

Chapter 13 Chemical Kinetics

Student: _____

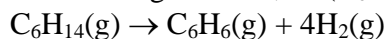
1. The units of "reaction rate" are

- A. $\text{L mol}^{-1} \text{s}^{-1}$.
- B. $\text{L}^2 \text{mol}^{-2} \text{s}^{-1}$.
- C. s^{-1} .
- D. s^{-2} .
- E. $\text{mol L}^{-1} \text{s}^{-1}$.

2. For the reaction $\text{BrO}_3^- + 5\text{Br}^- + 6\text{H}^+ \rightarrow 3\text{Br}_2 + 3\text{H}_2\text{O}$ at a particular time, $-\Delta[\text{BrO}_3^-]/\Delta t = 1.5 \times 10^{-2} \text{ M/s}$. What is $-\Delta[\text{Br}^-]/\Delta t$ at the same instant?

- A. 13 M/s
- B. $7.5 \times 10^{-2} \text{ M/s}$
- C. $1.5 \times 10^{-2} \text{ M/s}$
- D. $3.0 \times 10^{-3} \text{ M/s}$
- E. 330 M/s

3. For the following reaction, $\Delta P(\text{C}_6\text{H}_{14})/\Delta t$ was found to be $-6.2 \times 10^{-3} \text{ atm/s}$.



Determine $\Delta P(\text{H}_2)/\Delta t$ for this reaction at the same time.

- A. $6.2 \times 10^{-3} \text{ atm/s}$
- B. $1.6 \times 10^{-3} \text{ atm/s}$
- C. $2.5 \times 10^{-2} \text{ atm/s}$
- D. $-1.6 \times 10^{-3} \text{ atm/s}$
- E. $-2.5 \times 10^{-2} \text{ atm/s}$

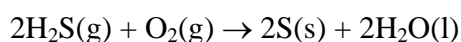
4. For the reaction $\text{C}_6\text{H}_{14}(\text{g}) \rightarrow \text{C}_6\text{H}_6(\text{g}) + 4\text{H}_2(\text{g})$, $\Delta P(\text{H}_2)/\Delta t$ was found to be $2.5 \times 10^{-2} \text{ atm/s}$, where $\Delta P(\text{H}_2)$ is the change in pressure of hydrogen. Determine $\Delta P(\text{C}_6\text{H}_{14})/\Delta t$ for this reaction at the same time.

- A. $2.5 \times 10^{-2} \text{ atm/s}$
- B. $-6.2 \times 10^{-3} \text{ atm/s}$
- C. $-2.5 \times 10^{-2} \text{ atm/s}$
- D. 0.10 atm/s
- E. $6.2 \times 10^{-3} \text{ atm/s}$

5. For the hypothetical reaction $A + 3B \rightarrow 2C$, the rate of appearance of C given by $(\Delta[C]/\Delta t)$ may also be expressed as

- A. $\Delta[C]/\Delta t = \Delta[A]/\Delta t$
- B. $\Delta[C]/\Delta t = -(3/2) \Delta[B]/\Delta t$
- C. $\Delta[C]/\Delta t = -(2/3) \Delta[B]/\Delta t$
- D. $\Delta[C]/\Delta t = -(1/2) \Delta[A]/\Delta t$

6. For the overall chemical reaction shown below, which one of the following statements can be rightly assumed?



- A. The reaction is third-order overall.
 - B. The reaction is second-order overall.
 - C. The rate law is, $\text{rate} = k[\text{H}_2\text{S}]^2 [\text{O}_2]$.
 - D. The rate law is, $\text{rate} = k[\text{H}_2\text{S}] [\text{O}_2]$.
 - E. The rate law cannot be determined from the information given.
7. The reaction $A + 2B \rightarrow \text{products}$ has been found to have the rate law, $\text{rate} = k[\text{A}] [\text{B}]^2$. While holding the concentration of A constant, the concentration of B is increased from x to $3x$. Predict by what factor the rate of reaction increases.

- A. 3
- B. 6
- C. 9
- D. 27
- E. 30

8. For the hypothetical reaction $A + 3B \rightarrow 2C$, the rate should be expressed as

- A. $\text{rate} = \Delta[\text{A}]/\Delta t$
- B. $\text{rate} = -\Delta[\text{C}]/\Delta t$
- C. $\text{rate} = -3(\Delta[\text{B}]/\Delta t)$
- D. $\text{rate} = (1/2)(\Delta[\text{C}]/\Delta t)$
- E. $\text{rate} = (1/3)(\Delta[\text{B}]/\Delta t)$

9. The reaction $A + 2B \rightarrow \text{products}$ has the rate law, $\text{rate} = k[\text{A}][\text{B}]^3$. If the concentration of B is doubled while that of A is unchanged, by what factor will the rate of reaction increase?

- A. 2
- B. 4
- C. 6
- D. 8
- E. 9

10. The reaction $A + 2B \rightarrow \text{products}$ was found to have the rate law, $\text{rate} = k[A][B]^2$. Predict by what factor the rate of reaction will increase when the concentration of A is doubled and the concentration of B is also doubled.
- A. 2
B. 4
C. 6
D. 8
E. 9
11. The reaction $A + 2B \rightarrow \text{products}$ was found to follow the rate law, $\text{rate} = k[A]^2[B]$. Predict by what factor the rate of reaction will increase when the concentration of A is doubled, the concentration of B is tripled, and the temperature remains constant.
- A. 5
B. 6
C. 12
D. 18
E. none of these
12. Appropriate units for a first-order rate constant are
- A. M/s.
B. $1/M \cdot s$.
C. 1/s.
D. $1/M^2 \cdot s$.
13. It takes 42.0 min for the concentration of a reactant in a first-order reaction to drop from 0.45 M to 0.32 M at 25°C. How long will it take for the reaction to be 90% complete?
- A. 13.0 min
B. 86.0 min
C. 137 min
D. 222 min
E. 284 min

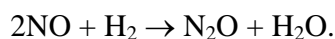
14. Nitric oxide gas (NO) reacts with chlorine gas according to the equation
 $\text{NO} + \frac{1}{2}\text{Cl}_2 \rightarrow \text{NOCl}$.

The following initial rates of reaction have been measured for the given reagent concentrations.

<u>Expt. #</u>	<u>Rate (M/hr)</u>	<u>NO (M)</u>	<u>Cl₂ (M)</u>
1	1.19	0.50	0.50
2	4.79	1.00	0.50
3	9.59	1.00	1.00

Which of the following is the rate law (rate equation) for this reaction?

- A. rate = $k[\text{NO}]$
B. rate = $k[\text{NO}][\text{Cl}_2]^{1/2}$
C. rate = $k[\text{NO}][\text{Cl}_2]$
D. rate = $k[\text{NO}]^2[\text{Cl}_2]$
E. rate = $k[\text{NO}]^2[\text{Cl}_2]^2$
15. Use the following data to determine the rate law for the reaction



<u>Expt. #</u>	<u>[NO]₀</u>	<u>[H₂]₀</u>	<u>Initial rate</u>
1	0.021	0.065	1.46 M/min
2	0.021	0.260	1.46 M/min
3	0.042	0.065	5.84 M/min

- A. rate = $k[\text{NO}]$
B. rate = $k[\text{NO}]^2$
C. rate = $k[\text{NO}][\text{H}_2]$
D. rate = $k[\text{NO}]^2[\text{H}_2]$
E. rate = $k[\text{NO}]^2[\text{H}_2]^2$

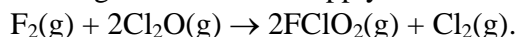
16. The data below were determined for the reaction $\text{S}_2\text{O}_8^{2-} + 3\text{I}^-(\text{aq}) \rightarrow 2\text{SO}_4^{2-} + \text{I}_3^-$.

<u>Expt. #</u>	<u>$[\text{S}_2\text{O}_8^{2-}]$</u>	<u>$[\text{I}^-]$</u>	<u>Initial Rate</u>
1	0.038	0.060	1.4×10^{-5} M/s
2	0.076	0.060	2.8×10^{-5} M/s
3	0.076	0.030	1.4×10^{-5} M/s

The rate law for this reaction must be:

- A. $\text{rate} = k[\text{S}_2\text{O}_8^{2-}][\text{I}^-]^3$.
B. $\text{rate} = k[\text{S}_2\text{O}_8^{2-}]$.
C. $\text{rate} = k[\text{S}_2\text{O}_8^{2-}]^2[\text{I}^-]^2$.
D. $\text{rate} = k[\text{I}^-]$.
E. $\text{rate} = k[\text{S}_2\text{O}_8^{2-}][\text{I}^-]$.
17. At 25°C the rate constant for the first-order decomposition of a pesticide solution is $6.40 \times 10^{-3} \text{ min}^{-1}$. If the starting concentration of pesticide is 0.0314 M, what concentration will remain after 62.0 min at 25°C ?
- A. 1.14×10^{-1} M
B. 47.4 M
C. -8.72.0 M
D. 2.11×10^{-2} M
E. 2.68×10^{-2} M
18. A certain first-order reaction $\text{A} \rightarrow \text{B}$ is 25% complete in 42 min at 25°C . What is the half-life of the reaction?
- A. 21 min
B. 42 min
C. 84 min
D. 120 min
E. 101 min

19. The following initial rate data apply to the reaction



Expt. #	[F ₂] (M)	[Cl ₂ O] (M)	Initial rate (M/s)
1	0.05	0.010	5.0×10^{-4}
2	0.05	0.040	2.0×10^{-3}
3	0.10	0.010	1.0×10^{-3}

Which of the following is the rate law (rate equation) for this reaction?

- A. rate = $k[\text{F}_2]^2[\text{Cl}_2\text{O}]^4$
B. rate = $k[\text{F}_2]^2[\text{Cl}_2\text{O}]$
C. rate = $k[\text{F}_2][\text{Cl}_2\text{O}]$
D. rate = $k[\text{F}_2][\text{Cl}_2\text{O}]^2$
E. rate = $k[\text{F}_2]^2[\text{Cl}_2\text{O}]^2$
20. Nitric oxide reacts with chlorine to form nitrosyl chloride, NOCl. Use the following data to determine the rate equation for the reaction.



Expt. #	[NO]	[Cl ₂]	Initial Rate
1	0.22	0.065	0.96 M/min
2	0.66	0.065	8.6 M/min
3	0.44	0.032	1.9 M/min

- A. rate = $k[\text{NO}]$
B. rate = $k[\text{NO}][\text{Cl}_2]^{1/2}$
C. rate = $k[\text{NO}][\text{Cl}_2]$
D. rate = $k[\text{NO}]^2[\text{Cl}_2]$
E. rate = $k[\text{NO}]^2[\text{Cl}_2]^2$
21. A first-order reaction has a rate constant of $3.00 \times 10^{-3} \text{ s}^{-1}$. The time required for the reaction to be 75.0% complete is
- A. 95.8 s.
B. 462 s.
C. 231 s.
D. 201 s.
E. 41.7 s.

22. A certain first-order reaction $A \rightarrow B$ is 25% complete in 42 min at 25°C. What is its rate constant?

- A. $6.8 \times 10^{-3} \text{ min}^{-1}$
- B. $8.3 \times 10^{-3} \text{ min}^{-1}$
- C. $3.3 \times 10^{-2} \text{ min}^{-1}$
- D. $-3.3 \times 10^{-2} \text{ min}^{-1}$
- E. 11 min^{-1}

23. Ammonium ion (NH_4^+) reacts with nitrite ion (NO_2^-) to yield nitrogen gas and liquid water. The following initial rates of reaction have been measured for the given reactant concentrations.

Expt. #	$[\text{NH}_4^+]$	$[\text{NO}_2^-]$	Initial rate (M/hr)
1	0.010	0.020	0.020
2	0.015	0.020	0.030
3	0.030	0.010	0.015

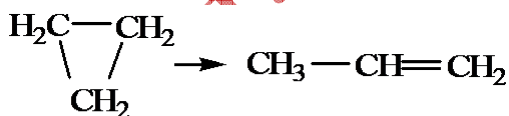
Which of the following is the rate law (rate equation) for this reaction?

- A. $\text{rate} = k [\text{NH}_4^+][\text{NO}_2^-]^4$
- B. $\text{rate} = k [\text{NH}_4^+][\text{NO}_2^-]$
- C. $\text{rate} = k [\text{NH}_4^+][\text{NO}_2^-]^2$
- D. $\text{rate} = k [\text{NH}_4^+]^2[\text{NO}_2^-]$
- E. $\text{rate} = k [\text{NH}_4^+]^{1/2}[\text{NO}_2^-]^{1/4}$

24. Appropriate units for a second-order rate constant are

- A. M/s.
- B. $1/\text{M}\cdot\text{s}$.
- C. 1/s.
- D. $1/\text{M}^2\cdot\text{s}$.

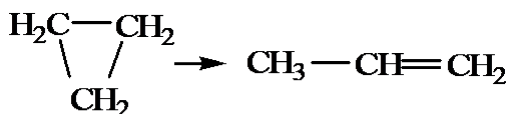
25. The isomerization of cyclopropane to form propene



is a first-order reaction. At 760 K, 15% of a sample of cyclopropane changes to propene in 6.8 min. What is the half-life of cyclopropane at 760 K?

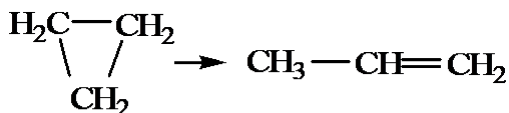
- A. $3.4 \times 10^{-2} \text{ min}$
- B. 2.5 min
- C. 23 min
- D. 29 min
- E. 230 min

26. The isomerization of cyclopropane to form propene



is a first-order reaction. At 760 K, 85% of a sample of cyclopropane changes to propene in 79.0 min. Determine the rate constant for this reaction at 760 K.

- A. $3.66 \times 10^{-2} \text{ min}^{-1}$
 B. $1.04 \times 10^{-2} \text{ min}^{-1}$
 C. 2.42 min^{-1}
 D. $2.06 \times 10^{-3} \text{ min}^{-1}$
 E. $2.40 \times 10^{-2} \text{ min}^{-1}$
27. The isomerization of cyclopropane to propene follows first-order kinetics. At 700 K, the rate constant for this reaction is $6.2 \times 10^{-4} \text{ min}^{-1}$. How many minutes are required for 10.0% of a sample of cyclopropane to isomerize to propene?



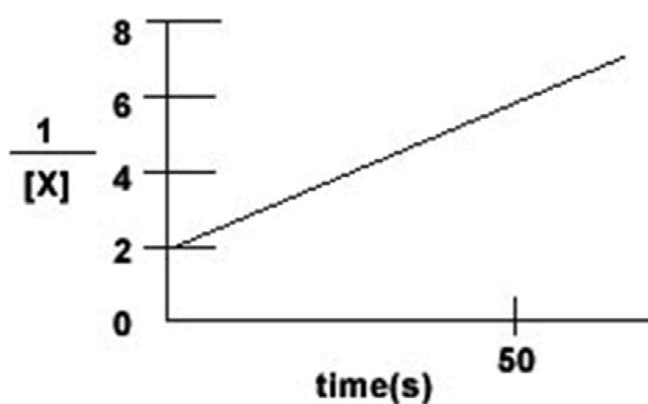
- A. 16,100 min
 B. 170 min
 C. 3,710 min
 D. $1.43 \times 10^{-3} \text{ min}$
 E. 1,120 min
28. At 700 K, the rate constant for the following reaction is $6.2 \times 10^{-4} \text{ min}^{-1}$. How many minutes are required for 20% of a sample of cyclopropane to isomerize to propene?



- A. 1,120 min
 B. 360 min
 C. 3710 min
 D. $1.4 \times 10^{-4} \text{ min}$
 E. 280 min

29. A first-order reaction has a rate constant of 7.5×10^{-3} /s. The time required for the reaction to be 60% complete is
- A. 3.8×10^{-3} s.
 - B. 6.9×10^{-3} s.
 - C. 68 s.
 - D. 120 s.
 - E. 130 s.
30. The first-order reaction $\text{SO}_2\text{Cl}_2 \rightarrow \text{SO}_2 + \text{Cl}_2$ is 10% complete in 80. min. How long would it take for the reaction to be 95% complete?
- A. 1.8 min
 - B. 104 min
 - C. 530 min
 - D. 2300 min
 - E. 990 min
31. Benzoyl chloride, $\text{C}_6\text{H}_5\text{COCl}$, reacts with water to form benzoic acid, $\text{C}_6\text{H}_5\text{COOH}$, and hydrochloric acid. This first-order reaction is 25% complete after 26 s. How much longer would one have to wait in order to obtain 99% conversion of benzoyl chloride to benzoic acid?
- A. 390 s
 - B. 420 s
 - C. 180 s
 - D. 290 s
 - E. 210 s
32. A certain reaction $\text{A} \rightarrow \text{products}$ is second order in A. If this reaction is 10.% complete after 20. s, how long would it take for the reaction to be 90.% complete?
- A. 180 s
 - B. 1600 s
 - C. 440 s
 - D. 18,000 s
 - E. 540 s
33. A certain reaction $\text{A} \rightarrow \text{products}$ is second order in A. If this reaction is 85% complete in 12 minutes, how long would it take for the reaction to be 15% complete?
- A. 110 s
 - B. 27 s
 - C. 62 s
 - D. 130 s
 - E. 22 s

34. At 25°C, the second-order reaction $\text{NOCl}(\text{g}) \rightarrow \text{NO}(\text{g}) + 1/2\text{Cl}_2(\text{g})$ is 50% complete after 5.82 hours when the initial concentration of NOCl is 4.46 mol/L. How long will it take for the reaction to be 75% complete?
- A. 8.22 hr
B. 11.6 hr
C. 15.5 hr
D. 17.5 hr
E. 23.0 hr
35. For the reaction $\text{X} + \text{Y} \rightarrow \text{Z}$, the reaction rate is found to depend only upon the concentration of X. A plot of $1/\text{X}$ versus time gives a straight line.



What is the rate law for this reaction?

- A. rate = $k[\text{X}]$
B. rate = $k[\text{X}]^2$
C. rate = $k[\text{X}][\text{Y}]$
D. rate = $k[\text{X}]^2[\text{Y}]$
36. The reaction $2\text{NO}_2(\text{g}) \rightarrow 2\text{NO}(\text{g}) + \text{O}_2(\text{g})$ is suspected to be second order in NO_2 . Which of the following kinetic plots would be the most useful to confirm whether or not the reaction is second order?
- A. a plot of $[\text{NO}_2]^{-1}$ vs. t
B. a plot of $\ln[\text{NO}_2]$ vs. t
C. a plot of $[\text{NO}_2]$ vs. t
D. a plot of $\ln[\text{NO}_2]^{-1}$ vs. t
E. a plot of $[\text{NO}_2]^2$ vs. t

37. The thermal decomposition of acetaldehyde, $\text{CH}_3\text{CHO} \rightarrow \text{CH}_4 + \text{CO}$, is a second-order reaction. The following data were obtained at 518°C .

time, s	Pressure CH_3CHO , mmHg
0	364
42	330
105	290
720	132

Calculate the rate constant for the decomposition of acetaldehyde from the above data.

- A. $2.2 \times 10^{-3}/\text{s}$
B. 0.70 mmHg/s
C. $2.2 \times 10^{-3}/\text{mmHg}\cdot\text{s}$
D. $6.7 \times 10^{-6}/\text{mmHg}\cdot\text{s}$
E. $5.2 \times 10^{-5}/\text{mmHg}\cdot\text{s}$
38. The thermal decomposition of acetaldehyde, $\text{CH}_3\text{CHO} \rightarrow \text{CH}_4 + \text{CO}$, is a second-order reaction. The following data were obtained at 518°C .

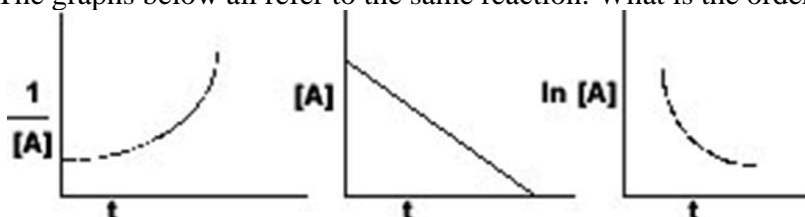
time, s	Pressure CH_3CHO , mmHg
0	364
42	330
105	290
720	132

Based on the data given, what is the half-life for the disappearance of acetaldehyde?

- A. $1.5 \times 10^5 \text{ s}$
B. 410 s
C. $5.4 \times 10^7 \text{ s}$
D. 520 s
E. 305 s
39. For the chemical reaction $\text{A} \rightarrow \text{B} + \text{C}$, a plot of $[\text{A}]_t$ versus time is found to give a straight line with a negative slope. What is the order of reaction with respect to A?
- A. zeroth
B. first
C. second
D. third
E. Such a plot cannot reveal the order of the reaction.

40. For the chemical reaction $A \rightarrow C$, a plot of $1/[A]_t$ versus time was found to give a straight line with a positive slope. What is the order of reaction?
- zeroth
 - first
 - second
 - Such a plot cannot reveal the order of the reaction.

41. The graphs below all refer to the same reaction. What is the order of this reaction?

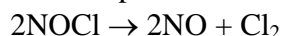


- zeroth order
 - first order
 - second order
 - unable to predict
42. For what order reaction does the half-life get longer as the initial concentration increases?
- zeroth order
 - first order
 - second order
 - none of them because half-life is always independent of the initial concentration
43. For a second order reaction, the half-life is equal to
- $t_{1/2} = 0.693/k$.
 - $t_{1/2} = k/0.693$.
 - $t_{1/2} = 1/k[A]_0$.
 - $t_{1/2} = k$.
 - $t_{1/2} = [A]_0/2k$.
44. Which one of the following changes would alter the rate constant (k) for the reaction $2A + B \rightarrow \text{products}$?
- increasing the concentration of A
 - increasing the concentration of B
 - increasing the temperature
 - measuring k again after the reaction has run for a while

45. The Arrhenius equation is $k = A e^{-(E_a/RT)}$. The slope of a plot of $\ln k$ vs. $1/T$ is equal to

- A. $-k$.
- B. k .
- C. E_a .
- D. $-E_a/R$.
- E. A .

46. What is the slope of an Arrhenius plot for the following reaction?



Temperature (K)	k (L/mol·s)
400.	6.6×10^{-4}
500.	2.9×10^{-1}
600.	16.3

- A. 8.18×10^{-2} K
- B. 5.06×10^{-2} K
- C. -1.22×10^4 K
- D. -1.96×10^4 K
- E. not enough information to calculate the slope

47. The activation energy for the reaction $\text{CH}_3\text{CO} \rightarrow \text{CH}_3 + \text{CO}$ is 71 kJ/mol. How many times greater is the rate constant for this reaction at 170°C than at 150°C?

- A. 0.40
- B. 1.1
- C. 2.5
- D. 4.0
- E. 5.0

48. If E_a for a certain biological reaction is 50. kJ/mol, by what factor (how many times will the rate of this reaction increase when body temperature increases from 37°C (normal) to 40°C (fever)?

- A. 1.15
- B. 1.20
- C. 2.0×10^5
- D. 1.0002
- E. 2.0

49. At 25°C, by what factor is the reaction rate increased by a catalyst that reduces the activation energy of the reaction by 1.00 kJ/mol?
- A. 1.63
B. 123
C. 1.04
D. 1.50
E. 2.53
50. At 30°C, by how much is a reaction's activation energy decreased by the addition of a catalyst if the catalyst triples the reaction rate?
- A. 2.77 kJ/mol
B. 274 J/mol
C. 2.70 J/mol
D. 119 J/mol
E. 1.20 kJ/mol
51. The activation energy for the following reaction is 60. kJ/mol.
 $\text{Sn}^{2+} + 2\text{Co}^{3+} \rightarrow \text{Sn}^{4+} + 2\text{Co}^{2+}$
- By what factor (how many times) will the rate constant increase when the temperature is raised from 10°C to 28°C?
- A. 1.002
B. 4.6
C. 5.6
D. 2.8
E. 696
52. The isomerization of cyclopropane follows first order kinetics. The rate constant at 700 K is $6.20 \times 10^{-4} \text{ min}^{-1}$, and the half-life at 760 K is 29.0 min. Calculate the activation energy for this reaction.
- A. 5.07 kJ/mol
B. 27.0 kJ/mol
C. 50.7 kJ/mol
D. 160. kJ/mol
E. 270. kJ/mol

53. The isomerization of methyl isocyanide, $\text{CH}_3\text{NC} \rightarrow \text{CH}_3\text{CN}$, follows first-order kinetics. The half-lives were found to be 161 min at 199°C and 12.5 min at 230°C . Calculate the activation energy for this reaction.

- A. 6.17×10^{-3} kJ/mol
- B. 31.4 kJ/mol
- C. 78.2 kJ/mol
- D. 124 kJ/mol
- E. 163 kJ/mol

54. Calculate the activation energy, in kJ/mol, for the redox reaction $\text{Sn}^{2+} + 2\text{Co}^{3+} \rightarrow \text{Sn}^{4+} + 2\text{Co}^{2+}$.

Data: Temp ($^\circ\text{C}$)	k (1/M·s)
2	3.12×10^3
27	27.0×10^3

- A. 59.2
- B. 0.477
- C. 5.37
- D. 163 kJ
- E. 48.1 kJ

55. The reaction $\text{C}_4\text{H}_{10} \rightarrow \text{C}_2\text{H}_6 + \text{C}_2\text{H}_4$ has an activation energy (E_a) of 350 kJ/mol, and the E_a of the reverse reaction is 260 kJ/mol. Estimate ΔH , in kJ/mol, for the reaction as written above.

- A. -90 kJ/mol
- B. +90 kJ/mol
- C. 350 kJ/mol
- D. -610 kJ/mol
- E. +610 kJ/mol

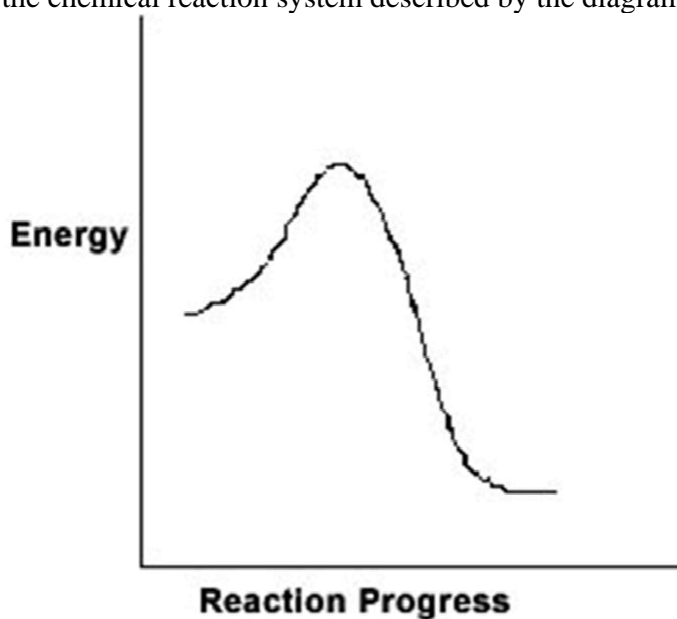
56. The activation energy for the following first-order reaction is 102 kJ/mol.



The value of the rate constant (k) is $1.35 \times 10^{-4} \text{ s}^{-1}$ at 35°C . What is the value of k at 0°C ?

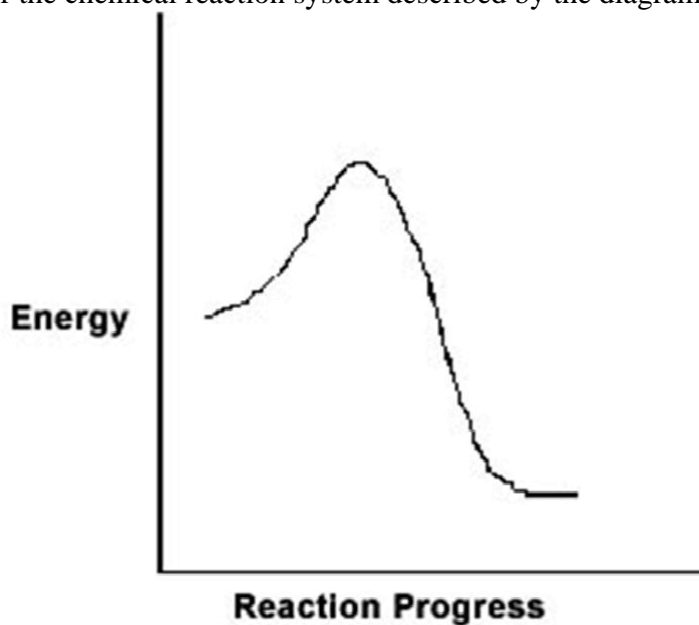
- A. $8.2 \times 10^{-7} \text{ s}^{-1}$
- B. $1.9 \times 10^{-5} \text{ s}^{-1}$
- C. $4.2 \times 10^{-5} \text{ s}^{-1}$
- D. $2.2 \times 10^{-2} \text{ s}^{-1}$
- E. none of these

57. Given that E_a for a certain biological reaction is 48 kJ/mol and that the rate constant is $2.5 \times 10^{-2} \text{ s}^{-1}$ at 15°C , what is the rate constant at 37°C ?
- A. $2.7 \times 10^{-2} \text{ s}^{-1}$
B. $2.5 \times 10^{-1} \text{ s}^{-1}$
C. $1.0 \times 10^{-1} \text{ s}^{-1}$
D. $6.0 \times 10^{-3} \text{ s}^{-1}$
E. 1.1 s^{-1}
58. The activation energy for the reaction $\text{O} + \text{O}_3 \rightarrow 2\text{O}_2$ is 25 kJ/mol, and the enthalpy change is $\Delta H = -388$ kJ/mol. What is the activation energy for the decomposition of O_2 by the reverse reaction?
- A. 413 kJ
B. 388 kJ
C. 363 kJ
D. 50 kJ
E. 25 kJ
59. For the chemical reaction system described by the diagram below, which statement is *true*?



- A. The forward reaction is endothermic.
B. The activation energy for the forward reaction is greater than the activation energy for the reverse reaction.
C. At equilibrium, the activation energy for the forward reaction is equal to the activation energy for the reverse reaction.
D. The activation energy for the reverse reaction is greater than the activation energy for the forward reaction.
E. The reverse reaction is exothermic.

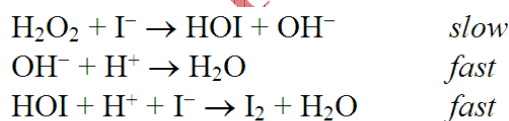
60. An increase in the temperature of the reactants causes an increase in the rate of reaction. The best explanation for this behavior is that as the temperature increases,
- A. the concentration of reactants increases.
 - B. the activation energy decreases.
 - C. the collision frequency increases.
 - D. the fraction of collisions with total kinetic energy $> E_a$ increases.
 - E. the activation energy increases.
61. For the chemical reaction system described by the diagram below, which statement is *true*?



If the E_a for the forward reaction is 25 kJ/mol and the enthalpy of reaction is -95 kJ/mol, what is E_a for the reverse reaction?

- A. 120 kJ/mol
- B. 70 kJ/mol
- C. 95 kJ/mol
- D. 25 kJ/mol
- E. -70 kJ/mol

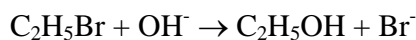
62. According to the collision theory, all collisions do not lead to reaction. Which choice gives *both* reasons why not all collisions between reactant molecules lead to reaction?
1. The total energy of two colliding molecules is less than some minimum amount of energy.
 2. Molecules cannot react with each other unless a catalyst is present.
 3. Molecules that are improperly oriented during collision will not react.
 4. Solids cannot react with gases.
- A. 1 and 2
B. 1 and 3
C. 1 and 4
D. 2 and 3
E. 3 and 4
63. When the concentrations of reactant molecules are increased, the rate of reaction increases. The best explanation for this phenomenon is that as the reactant concentration increases,
- A. the average kinetic energy of molecules increases.
 - B. the frequency of molecular collisions increases.
 - C. the rate constant increases.
 - D. the activation energy increases.
 - E. the order of reaction increases.
64. A reaction mechanism usually is
- A. the same as the balanced chemical equation.
 - B. restricted to only one possible explanation.
 - C. obvious if the reaction order is known.
 - D. difficult, if not impossible, to prove.
 - E. obvious if the activation energy is known.
65. The rate law for the reaction $\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{I}^- \rightarrow \text{I}_2 + 2\text{H}_2\text{O}$ is $\text{rate} = k[\text{H}_2\text{O}_2][\text{I}^-]$. The following mechanism has been suggested.



Identify all intermediates included in this mechanism.

- A. H^+ and I^-
- B. H^+ and HOI
- C. HOI and OH^-
- D. H^+ only
- E. H_2O and OH^-

66. The following reaction in aqueous solution was found to be first order in $[\text{OH}^-]$, first order in $[\text{C}_2\text{H}_5\text{Br}]$, and inverse first order in $[\text{Br}^-]$.



Which one of the following mechanisms is consistent with the observed reaction order?

- A. $\text{C}_2\text{H}_5\text{Br} \rightleftharpoons \text{C}_2\text{H}_5^+ + \text{Br}^-$ *fast*
 $\text{C}_2\text{H}_5^+ + \text{OH}^- \rightarrow \text{C}_2\text{H}_5\text{OH}$ *slow*
- B. $\text{C}_2\text{H}_5\text{Br} + \text{H}_2\text{O} \rightarrow \text{C}_2\text{H}_5\text{OH} + \text{H}^+ + \text{Br}^-$ *slow*
 $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$ *fast*
- C. $\text{C}_2\text{H}_5\text{Br} \rightleftharpoons \text{C}_2\text{H}_5^+ + \text{Br}^-$ *slow*
 $\text{C}_2\text{H}_5^+ + \text{OH}^- \rightarrow \text{C}_2\text{H}_5\text{OH}$ *fast*
- D. $\text{C}_2\text{H}_5\text{Br} \rightleftharpoons \text{C}_2\text{H}_5^+ + \text{Br}^-$ *slow*
 $\text{OH}^- + \text{Br}^- \rightarrow \text{HOBr}$ *fast*
 $\text{HOBr} + \text{C}_2\text{H}_5^+ \rightarrow \text{C}_2\text{H}_5\text{OH} + \text{Br}^-$ *fast*

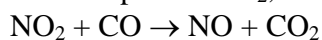
67. The rate law for the reaction $2\text{NO}_2 + \text{O}_3 \rightarrow \text{N}_2\text{O}_5 + \text{O}_2$ is rate = $k[\text{NO}_2][\text{O}_3]$. Which one of the following mechanisms is consistent with this rate law?

- A. $\text{NO}_2 + \text{NO}_2 \rightarrow \text{N}_2\text{O}_4$ *fast*
 $\text{N}_2\text{O}_4 + \text{O}_3 \rightarrow \text{N}_2\text{O}_5 + \text{O}_2$ *slow*
- B. $\text{NO}_2 + \text{O}_3 \rightarrow \text{NO}_5$ *fast*
 $\text{NO}_5 + \text{NO}_2 \rightarrow \text{N}_2\text{O}_5 + (5/2)\text{O}_2$ *slow*
- C. $\text{NO}_2 + \text{O}_3 \rightarrow \text{NO}_3 + \text{O}_2$ *slow*
 $\text{NO}_3 + \text{NO}_2 \rightarrow \text{N}_2\text{O}_5$ *fast*
- D. $\text{NO}_2 + \text{NO}_2 \rightarrow \text{N}_2\text{O}_2 + \text{O}_2$ *slow*
 $\text{N}_2\text{O}_2 + \text{O}_3 \rightarrow \text{N}_2\text{O}_5$ *fast*

68. For the reaction $\text{X}_2 + \text{Y} + \text{Z} \rightarrow \text{XY} + \text{XZ}$, it is found that the rate equation is rate = $k[\text{X}_2][\text{Y}]$. Why does the concentration of Z have no effect on the rate?

- A. The concentration of Z is very small and the others are very large.
 B. Z must react in a step after the rate determining step.
 C. Z is an intermediate.
 D. The fraction of molecules of Z that have very high energies is zero.
 E. The activation energy for Z to react is very high.

69. The gas phase reaction of nitrogen dioxide and carbon monoxide was found by experiment to be second-order with respect to NO_2 , and zeroth-order with respect to CO below 25°C .



Which one of the following mechanisms is consistent with the observed reaction order?

- A. $\text{NO}_2 + 2\text{CO} \rightleftharpoons \text{N} + 2\text{CO}_2$ *fast*
 $\text{N} + \text{NO}_2 \rightarrow 2\text{NO}$ *slow*
- B. $\text{NO}_2 + 2\text{CO} \rightarrow \text{N} + 2\text{CO}_2$ *slow*
 $\text{N} + \text{NO}_2 \rightarrow 2\text{NO}$ *fast*
- C. $\text{NO}_2 + \text{NO}_2 \rightarrow \text{NO}_3 + \text{NO}$ *fast*
 $\text{NO}_3 + \text{CO} \rightarrow \text{NO}_2 + \text{CO}_2$ *slow*
- D. $\text{NO}_2 + \text{NO}_2 \rightarrow \text{NO}_3 + \text{NO}$ *slow*
 $\text{NO}_3 + \text{CO} \rightarrow \text{NO}_2 + \text{CO}_2$ *fast*

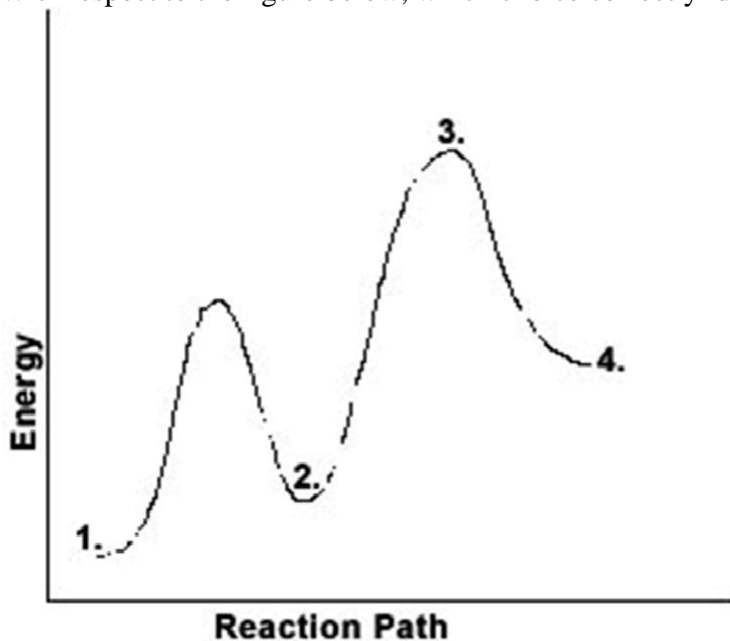
70. Which of the following statements is *false*?

- A. A catalyst increases the rate of the forward reaction, but does not alter the reverse rate.
B. A catalyst alters the mechanism of reaction.
C. A catalyst alters the activation energy.
D. A catalyst may be altered in the reaction, but is always regenerated.
E. A catalyst increases the rate of reaction, but is not consumed.

71. Complete the following statement: A catalyst

- A. increases the activation energy.
B. alters the reaction mechanism.
C. increases the average kinetic energy of the reactants.
D. increases the concentration of reactants.
E. increases the collision frequency of reactant molecules.

72. With respect to the figure below, which choice correctly identifies all the numbered positions?



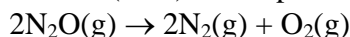
	1.	2.	3.	4.
Row 1	catalyst	intermediate	activated complex	product
Row 2	reactants	activated complex	intermediate	product
Row 3	reactants	activated complex	catalyst	product
Row 4	reactants	intermediate	activated complex	product
Row 5	reactants	intermediate	activated complex	catalyst

- A. Row 1
 B. Row 2
 C. Row 3
 D. Row 4
 E. Row 5

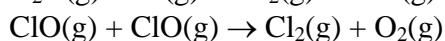
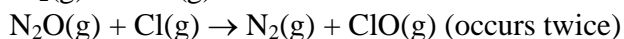
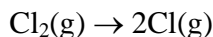
73. The activation energy of a certain uncatalyzed reaction is 64 kJ/mol. In the presence of a catalyst, the E_a is 55 kJ/mol. How many times faster is the catalyzed than the uncatalyzed reaction at 400°C? Assume that the frequency factor remains the same.

- A. 5.0 times
 B. 1.16 times
 C. 15 times
 D. 2.0 times
 E. 0.2 times

74. Nitrous oxide (N_2O) decomposes at 600°C according to the balanced equation

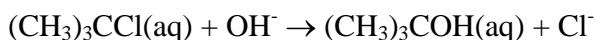


A reaction mechanism involving three steps is shown below. Identify all of the catalysts in the following mechanism.



- A. Cl
B. Cl_2
C. ClO
D. N_2O
E. ClO and Cl
75. Peroxodisulfate ion can oxidize iodide ions to iodine according to the balanced equation
- $$\text{S}_2\text{O}_8^{2-} + 2\text{I}^- \rightarrow 2\text{SO}_4^{2-} + \text{I}_2$$
- The reaction is catalyzed by certain chemical species. Identify the catalyst in the following mechanism:
- step 1: $\text{Fe}^{3+} + 2\text{I}^- \rightarrow \text{Fe}^{2+} + \text{I}_2$
step 2: $\text{S}_2\text{O}_8^{2-} + \text{Fe}^{2+} \rightarrow 2\text{SO}_4^{2-} + \text{Fe}^{3+}$
- A. Fe^{3+}
B. I^-
C. $\text{S}_2\text{O}_8^{2-}$
D. Fe^{2+}
E. SO_4^{2-}
76. In which of the forms listed below would 0.5 g aluminum react the fastest with gaseous chlorine at 25°C ?
- A. 0.5 g aluminum in one piece
B. 0.5 g aluminum divided into 10 pieces
C. 0.5 g aluminum divided into 100 pieces
D. 0.5 g aluminum divided into 1,000 pieces
E. All the choices will react at the same rate since the temperature is the same.
77. For the reaction whose rate law is $\text{rate} = k[\text{X}]$, a plot of which of the following is a straight line?
- A. $[\text{X}]$ versus time
B. $\log [\text{X}]$ versus time
C. $1/[\text{X}]$ versus time
D. $[\text{X}]$ versus $1/\text{time}$
E. $\log [\text{X}]$ versus $1/\text{time}$

78. For the reaction represented below, the experimental rate law is given by $\text{rate} = k [(\text{CH}_3)_3\text{CCl}]$.



If some solid sodium hydroxide were added to a solution in which $[(\text{CH}_3)_3\text{CCl}] = 0.01 \text{ M}$ and $[\text{NaOH}] = 0.10 \text{ M}$, which of the following would be *true*? (Assume the temperature and volume remain constant.)

- A. Both the reaction rate and k would increase.
B. Both the reaction rate and k would decrease.
C. Both the reaction rate and k would remain the same.
D. The reaction rate would increase but k would remain the same.
E. The reaction rate would decrease but k would remain the same.
79. At a particular temperature the first-order gas-phase reaction $2\text{N}_2\text{O}_5 \rightarrow 2\text{N}_2\text{O}_4 + \text{O}_2$ has a half-life for the disappearance of dinitrogen pentoxide of 3240 s. If 1.00 atm of N_2O_5 is introduced into an evacuated 5.00 L flask, what will be the total pressure of the gases in the flask after 1.50 hours?
- A. 0.685 atm
B. 1.00 atm
C. 0.315 atm
D. 1.68 atm
E. 1.34 atm
80. At a particular temperature the first-order gas-phase reaction $\text{N}_2\text{O}_5 \rightarrow 2\text{NO}_2 + (1/2)\text{O}_2$ has a half-life for the disappearance of dinitrogen pentoxide of 5130 s. Suppose 0.450 atm of N_2O_5 is introduced into an evacuated 2.00 L flask. What will be the total gas pressure inside the flask after 3.00 hours?
- A. 0.969 atm
B. 0.105 atm
C. 0.795 atm
D. 1.14 atm
E. 0.864 atm
81. When acetaldehyde at a pressure of 364 mmHg is introduced into an evacuated 500. mL flask at 518°C , the half-life for the second-order decomposition process, $\text{CH}_3\text{CHO} \rightarrow \text{CH}_4 + \text{CO}$, is 410. s. What will the total pressure in the flask be after 1.00 hour?
- A. 327 mmHg
B. 654 mmHg
C. 37 mmHg
D. 691 mmHg
E. 728 mmHg

82. The first-order decomposition of SO_2Cl_2 to sulfur dioxide gas and chlorine gas at 320°C has a half-life of 8.75 hr. If one begins with 600. mmHg of pure SO_2Cl_2 in a 5.00-L flask at 320°C , how long does it take for the total pressure in the flask to rise to 1.000 atm?
- A. 6.45 hr
B. 11.1 hr
C. 3.91 hr
D. 20.3 hr
E. 6.91 hr
83. The first-order decomposition of phosphene to phosphorus and hydrogen
 $4\text{PH}_3(\text{g}) \rightarrow \text{P}_4(\text{g}) + 6\text{H}_2(\text{g})$
has a half-life of 35.0 s at 680°C . Starting with 520 mmHg of pure phosphene in an 8.00-L flask at 680° , how long will it take for the total pressure in the flask to rise to 1.000 atm?
- A. 628 s
B. 33.4 s
C. 51.2 s
D. 111 s
E. 48.3 s
84. Use the table of data shown below to calculate the average rate of the reaction $\text{A} \rightarrow \text{B}$ between 10 s and 20 s.

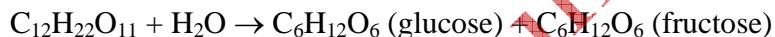
time (s)	[A] mol/L
0	0.20
5	0.14
10	0.10
15	0.071
20	0.050

85. What is the rate law that corresponds to the data shown for the reaction $2A + B \rightarrow C$?

Exp.	Initial [A]	Initial [B]	Initial rate
1	0.015	0.022	0.125
2	0.030	0.044	0.500
3	0.060	0.044	0.500
4	0.060	0.066	1.125

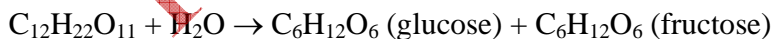
86. The rate constant for a certain first-order reaction is 0.40/min. What is the initial rate in mole/L-min, if the initial concentration of the compound involved is 0.50 mol/L?

87. Sucrose, $C_{12}H_{22}O_{11}$, reacts slowly with water in the presence of an acid to form two other sugars, glucose and fructose, both of which have the same molecular formulas, but different structures.



The reaction is first order and has a rate constant of $6.2 \times 10^{-5}/s$ at $35^\circ C$ when the H^+ concentration is 0.10 M. Suppose that the initial concentration of sucrose in the solution is 0.40 M. What will the sucrose concentration be after 2.0 hours?

88. Sucrose, $C_{12}H_{22}O_{11}$, reacts slowly with water in the presence of an acid to form two other sugars, glucose and fructose, both of which have the same molecular formulas, but different structures.



The reaction is first order and has a rate constant of $6.2 \times 10^{-5}/s$ at $35^\circ C$ when the H^+ concentration is 0.10 M. Suppose that the initial concentration of sucrose in the solution is 0.40 M. How many minutes will it take for the sucrose concentration to drop to 0.30 M?

89. At a certain temperature, the data below were collected for the reaction $2\text{ICl} + \text{H}_2 \rightarrow \text{I}_2 + 2\text{HCl}$. Determine the rate law for the reaction.

Initial concentrations (M)		Initial Rate of Formation of I_2
$[\text{ICl}]$	$[\text{H}_2]$	mol/L·s
0.10	0.10	0.0015
0.20	0.10	0.0030
0.10	0.050	0.00075

90. At a certain temperature, the data below were collected for the reaction $2\text{ICl} + \text{H}_2 \rightarrow \text{I}_2 + 2\text{HCl}$. Determine the rate constant for the reaction.

Initial concentrations (M)		Initial Rate of Formation of I_2
$[\text{ICl}]$	$[\text{H}_2]$	mol/L·s
0.10	0.10	0.0015
0.20	0.10	0.0030
0.10	0.050	0.00075

91. Nitrogen pentoxide decomposes by a first-order process yielding N_2O_4 and oxygen.
 $2\text{N}_2\text{O}_5 \rightarrow 2\text{N}_2\text{O}_4 + \text{O}_2$

At a given temperature, the half-life of N_2O_5 is 0.90 hr. What is the first-order rate constant for N_2O_5 decomposition?

92. For the first-order reaction $2\text{N}_2\text{O}_5 \rightarrow 2\text{N}_2\text{O}_4 + \text{O}_2$ at a particular temperature, the half-life of N_2O_5 is 0.90 hr. What fraction of the initial concentration of N_2O_5 will remain after 2.4 hours?
93. The rate constant for the first-order decomposition of C_4H_8 at 500°C is $9.2 \times 10^{-3} \text{ s}^{-1}$. How long will it take for 10.0% of a 0.100 M sample of C_4H_8 to decompose at 500°C ?
94. The activation energy for a certain reaction is 113 kJ/mol. By what factor (how many times) will the rate constant increase when the temperature is raised from 310 K to 325 K?
95. Nitric acid is formed by the gas-phase hydrolysis of N_2O_5 . Draw an energy profile curve for the reaction $\text{N}_2\text{O}_5 + \text{H}_2\text{O} \rightarrow 2\text{HNO}_3$ given that $E_a(\text{forward}) = 15 \text{ kJ}$ and $E_a(\text{reverse}) = 51 \text{ kJ/mol}$. Label the activation energies of the forward and reverse reactions and label ΔH for the reaction.
96. Nitric acid is formed by the gas-phase hydrolysis of N_2O_5 . For the reaction $\text{N}_2\text{O}_5 + \text{H}_2\text{O} \rightarrow 2\text{HNO}_3$, $E_a(\text{forward}) = 15 \text{ kJ/mol}$ and $E_a(\text{reverse}) = 51 \text{ kJ/mol}$. Calculate ΔH_{rxn} .

97. The oxidation of iodide ions by arsenic acid in acidic aqueous solution occurs according to the net reaction $\text{H}_3\text{AsO}_4 + 3\text{I}^- + 2\text{H}_3\text{O}^+ \rightarrow \text{H}_3\text{AsO}_3 + \text{I}_3^- + \text{H}_2\text{O}$. The experimental rate law for this reaction is $\text{Rate} = k[\text{H}_3\text{AsO}_4][\text{I}^-][\text{H}_3\text{O}^+]$.

What is the order of the reaction with respect to I^- ?

98. The oxidation of iodide ions by arsenic acid in acidic aqueous solution occurs according to the net reaction $\text{H}_3\text{AsO}_4 + 3\text{I}^- + 2\text{H}_3\text{O}^+ \rightarrow \text{H}_3\text{AsO}_3 + \text{I}_3^- + \text{H}_2\text{O}$. The experimental rate law for this reaction is $\text{Rate} = k[\text{H}_3\text{AsO}_4][\text{I}^-][\text{H}_3\text{O}^+]$.

According to the rate law for the reaction, an increase in the concentration of hydronium ion has what effect on this reaction?

99. For the following exothermic reaction, the rate law at 298 K is $\text{Rate} = k [\text{H}_2] [\text{I}_2]$:
 $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightarrow 2\text{HI}(\text{g})$

Predict the effect of each of the following changes on the initial rate of the reaction.

- Addition of hydrogen gas at constant temperature and volume
- Increase in volume of the reaction vessel at constant temperature
- Addition of a catalyst
- Increase in temperature

100. In the reaction, $2\text{N}_2\text{O} \rightarrow 2\text{N}_2 + \text{O}_2$, oxygen and nitrogen gases are formed at the same rate (mol/L·s) .

True False

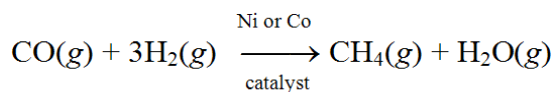
101. The rate constant of a first-order reaction, $\text{A} \rightarrow \text{products}$, can be determined from a graph of $\ln[\text{A}]$ versus t .

True False

102. For the first-order reaction, $A \rightarrow \text{products}$, if half of the initial concentration of A reacts in 20 min, then the remaining half will completely react in the next 20 min.

True False

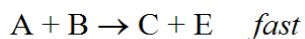
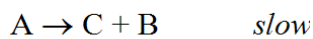
103. Substitute natural gas can be synthesized by passing carbon monoxide and hydrogen over Ni or Co at 400°C.



This process is an example of homogeneous catalysis.

True False

104. The rate law predicted by the following two-step mechanism is $\text{Rate} = k[\text{A}][\text{B}]$.



True False

Prof. Abdullah M. Asiri

Chapter 13 Chemical Kinetics **Key**

1.E

2.B

3.C

4.B

5.C

6.E

7.C

8.D

9.D

10.D

11.C

12.C

13.E

14.D

15.B

16.E

17.D

18.E

19.C

20.D

21.B

22.A

23.C

24.B

25.D

26.E

27.B

28.B

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- 29.D
- 30.D
- 31.A
- 32.B
- 33.E
- 34.D
- 35.B
- 36.A
- 37.D
- 38.B
- 39.A
- 40.C
- 41.A
- 42.A
- 43.C
- 44.C
- 45.D
- 46.C
- 47.C
- 48.B
- 49.D
- 50.A
- 51.B
- 52.E
- 53.E
- 54.A
- 55.B
- 56.A
- 57.C
- 58.A
- 59.D
- 60.D
- 61.A

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62.B

63.B

64.D

65.C

66.A

67.C

68.B

69.D

70.A

71.B

72.D

73.A

74.B

75.A

76.D

77.B

78.C

79.E

80.A

81.D

82.C

83.E

84. 5×10^{-3} M/s85. Rate = $k[B]^2$

86. 0.20

87. 0.26 M

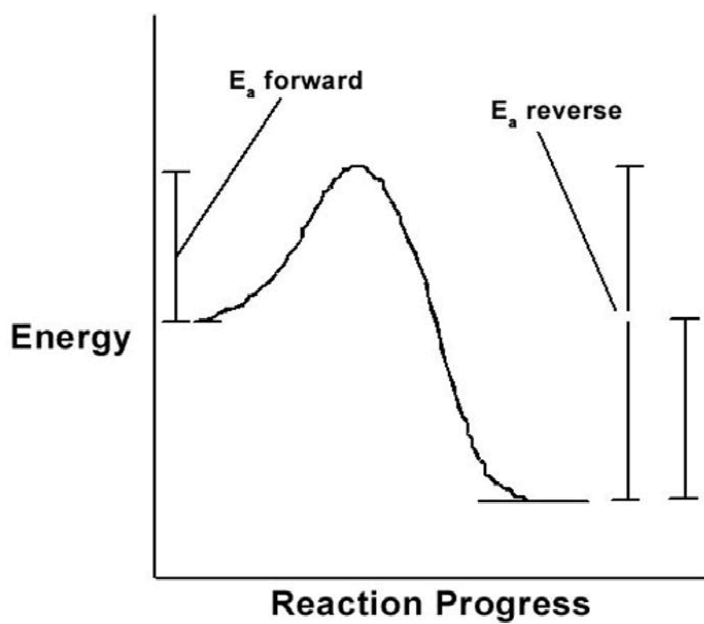
88. 77 min

89. rate = $k[ICl][H_2]$ 90. 1.5×10^{-1} L/mol·s91. 0.77 hr^{-1}

92. 0.16

93. 11 s

94. 7.6



95.

96. -36 kJ/mol

97. first

98. The rate of reaction increases.

99. a. The initial rate increases.

b. The initial rate decreases.

c. The initial rate increases

d. The initial rate increases.

100. FALSE

101. TRUE

102. FALSE

103. FALSE

104. FALSE

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