Unit 9 Kinetics Partner Multiple Choice Name(s):

1) Which of the following best helps explain why an increase in temperature increases the rate of a chemical reaction?
A) At higher temperatures, reactions have a lower activation energy. no
B) At higher temperatures, reactions have a higher activation energy. no
C) At higher temperatures, every collision results in the formation of product. no
(D) At higher temperatures, high-energy collisions happen more frequently.

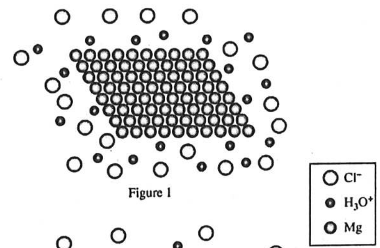


Figure 1

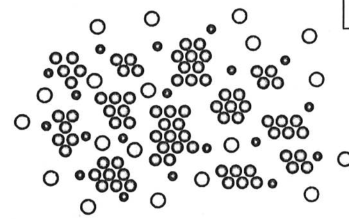


Figure 2

2) Two samples of Mg(s) of equal mass were placed in equal amounts of HCl(aq) contained in two separate reaction vessels. Particle representations of the mixing of Mg(s) and HCl(aq) in the two reaction vessels are shown in Figure 1 and 2. Water molecules are not included in the particle representations. Which of the reactions will initially proceed faster, and why?

A) The reaction in Figure 1, because the atoms of Mg are more concentrated than those in Figure 2 opposite
B) The reaction in Figure 1, because the Mg(s) in Figure 1 has a larger mass than the Mg(s) in Figure 2 it doesn't
(C) The reaction in Figure 2, because more Mg atoms are exposed to HCl(aq) in Figure 2 than in Figure 1
D) The reaction in Figure 2, because the Mg(s) in Figure 2 has ~~less~~ surface area than the Mg(s) in Figure 1
more

3) Which of the following will most likely increase the rate of the following reaction? $C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$
A) Decreasing the temperature of the reaction system (B) Adding a heterogeneous catalyst to the reaction system
C) Increasing the volume of the reaction vessel using a piston D) Removing some $H_2(g)$ from the reaction system

$$2\ NO_2(g) + F_2(g) \rightarrow 2\ NO_2F(g)$$

4) The rate law for the reaction represented by the equation above is rate = $k\,[NO_2][F_2]$. Which of the following could be the first elementary step of a two-step mechanism for the reaction if the first step is slow and the second step is fast?
A) $F_2(g) \rightarrow 2\ F(g)$ (B) $NO_2(g) + F_2(g) \rightarrow NO_2F(g) + F(g)$ both first order
C) $NO_2(g) + F(g) \rightarrow NO_2F(g)$ D) $2\ NO_2(g) + F_2(g) \rightarrow 2\ NO_2F(g)$

$$NO_2(g) + CO(g) \rightarrow NO_2(g) + CO_2(g)$$

5) The reaction between $NO_2(g)$ and $CO(g)$ is represented above. The elementary steps of a proposed reaction mechanism are represented here:
Step 1: $2\ NO_2(g) \rightarrow NO(g) + NO_3(g)$ (*slow*)
Step 2: $NO_3(g) + CO(g) \rightarrow NO_2(g) + CO_2(g)$ (*fast*)

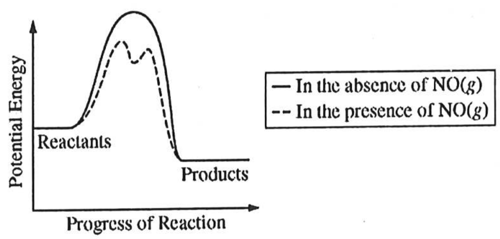
CO is not in rate determining step

Which of the following is the rate law for the overall reaction that is consistent with the proposed mechanism?
A) Rate = $k\,[NO_2][CO]$ (B) Rate = $k\,[NO_2]^2$ C) Rate = $k\,[NO_3][CO]$ D) Rate = $k\,[NO_2][NO_3][CO]$

$$Cl^-(aq) + ClO^-(aq) + 2\ H^+(aq) \rightarrow Cl_2(g) + H_2O(l)$$

6) What effect will increasing $[H^+]$ at constant temperature have on the reaction represented above?
A) The activation energy of the reaction will increase.
B) The activation energy of the reaction will decrease. no
(C) The frequency of collisions between $H^+(aq)$ ions and $ClO^-(aq)$ ions will increase.
D) The value of the rate constant will increase.

7)



The decomposition of $O_3(g)$ in the upper atmosphere is represented by the equation $O_3(g) + O(g) \rightarrow 2\ O_2(g)$. The potential energy diagram for the decomposition of $O_3(g)$ in the presence and absence of NO(g) is given. Which of the following mechanisms for the catalyzed reaction is consistent with the equation and diagram?

A) $2\ O_3(g) + 2\ NO(g) \rightarrow 4\ O_2(g) + N_2(g)$ *slow*

(B) $O_3(g) + NO(g) \rightarrow NO_2(g) + O_2(g)$ *slow*
$NO_2(g) + O(g) \rightarrow NO(g) + O_2(g)$ *fast*

C) $NO_2(g) + O_3(g) \rightarrow NO(g) + 2\ O_2(g)$ *slow*
$NO(g) + O(g) \rightarrow NO_2(g)$ *fast*

D) $NO_2(g) + O(g) \rightarrow NO_3(g)$ *slow*
$NO_3(g) + O_3(g) \rightarrow NO_2(g) + 2\ O_2(g)$ *fast*

NO is a catalyst consumed then reformed

Questions 8-10 refer to the investigation described below. $C_{25}H_{30}N_3^+\,(aq) + OH^-\,(aq) \rightarrow C_{25}H_{30}N_3OH\,(aq)$

violet *colorless*

The reaction between $C_{25}H_{30}N_3^+(aq)$ and $OH^-(aq)$, as represented above, is first order with respect to $C_{25}H_{30}N_3^+(aq)$ in the presence of excess $OH^-(aq)$. A 10.0 mL sample of 0.10 M NaOH(aq) is mixed with a 10.0 mL sample of 2.5×10^{-5} M $C_{25}H_{30}N_3^+(aq)$. A 5.0 mL sample of the mixture is quickly transferred to a clean cuvette and placed in a spectrophotometer, and the progress of the reaction is measured. The data are given in the table below.

Time (s)	0	30	60	90	120	150	180	210	240	270	300
Absorbance	0.62	0.54	0.47	0.41	0.36	0.31	0.27	0.23	0.20	0.17	0.15

Half-Life

8) Approximately how long did it take for 75 percent of the initial amount of $C_{25}H_{30}N_3^+(aq)$ to react?

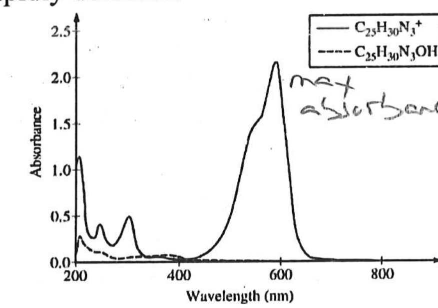
A) 75 s B) 225 s C) 300 s D) 600 s

9) What would be the effect on the reaction rate if the solution of $C_{25}H_{30}N_3^+(aq)$ is diluted by a factor of two?

A) It would be higher. B) It would be lower. simple Kinetics

C) It would not change. D) It would initially be higher but then rapidly decrease.

10) To choose a wavelength to analyze the progress of the reaction, a student records the absorbance spectra of both $C_{25}H_{30}N_3^+(aq)$ and $C_{25}H_{30}N_3OH(aq)$ in the range of 200-800 nm. The two spectra are presented in the graph.



max absorbance

The student wants to use the spectrophotometer to measure $[C_{25}H_{30}N_3^+]$ with the greatest sensitivity as the reaction progresses. Which of the following indicates the best wavelength setting and explains why it is best?

A) 205 nm, because the colorless form of the molecule will absorb significantly at this wavelength
B) 205 nm, because both forms of the molecule will absorb significantly at this wavelength
C) 590 nm, because only the violet form of the molecule will absorb significantly at this wavelength
D) 590 nm, because this wavelength falls in the violet region of the visible light spectrum it doesn't

Use this equation for #11 - 13: $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

11) Nitrogen monoxide and oxygen gas were combined in a flask at 25°C and allowed to react as shown above. The concentration of the reactants were varied according to the table to the right, and initial rates were calculated. Which of the following is the rate law for the reaction?

Trial	$[O_2]$	[NO]	Initial Rate of Formation of NO_2 ($M\,s^{-1}$)
1	0.020	0.050	x4 0.038
2	0.020	0.100	0.152 4
3	0.080	0.100	0.608

A) Rate = $k[O_2]^2[NO]^2$
B) Rate = $k[NO]^2$
C) Rate = $k[O_2][NO]^2$
D) Rate = $k[O_2][NO]$

12) Which mechanism agrees with the rate law above?

A) Step 1: $NO(g) + O_2(g) \rightarrow NO_2(g) + O(g)$ (*slow*)
Step 2: $NO(g) + O(g) \rightarrow NO_2(g)$ (*fast*)
first order w/ respect to NO

B) Step 1: $NO(g) + O_2(g) \rightarrow NO_2(g) + O(g)$ (*fast*)
Step 2: $NO(g) + O(g) \rightarrow NO_2(g)$ (*slow*)
2nd order w/ respect to NO

13) Which substance is acting as intermediate in each mechanism?

A) NO B) NO_2 C) O D) O_2

1) A rate study of the reaction represented below was conducted at 25°C. The data that were obtained are shown in the table. 1999 3

$$2\ NO(g) + Br_2(g) \rightarrow 2\ NOBr(g)$$

Experiment	Initial [NO] (mol L^{-1})	Initial [Br_2] (mol L^{-1})	Initial Rate of Appearance of NOBr (mol L^{-1} s^{-1})
1	0.0160	0.0120	3.24 x 10^{-4}
2	0.0160	0.0240	6.38 x 10^{-4}
3	0.0320	0.0060	6.42 x 10^{-4}

4 (
2 (
2 (

a) Calculate the initial rate of disappearance of $Br_2(g)$ in experiment 1. *(1pt)*

2:1 ratio

$$\frac{3.24\times10^{-4}}{2} = \boxed{-1.62\times10^{-4}\ \frac{mol}{L\cdot s}}$$

b) Determine the order of the reaction with respect to each reactant, $Br_2(g)$ and NO(g). In each case, explain your reasoning. *(3pts)*

1st order w/ respect to Br_2. when [NO] was held constant, [Br_2] was doubled and the rate also doubled

2nd order w/ respect to [NO]. The rate remained constant when [NO] was doubled and [Br_2] was decreased by 4. The rate would stay constant if [NO] is 2nd order.

c) For the reaction,

i) write the rate law that is consistent with the data, and *(1pt)*

$$R = k[NO]^2[Br_2]$$

ii) calculate the value of the specific rate constant, *k*, and specify units. *(2pts)*

$$K = \frac{R}{[NO]^2[Br_2]} = \frac{3.24\times10^{-4}\ \frac{M}{s}}{[0.0160\,M]^2\,[0.0120\,M]} = \boxed{105\ M^{-2}s^{-1}}$$

d) The following mechanism was proposed for the reaction:

$Br_2(g) + NO(g) \rightarrow NOBr_2(g)$ *slow*
$NOBr_2(g) + NO(g) \rightarrow 2\ NOBr(g)$ *fast*

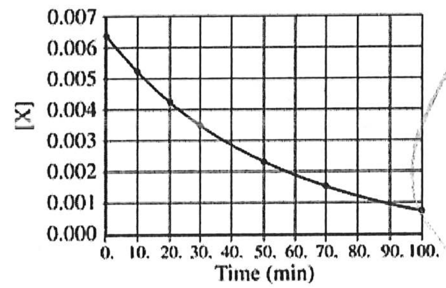
Is this mechanism consistent with the given experimental observations? Justify your answer. *(2pt)*

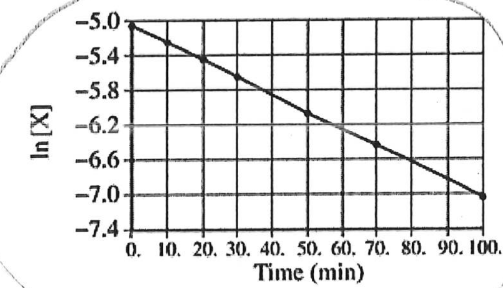
No, the mechanism suggests that NO is first order based on the slow step. It is 2nd order, so the second step should be the rate-determining step

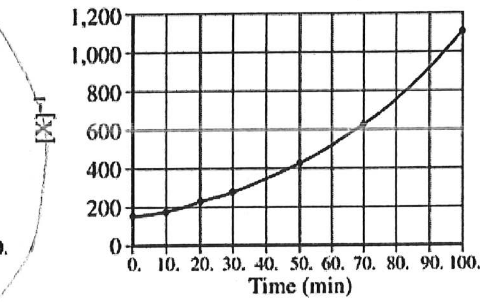
$X \rightarrow 2Y + Z$

Time (minutes)	[X] (mol L^{-1})	ln [X]	$[X]^{-1}$ (L mol^{-1})
0	0.00633	−5.062	158
10.	0.00520	−5.259	192
20.	0.00427	−5.456	234
30.	0.00349	−5.658	287
50.	0.00236	−6.049	424
70.	0.00160	−6.438	625
100.	0.000900	−7.013	1,110

5) The decomposition of gas X to produce gases Y and Z is represented by the equation above. In a certain experiment, the reaction took place in a 5.00 L flask at 428 K. Data from this experiment were used to produce the information in the table below, which is plotted in the graphs that follow. 2005B 3







a) How many moles of X were initially in the flask? *(1pt)*

$M = \frac{mol}{L}$ $\quad 0.00633 = \frac{mol}{5.00}$ $\quad \boxed{0.0317 \text{ mol } X}$

b) How many molecules of Y were produced in the first 20. minutes of the reaction? *(2pts)*

$0.00427 = \frac{mol}{5.00} = 0.0214 \text{ mol}$ $\quad 0.0317 \text{ mol} - 0.0214 \text{ mol} = 0.0104$ mol used

$0.0104 \text{ mol } X \times \frac{2 \text{ mol } Y}{1 \text{ mol } X} \times \frac{6.022 \times 10^{23} \text{ molec } Y}{1 \text{ mol } Y} = \boxed{1.25 \times 10^{22} \text{ molecules } Y}$

c) What is the order of this reaction with respect to X? Justify your answer. *(1pt)*

1^{st} Order – Linear when ln[X]

d) Write the rate law for this reaction. *(1pt)*

Rate = k[X]

e) Calculate the specific rate constant for this reaction. Specify units. *(2pts)*

$\ln[A]_t - \ln[A]_0 = -kt$

$-5.456 - (-5.062) = -k(1200 s)$

$\boxed{k = 3.28 \times 10^{-4} \; s^{-1}}$

or

$0.0197 \; min^{-1}$

f) Calculate the concentration of X in the flask after a total of 150. minutes of reaction. *(2pts)*

$\ln[A]_t - \ln[A]_0 = -kt$

$\ln[A]_t - (-5.062) = -(3.28 \times 10^{-4} s^{-1})(9000 s)$

$\ln[A]_t = -8.014$

$\boxed{[X] = 3.31 \times 10^{-4} M}$

$H_2(g) + Cl_2(g) \rightarrow 2\,HCl(g)$

9) The table gives data for a reaction rate study of the reaction represented above.

2010B 6

Experiment	Initial $[H_2]$ (mol L^{-1})	Initial $[Cl_2]$ (mol L^{-1})	Initial Rate of Formation of HCl (mol L^{-1} s^{-1})
1	0.00100 ×2	0.000500 c	1.82×10^{-12} ×2
2	0.00200	0.000500	3.64×10^{-12}
3	0.00200 c	0.000250 ÷2	1.82×10^{-12} ÷2

a) Determine the order of the reaction with respect to H_2 and justify your answer. *(1pt)*

1st Order - when $[H_2]$ is doubled and $[Cl_2]$ remains constant, the rate doubles

b) Determine the order of the reaction with respect to Cl_2 and justify your answer. *(1pt)*

1st Order - when $[Cl_2]$ is halved and $[H_2]$ is constant, the rate is also halved.

c) Write the overall rate law for the reaction. $R = K[H_2][Cl_2]$ *(1pt)*

d) Write (determine) the units of the rate constant. *(1pt)*

$\frac{M}{s} = K(M)(M)$ $\boxed{K = \frac{1}{M \cdot s}}$ or $\frac{L}{mol \cdot s}$

e) Predict the initial rate of the reaction if the initial concentration of H_2 is 0.00300 mol L^{-1} and the initial concentration of Cl_2 is 0.000500 mol L^{-1}. *(1pt)*

Find K:

$1.82 \times 10^{-12}\ \frac{M}{s} = K(0.00100M)(0.000500M)$

$K = 3.64 \times 10^{-6}\ \frac{1}{M \cdot s}$

Find Rate:

$R = \left(3.64 \times 10^{-6}\ \frac{1}{M \cdot s}\right)(0.00300M)(0.000500M)$

$\boxed{R = 5.46 \times 10^{-12}\ \frac{M}{s}}$

(or do 3× the first rate based on data)

The gas-phase decomposition of nitrous oxide has the following two-step mechanism.

Step 1: $N_2O \rightarrow N_2 + \cancel{O}$

Step 2: $\cancel{O} + N_2O \rightarrow N_2 + O_2$

f) Write the balanced equation for the overall reaction. *(1pt)*

$2N_2O \rightarrow 2N_2 + O_2$

g) Is the oxygen atom, O, a catalyst for the reaction or is it an intermediate? Explain. *(1pt)*

Intermediate - produced then immediately reacted

h) Identify the slower step in the mechanism if the rate law for the reaction was determined to be *rate* = *k* $[N_2O]$. Justify your answer. *(1pt)*

Step 1 - only 1 N_2O appears. This is consistent with the first order rate law.

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 5.00×10^3 cm^{-1} M^{-1} and the path length of the cuvette containing the reaction mixture is 1.00 cm. The data from the experiment are given in the table. 2004 3

[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	?

a) Calculate the initial concentration of the colored species. *(1pt)*

$A = \varepsilon b c$

$$c = \frac{A}{\varepsilon b} = \frac{(0.600)}{(5.00\times10^3\ cm^{-1}M^{-1})(1.00\ cm)} = \boxed{1.20\times10^{-4}\ M}$$

b) Calculate the rate constant for the first-order reaction using the values given for concentration and time. Include units with your answer. *(2pts)*

$\ln[A]_t - \ln[A]_0 = -kt$

$\boxed{k = 0.0314\ min^{-1}}$

$\ln[4.00\times10^{-5}\ \cancel{M}] - \ln[1.20\times10^{-4}\ \cancel{M}] = -k(35\ min)$

c) Calculate the number of minutes it takes for the absorbance to drop from 0.600 to 0.075. *(2pts)*

$(1.20\times10^{-4}\ M)$ $(1.50\times10^{-5}\ M)$

$\ln[1.50\times10^{-5}\ M] - \ln[1.20\times10^{-4}] = -(0.0314)(t)$

$\boxed{t = 66.2\ min}$

d) Calculate the half-life of the reaction. Include units with your answer. *(2pts)*

$$t_{\frac{1}{2}} = \frac{0.693}{k} = \frac{0.693}{0.0314\ min^{-1}} = \boxed{22.1\ min}$$