[A_2]^1[B_2]^n$$

$$k_3 = \text{initial rate}_3/[A_3]^1[B_3]^n$$

$$k_2 = 1.3 \times 10^{-3}/0.75[B_2]^n = 1.7 \times 10^{-3}/[0.75]^n$$

$$k_3 = 5.3 \times 10^{-3}/1.50[B_3]^n = 3.5 \times 10^{-3}/[1.50]^n$$

$$1.7 \times 10^{-3}/[0.75]^n = 3.5 \times 10^{-3}/[1.50]^n$$

$$[1.50]^n/[0.75]^n = 3.5 \times 10^{-3}/1.7 \times 10^{-3} = 2.0/1 \text{ when } n = 1$$

$$(b) \text{ rate} = k[A]^1[B]^1; k = \text{rate}/[A][B]$$

$$k = 5.3 \times 10^{-3} \text{ mol L}^{-1} \text{ min}^{-1}/(1.50 \text{ mol L}^{-1} \times 1.50 \text{ mol L}^{-1})$$

$$k = 2.4 \times 10^{-3} \text{ L mol}^{-1} \cdot \text{min}^{-1}$$

- since there is a 2:1 ratio of A:C and since the formation of C is the disappearance of A then:

$$\text{rate} = -\text{rate} \times 2 = -(5.3 \times 10^{-3}) \times 2$$

$$= -1.1 \times 10^{-2} \text{ mol L}^{-1} \text{ min}^{-1}$$

$$(d) \text{ rate} = k[A]^1[B]^1; [B] = \text{rate}/k[A]$$

$$[B] = 8.0 \times 10^{-3} \text{ M min}^{-1}/(2.4 \times 10^{-3} \text{ L mol}^{-1} \cdot \text{min}^{-1} \times 1.75 \text{ M})$$

$$= 1.9 \text{ M}$$

- mechanism 2.

$$\text{step 1: rate}_f = \text{rate}_r; k_f[B] = k_r[M]$$

$$[M] = k_f/k_r \cdot [B]$$

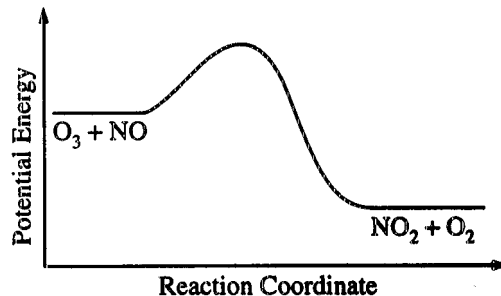
$$\text{step 2: rate} = k_3[M][A] = k k_r \cdot k_3[B][A]$$

$$k k_r \cdot k_3 = k$$

1998 D

Answer the following questions regarding the kinetics of chemical reactions.

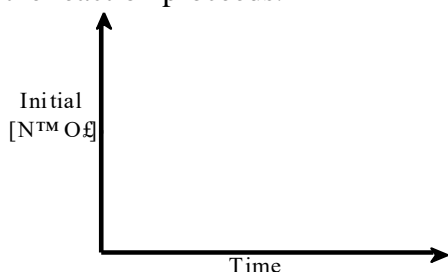
- The diagram below at right shows the energy pathway for the reaction $O_3 + NO \rightarrow NO_2 + O_2$. Clearly label the following directly on the diagram.



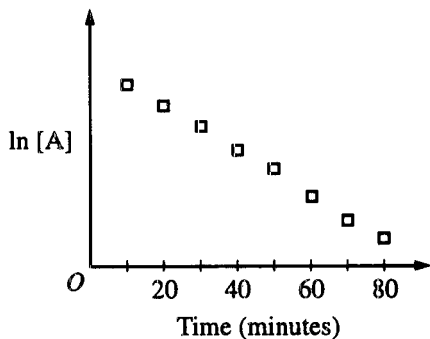
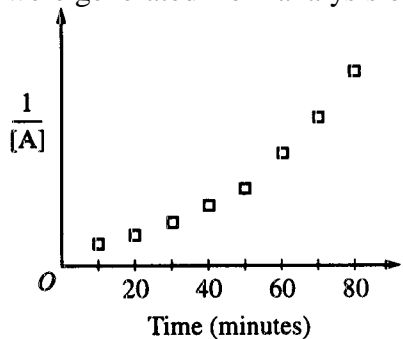
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- (i) The activation energy (E_a) for the forward reaction
- (ii) The enthalpy change (ΔH) for the reaction
- (b) The reaction $2 \text{N}_2\text{O}_5 \rightarrow 4 \text{NO}_2 + \text{O}_2$ is first order with respect to N_2O_5 .
- (i) Using the axes at right, complete the graph that represents the change in $[\text{N}_2\text{O}_5]$ over time as the reaction proceeds.



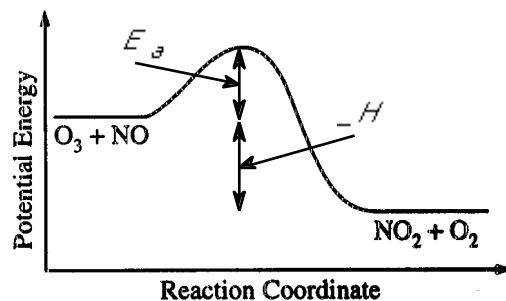
- (ii) Describe how the graph in (i) could be used to find the reaction rate at a given time, t .
- (iii) Considering the rate law and the graph in (i), describe how the value of the rate constant, k , could be determined.
- (iv) If more N_2O_5 were added to the reaction mixture at constant temperature, what would be the effect on the rate constant, k ? Explain.
- (c) Data for the chemical reaction $2\text{A} \rightarrow \text{B} + \text{C}$ were collected by measuring the concentration of A at 10-minute intervals for 80 minutes. The following graphs were generated from analysis of the data.



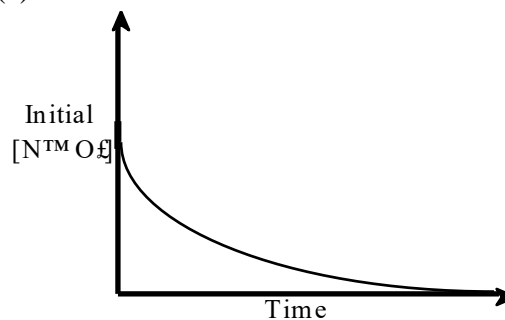
Use the information in the graphs above to answer the following.

- (i) Write the rate-law expression for the reaction. Justify your answer.
- (ii) Describe how to determine the value of the rate constant for the reaction.

Answer



(a)



(b) i

ii the rate at time, t , is the slope of the tangent to the curve at time t

iii since the reaction is 1st order:

$$\ln[\text{N}_2\text{O}_3]_t - \ln[\text{N}_2\text{O}_3]_o = -kt$$

$$-\ln \frac{[\text{N}_2\text{O}_3]_t}{[\text{N}_2\text{O}_3]_o}$$

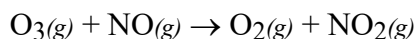
$$k = \frac{t}{t}$$

iv k would remain unchanged, it is temperature dependent, not concentration dependent.

(c)i since the graph of $\ln[\text{A}]$ is a straight line, this indicates that it is 1st order with respect to A, \therefore , rate = $k[\text{A}]$

ii $k = -\text{slope of the straight line of the } \ln[\text{A}] \text{ vs. time graph}$

2000 D Required



Consider the reaction represented above.

- (a) Referring to the data in the table below, calculate the standard enthalpy change, for the reaction at 25°C. Be sure to show your work.

Kinetics

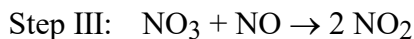
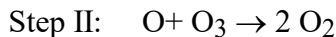
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	O _{3(g)}	NO(g)	NO _{2(g)}
Standard enthalpy of formation, ΔH_f° at 25°C (kJ mol ⁻¹)	143	90.	33

- (b) Make a qualitative prediction about the magnitude of the standard entropy change, ΔS° , for the reaction at 25°C. Justify your answer.
- (c) On the basis of your answers to parts (a) and (b), predict the sign of the standard free-energy change, for the reaction at 25°C. Explain your reasoning.
- (d) Use the information in the table below to write the rate-law expression for the reaction, and explain how you obtained your answer.

Experiment Number	Initial [O ₃] (mol L ⁻¹)	Initial [NO] (mol L ⁻¹)	Initial Rate of Formation of [NO ₂] (mol L ⁻¹ s ⁻¹)
1	0.0010	0.0010	x
2	0.0010	0.0020	2x
3	0.0020	0.0010	2x
4	0.0020	0.0020	4x

- (e) The following three-step mechanism is proposed for the reaction. Identify the step that must be the slowest in order for this mechanism to be consistent with the rate-law expression derived in part (d). Explain.



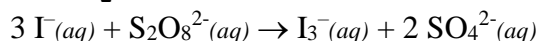
Answer:

- (a) $\Delta H = \sum \Delta H_f^\circ(\text{products}) - \sum \Delta H_f^\circ(\text{reactants})$
 $= (33 + 0) - (143 + 90) \text{ kJ} = -200 \text{ kJ}$
- (b) considering that there is: (1) no change in phase [*i.e.*, two kinds of gases on each side of the equation], (2) no change in the number of gas molecules [*i.e.*, 3 gas molecules on each side], (3) no change in the complexity of the gases [*i.e.*, a triatomic and biatomic molecule on each side], THEN

there should be a VERY SMALL magnitude change in the standard free entropy.

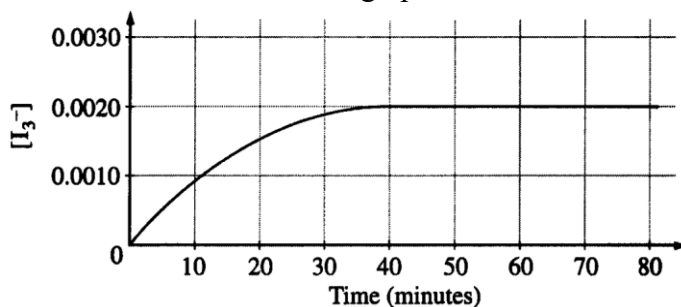
- (c) ΔG° should be negative (-). $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$; with $\Delta H^\circ < 0$ and $-T\Delta S^\circ$ as negligible, then ΔG° should be < 0 also.
- (d) $\text{rate} = k [\text{O}_3] [\text{NO}]$. Since the initial rate in expt. 2 doubles compared to expt. 1, while the initial [NO] is doubled and [O₃] stays the same, this indicates that the rate is first order with respect to NO.
- Since the initial rate in expt. 3 doubles compared to expt. 1, while the initial [O₃] is doubled and [NO] stays the same, this indicates that the rate is first order with respect to O₃.
- (e) slowest is step I. The slowest step is the rate determining step and, therefore, controls the speed of the overall reaction. The rate for step I = $k [\text{O}_3] [\text{NO}]$ and matches the rate law expression given in (d).

2001 D Required



Iodide ion, $\text{I}^-(\text{aq})$, reacts with peroxydisulfate ion, $\text{S}_2\text{O}_8^{2-}(\text{aq})$, according to the equation above. Assume that the reaction goes to completion.

- (a) Identify the type of reaction (combustion, disproportionation, neutralization, oxidation-reduction, precipitation, etc.) represented by the equation above. Also, give the formula of another substance that could convert $\text{I}^-(\text{aq})$ to $\text{I}_3^-(\text{aq})$.
- (b) In an experiment, equal volumes of 0.0120 M $\text{I}^-(\text{aq})$ and 0.0040 M $\text{S}_2\text{O}_8^{2-}(\text{aq})$ are mixed at 25°C. The concentration of $\text{I}_3^-(\text{aq})$ over the following 80 minutes is shown in the graph below.



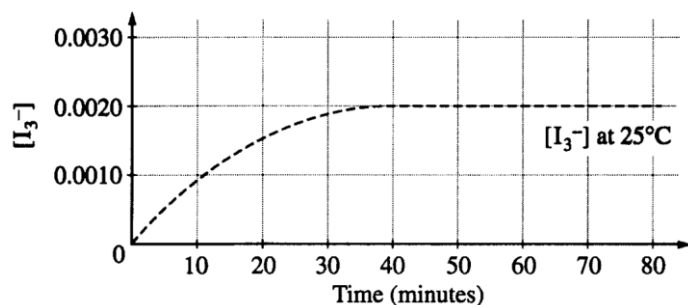
- (i) Indicate the time at which the reaction first reaches completion by marking an "X" on the

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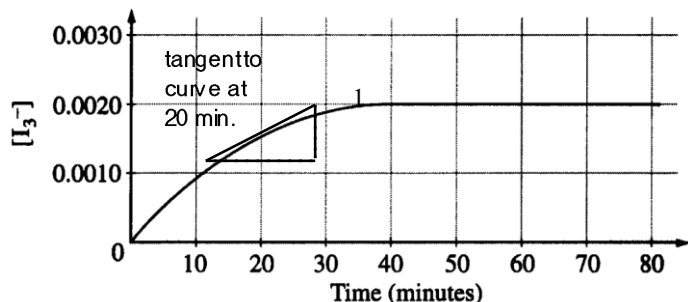
curve above at the point that corresponds to this time. Explain your reasoning.

- (ii) Explain how to determine the instantaneous rate of formation of $\text{I}_3^- (\text{aq})$ at exactly 20 minutes. Draw on the graph above as part of your explanation.
- (c) Describe how to change the conditions of the experiment in part (b) to determine the order of the reaction with respect to $\text{I}^- (\text{aq})$ and with respect to $\text{S}_2\text{O}_8^{2-} (\text{aq})$.
- (d) State clearly how to use the information from the results of the experiments in part (c) to determine the value of the rate constant, k , for the reaction.
- (e) On the graph below (which shows the results of the initial experiment as a dashed curve), draw in a curve for the results you would predict if the initial experiment were to be carried out at 35°C rather than at 25°C .



Answer :

- (a) redox; H_2O_2 , MnO_4^- , $\text{Cr}_2\text{O}_7^{2-}$, I_2
- (b) (i) 35-40 minutes. When mixed in equal amounts of solution, the maximum $[\text{I}_3^-]$ that can be reached is $2.00 \times 10^{-3} \text{ M}$. No further change occurs after this time.



- (ii) determine the slope of the tangent to the curve at 20 minutes.
- (c) set up a series of reactions in which the concentrations of each ion is changed and measure the initial reaction rate for each. Such as

expt.	$[\text{I}^-]$	$[\text{S}_2\text{O}_8^{2-}]$	Initial rate
1	0.0120	0.0040	R_1
2	0.0240	0.0040	R_2
3	0.0120	0.0080	R_3

to determine the order with respect to $[\text{S}_2\text{O}_8^{2-}]$,

$$\frac{R_1}{[\text{S}_2\text{O}_8^{2-}]^n} = \frac{R_3}{[\text{S}_2\text{O}_8^{2-}]^n}$$

solve for n , when:

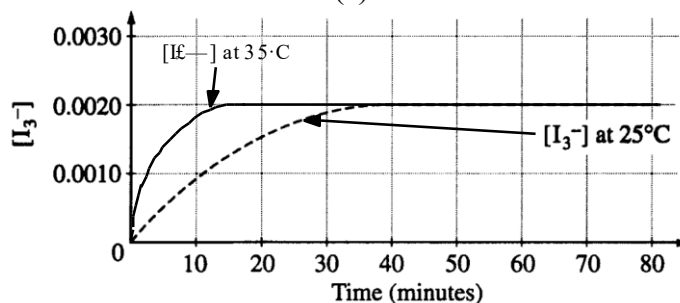
to determine the order with respect to $[\text{I}^-]$, solve

$$\frac{R_1}{[\text{I}^-]^m} = \frac{R_2}{[\text{I}^-]^m}$$

for m , when:

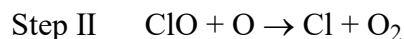
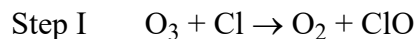
$$(d) R_1 = k [\text{I}^-]^m [\text{S}_2\text{O}_8^{2-}]^n; k = \frac{R_1}{[\text{I}^-]^m [\text{S}_2\text{O}_8^{2-}]^n}$$

(e)



2002 D

An environmental concern is the depletion of O_3 in Earth's upper atmosphere, where O_3 is normally in equilibrium with O_2 and O . A proposed mechanism for the depletion of O_3 in the upper atmosphere is shown below.

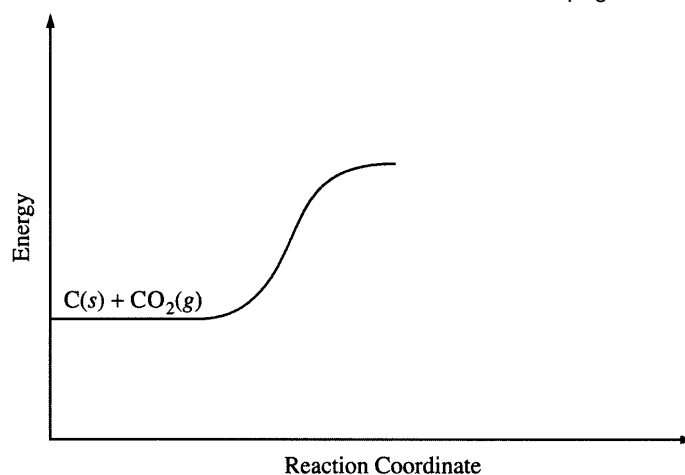


- (a) Write a balanced equation for the overall reaction represented by Step I and Step II above.
- (b) Clearly identify the catalyst in the mechanism above. Justify your answer.
- (c) Clearly identify the intermediate in the mechanism above. Justify your answer.
- (d) If the rate law for the overall reaction is found to be $\text{rate} = k[\text{O}_3][\text{Cl}]$, determine the following.
- (i) The overall order of the reaction
- (ii) Appropriate units for the rate constant, k

- (iii) The rate-determining step of the reaction, along with justification for your answer

Answer:

- (a) $\text{O}_3 + \text{O} \rightarrow 2 \text{O}_2$
- (b) Cl; used in step I and regenerated in step II, the amount at the end is the same as the beginning
- (c) ClO; product of step I and used in step II, an intermediate is a material that is produced by a step and consumed later, it does not show as either a product or reactant in the overall equation.
- (d) (i) second order overall
 (ii) k unit is $M^{-1}\text{time}^{-1}$
 (iii) step 1. the rate law applies to the concentration of the materials in the slowest step, the rate determining step.

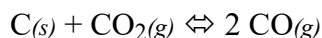


- (d) If the initial amount of $\text{C}(s)$ were doubled, what would be the effect on the percent of CO in the equilibrium mixture? Justify your answer.

Answer:

- (a) $\Delta S = +$; a low entropy solid and a single gas is changing to 2 molecules of gas, and increase in entropy
- (b) $\Delta H = +$; since an increase in yield is indicated by an increase in temperature, the reaction is likely endothermic.

2002 D

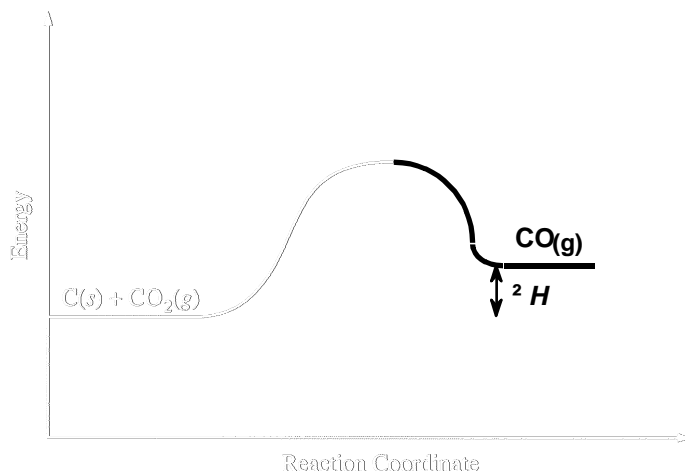


Carbon (graphite), carbon dioxide, and carbon monoxide form an equilibrium mixture, as represented by the equation above.

- (a) Predict the sign for the change in entropy, ΔS , for the reaction. Justify your prediction.
- (b) In the table below are data that show the percent of CO in the equilibrium mixture at two different temperatures. Predict the sign for the change in enthalpy, ΔH , for the reaction. Justify your prediction.

Temperature	% CO
700°C	60
850°C	94

- (c) Appropriately complete the potential energy diagram for the reaction by finishing the curve on the graph below. Also, clearly indicate ΔH for the reaction on the graph.



- (d) no change. a amount of solid will not change the equilibrium, the concentration of a solid is constant

2003 B



In a study of the kinetics of the reaction represented above, the following data were obtained at 298 K.

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Experiment	Initial [Br ⁻] (mol L ⁻¹)	Initial [BrO ₃ ⁻] (mol L ⁻¹)	Initial [H ⁺] (mol L ⁻¹)	Rate of Disappearance of BrO ₃ ⁻ (mol L ⁻¹ s ⁻¹)
1	0.00100	0.00500	0.100	2.50×10 ⁻⁴
2	0.00200	0.00500	0.100	5.00×10 ⁻⁴
3	0.00100	0.00750	0.100	3.75×10 ⁻⁴
4	0.00100	0.01500	0.200	3.00×10 ⁻³

(a) From the data given above, determine the order of the reaction for each reactant listed below. Show your reasoning.

- (i) Br⁻
- (ii) BrO₃⁻
- (iii) H⁺

(b) Write the rate law for the overall reaction.

(c) Determine the value of the specific rate constant for the reaction at 298 K. Include the correct units.

(d) Calculate the value of the standard cell potential, E°, for the reaction using the information in the table below.

Half-reaction	E° (V)
Br ₂ (l) + 2e ⁻ → 2 Br ⁻ (aq)	+1.065
BrO ₃ ⁻ (aq) + 6 H ⁺ (aq) + 5e ⁻ → 1/2 Br ₂ (l) + 3 H ₂ O(l)	+1.52

(e) Determine the total number of electrons transferred in the overall reaction.

Answer:

(a) (i) 1st order with respect to Br⁻; in experiments 1 and 2, a doubling of the [Br⁻] results in the doubling of the initial rate, and indication of 1st order

(ii) 1st order with respect to BrO₃⁻

using expt. 1 & 3

$$\text{rate}_1 = k[\text{Br}^-]^1[\text{BrO}_3^-]^m[\text{H}^+]^n$$

$$\frac{\text{rate}_1}{[\text{BrO}_3^-]^m} = k[\text{Br}^-]^1[\text{H}^+]^n$$

$$\text{rate}_3 = k[\text{Br}^-]^1[\text{BrO}_3^-]^m[\text{H}^+]^n$$

$$\frac{\text{rate}_3}{[\text{BrO}_3^-]^m} = k[\text{Br}^-]^1[\text{H}^+]^n$$

$$\frac{\text{rate}_1}{[\text{BrO}_3^-]^m} = \frac{\text{rate}_3}{[\text{BrO}_3^-]^m}$$

$$\frac{0.000250}{(0.0050)^m} = \frac{0.000375}{(0.0075)^m}$$

$$m = 1$$

(iii) 2nd order with respect to H⁺

using expt. 3 & 4

$$\text{rate}_3 = k[\text{Br}^-]^1[\text{BrO}_3^-]^1[\text{H}^+]^n$$

$$\frac{\text{rate}_3}{[\text{BrO}_3^-]^1[\text{H}^+]^n} = k[\text{Br}^-]^1$$

$$\text{rate}_4 = k[\text{Br}^-]^1[\text{BrO}_3^-]^1[\text{H}^+]^n$$

$$\frac{\text{rate}_4}{[\text{BrO}_3^-]^1[\text{H}^+]^n} = k[\text{Br}^-]^1$$

$$\frac{\text{rate}_3}{[\text{BrO}_3^-]^1[\text{H}^+]^n} = \frac{\text{rate}_4}{[\text{BrO}_3^-]^1[\text{H}^+]^n}$$

$$\frac{0.000375}{(0.00750)(0.100)^n} = \frac{0.00300}{(0.0150)(0.200)^n}$$

$$n = 2$$

$$(b) \text{ rate} = k[\text{Br}^-]^1[\text{BrO}_3^-]^1[\text{H}^+]^2$$

$$(c) 2.50 \times 10^{-4} = k(0.00100)(0.00500)(0.100)^2$$

$$k = 5000 \text{ mol}^{-3}\text{L}^3\text{s}^{-1}$$

$$(d) E^\circ = 1.52 + -1.065 \text{ V} = 0.455 \text{ V}$$

(e) the overall reaction can be made by reversing the first half-reaction and multiplying by 2.5, therefore, there are 5 electrons transferred.

2004 B

The first-order decomposition of a colored chemical species, X, into colorless products is monitored with a spectrophotometer by measuring changes in absorbance over time. Species X has a molar absorptivity constant of 5.00×10³ cm⁻¹M⁻¹ and the pathlength of the cuvette containing the reaction mixture is 1.00 cm. The data from the experiment are given in the table below.

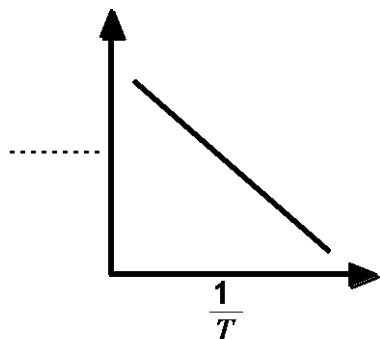
[X] (M)	Absorbance	Time (min)
?	0.600	0.0
4.00×10 ⁻⁵	0.200	35.0
3.00×10 ⁻⁵	0.150	44.2
1.50×10 ⁻⁵	0.075	?

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- Calculate the initial concentration of the unknown species.
- Calculate the rate constant for the first order reaction using the values given for concentration and time. Include units with your answers.
- Calculate the minutes it takes for the absorbance to drop from 0.600 to 0.075.
- Calculate the half-life of the reaction. Include units with your answer.
- Experiments were performed to determine the value of the rate constant for this reaction at various temperatures. Data from these experiments were used to produce the graph below, where T is temperature. This graph can be used to determine E_a , the activation energy.

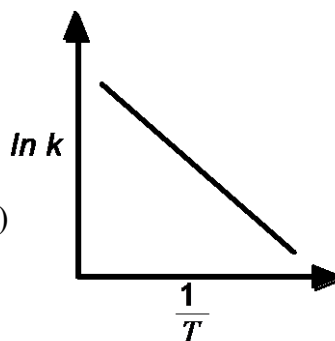
- Label the vertical axis of the graph
- Explain how to calculate the activation energy from this graph.



Answer:

- $A = abc$; $0.600 = (5000 \text{ cm}^{-1} \text{M}^{-1})(1.00 \text{ cm})(c)$
 $c = 1.20 \times 10^{-4} \text{ M}$
- $\ln[X]_t - \ln[X]_0 = -kt$
 $\ln(4.00 \times 10^{-5}) - \ln(1.20 \times 10^{-4}) = -k(35 \text{ min})$
 $k = 0.0314 \text{ min}^{-1}$
- $\ln[X]_t - \ln[X]_0 = -kt$
 $\ln[1.50 \times 10^{-5}] - \ln[1.20 \times 10^{-4}] = -0.0314 \text{ min}^{-1}t$
 $t = 66.2 \text{ min.}$
- $t_{1/2} = \frac{0.693}{k} = \frac{0.693}{0.0314} = 22.1 \text{ min}$

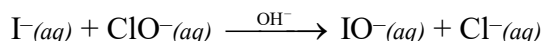
(e) (i)



- $\frac{-E_a}{R} = \text{slope of the line, multiply the slope by } -R \text{ to obtain } E_a$

2005 B

Answer the following questions related to the kinetics of chemical reactions.

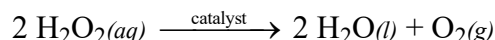


Iodide ion, I^- , is oxidized to hypoiodite ion, IO^- , by hypochlorite, ClO^- , in basic solution according to the equation above. Three initial-rate experiments were conducted; the results shown in the following table.

Experiment	$[\text{I}^-]$ (mol L ⁻¹)	$[\text{ClO}^-]$ (mol L ⁻¹)	Initial Rate of Formation of IO^- (mol L ⁻¹ s ⁻¹)
1	0.017	0.015	0.156
2	0.052	0.015	0.476
3	0.016	0.061	0.596

- Determine the order of the reaction with respect to each reactant listed below. Show your work.
 - $\text{I}^-(\text{aq})$
 - $\text{ClO}^-(\text{aq})$
- For the reaction,
 - write the rate law that is consistent with the calculations in part (a);
 - calculate the value of the specific rate constant, k , and specify units.

The catalyzed decomposition of hydrogen peroxide, $\text{H}_2\text{O}_2(\text{aq})$, is represented by the following equation.



The kinetics of the decomposition reaction were studied and the analysis of the results show that it is a first-

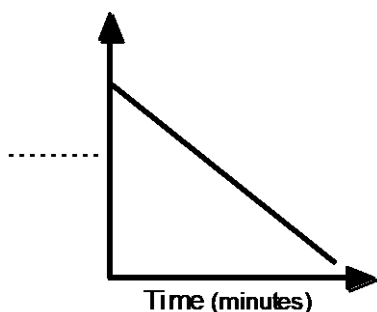
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order reaction. Some of the experimental data are shown in the table below.

[H ₂ O ₂] (mol L ⁻¹)	Time (minutes)
1.00	0.0
0.78	5.0
0.61	10.0

- (c) During the analysis of the data, the graph below was produced.



- Label the vertical axis of the graph
- What are the units of the rate constant, k , for the decomposition of H₂O₂(aq) ?
- On the graph, draw the line that represents the plot of the uncatalyzed first-order decomposition of 1.00 M H₂O₂(aq).

Answer:

- (a) (i) comparing expt. 1 to expt. 2, while the hypochlorite concentration remains constant, the iodide concentration is essentially tripled { $\frac{0.052}{0.017} =$

$\frac{3.06}{1}$ } and the initial rate is essentially tripled { $\frac{0.476}{0.156} = \frac{3.05}{1}$ }. This indicates a first order with respect to the iodide ion.

(ii) comparing expt. 1 to expt. 3, while the iodide concentration remains essentially constant (a 2.7% drop), the hypochlorite concentration is essentially quadrupled { $\frac{0.061}{0.015} = \frac{4.07}{1}$ } and the initial

rate is essentially quadrupled { $\frac{0.596}{0.156} = \frac{3.82}{1}$

}. This indicates a first order with respect to the hypochlorite ion.

OR

- (i) from experiments 1 & 2

$$\frac{\text{rate}_2}{\text{rate}_1} = \frac{k[\text{I}]_2^m[\text{ClO}^-]_2^n}{k[\text{I}]_1^m[\text{ClO}^-]_1^n}$$

$$\frac{0.476}{0.156} = \frac{k(0.052)^m(0.015)^n}{k(0.017)^m(0.015)^n}$$

$$3.05 = \frac{0.052^m}{0.017^m} = 3.1^m, \text{ where } m = 1$$

- (ii) from experiments 1 & 3

$$\frac{\text{rate}_3}{\text{rate}_1} = \frac{k[\text{I}]_3^m[\text{ClO}^-]_3^n}{k[\text{I}]_1^m[\text{ClO}^-]_1^n}$$

$$\frac{0.596}{0.156} = \frac{k(0.016)^m(0.061)^n}{k(0.017)^m(0.015)^n}$$

$$3.82 = (0.94) \frac{0.061^n}{0.015^n}$$

$$4.06 = 4.1^n, \text{ where } n = 1$$

- (b) (i) rate = $k[\text{I}^-][\text{ClO}^-]$

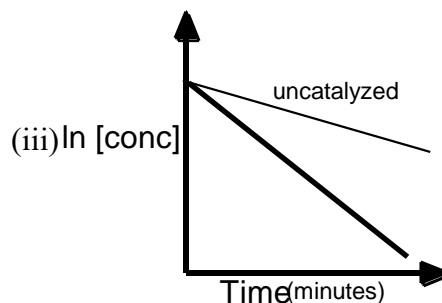
$$(ii) k = \frac{\text{rate}}{[\text{I}^-][\text{ClO}^-]} =$$

$$= \frac{0.156 \text{ mol L}^{-1} \text{ s}^{-1}}{(0.017 \text{ mol L}^{-1})(0.015 \text{ mol L}^{-1})} =$$

$$k = 610 \text{ L mol}^{-1} \text{ s}^{-1}$$

- (c) (i) vertical axis is "ln of [H₂O₂]"

- (ii) units for k are min⁻¹



(iii) ln [conc]