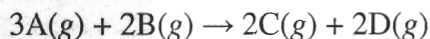


The people sitting on either side of you must have a DIFFERENT COLORED exam than you. Be sure to write your name on BOTH your EXAM and your GREEN SCANTRON. You may write on the exam but must transfer your answers to your green scantron and you must WRITE THE EXAM COLOR ON THE GREEN SCANTRON..

1. For the reaction



the following data was collected at constant temperature. Determine the correct rate law for this reaction.

Trial	Initial [A] (mol/L)	Initial [B] (mol/L)	Initial Rate (mol/L·min)
1	0.200	0.100	$6.00 \times 10^{-2}$
2	0.100	0.100	$1.50 \times 10^{-2}$
3	0.200	0.200	$1.20 \times 10^{-1}$
4	0.300	0.200	$2.70 \times 10^{-1}$

(A)  $\frac{2}{1} = \left(\frac{6}{1.5}\right)^x \Rightarrow x = 2$

(B)  $\frac{2}{1} = \left(\frac{1.20 \times 10^{-1}}{6.00 \times 10^{-2}}\right)^2 \Rightarrow x = 1$

- A. Rate =  $k[A][B]$
- B. Rate =  $k[A][B]^2$
- C. Rate =  $k[A]^3[B]^2$
- D. Rate =  $k[A]^{1.5}[B]$
- E. Rate =  $k[A]^2[B]$

2. Dinitrogen tetroxide,  $N_2O_4$ , decomposes to nitrogen dioxide,  $NO_2$ , in a first-order process. If  $k = 2.5 \times 10^3 \text{ s}^{-1}$  at  $-5^\circ\text{C}$  and  $k = 3.5 \times 10^4 \text{ s}^{-1}$  at  $25^\circ\text{C}$ , what is the activation energy for the decomposition?

- A. 0.73 kJ/mol
- B. 58 kJ/mol
- C. 140 kJ/mol
- D. 580 kJ/mol
- E. > 1000 kJ/mol

$$\ln \frac{2.5 \times 10^3}{3.5 \times 10^4} = \frac{E_a}{8.314 \text{ J/mol}\cdot\text{K}} \left( \frac{1}{298} - \frac{1}{268} \right)$$

$$E_a = 58410 \text{ J/mol}$$

3. Tetrafluoroethylene,  $C_2F_4$ , can be converted to octafluorocyclobutane which can be used as a refrigerant or an aerosol propellant. A plot of  $1/[C_2F_4]$  vs. time gives a straight line with a slope of  $0.0448 \text{ L mol}^{-1}\text{s}^{-1}$ . What is the rate law for this reaction?

- A. Rate =  $0.0448 \text{ (L mol}^{-1}\text{s}^{-1})[C_2F_4]$
- B. Rate =  $22.3 \text{ (mol L}^{-1}\text{s)}[C_2F_4]$
- C. Rate =  $0.0448 \text{ (L mol}^{-1}\text{s}^{-1})[C_2F_4]^2$
- D. Rate =  $22.3 \text{ (mol L}^{-1}\text{s)}[C_2F_4]^2$
- E. Rate =  $0.0448 \text{ s}^{-1} [C_2F_4]$

2nd order

Key

4. A reaction has the following rate law:

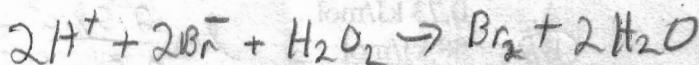
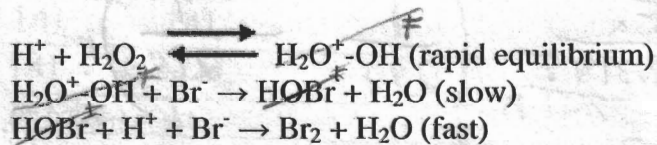
$$\text{Rate} = k[A][B]^2$$

In experiment 1, the concentrations of A and B are both  $0.10 \text{ mol L}^{-1}$ ; in experiment 2, the concentrations are both  $0.30 \text{ mol L}^{-1}$ . If the temperature stays constant, what is the value of the ratio,  $\text{Rate}(2)/\text{Rate}(1)$ ?

- A. 3.0
- B. 6.0
- C. 9.0
- D. 18
- E. 27

$$\text{rate} = k \left[ \frac{0.30}{0.10} \right] \left[ \frac{0.30}{0.10} \right]^2$$
$$3(3)^2 = 27$$

5. The kinetics of the decomposition of dinitrogen pentaoxide is studied at  $50^\circ\text{C}$  and at  $75^\circ\text{C}$ . Which of the following statements concerning the studies is correct?
- A. The rate at  $75^\circ\text{C}$  will be greater than the rate at  $50^\circ\text{C}$  because the activation energy will be lower at  $75^\circ\text{C}$  than at  $50^\circ\text{C}$ .
  - B. The rate at  $75^\circ\text{C}$  will be greater than the rate at  $50^\circ\text{C}$  because the activation energy will be higher at  $75^\circ\text{C}$  than at  $50^\circ\text{C}$ .
  - C. The rate at  $75^\circ\text{C}$  will be less than the rate at  $50^\circ\text{C}$  because the molecules at higher speeds do not interact as well as those at lower speeds.
  - D. The rate at  $75^\circ\text{C}$  will be greater than at  $50^\circ\text{C}$  because the concentration of a gas increases with increasing temperature.
  - E. The rate at  $75^\circ\text{C}$  will be greater than the rate at  $50^\circ\text{C}$  because the number of molecules with enough energy to react increases with increasing temperature.
6. Consider the following mechanism for the oxidation of bromide ions by hydrogen peroxide in aqueous acid solution.



What is the overall reaction equation for this process?

- A.  $2\text{H}_2\text{O}^+-\text{OH} + 2\text{Br}^- \rightarrow \text{H}_2\text{O}_2 + \text{Br}_2 + 2\text{H}_2\text{O}$
  - B.  $2\text{H}^+ + 2\text{Br}^- + \text{H}_2\text{O}_2 \rightarrow \text{Br}_2 + 2\text{H}_2\text{O}$
  - C.  $2\text{H}^+ + \text{H}_2\text{O}_2 + \text{Br}^- + \text{HOBr} \rightarrow \text{H}_2\text{O}^+-\text{OH} + \text{Br}_2 + \text{H}_2\text{O}$
  - D.  $\text{H}_2\text{O}^+-\text{OH} + \text{Br}^- + \text{H}^+ \rightarrow \text{Br}_2 + \text{H}_2\text{O}$
  - E. none of these choices is correct
7. Which of the following sets of units could be appropriate for a zero-order rate constant?

- A.  $\text{s}^{-1}$
- B.  $\text{L mol}^{-1} \text{s}^{-1}$
- C.  $\text{L}^2 \text{mol}^{-2} \text{s}^{-1}$
- D.  $\text{L}^3 \text{mol}^{-3} \text{s}^{-1}$
- E.  $\text{mol L}^{-1} \text{s}^{-1}$

$$\text{rate} = k$$
$$\frac{M}{t} = \frac{M}{t}$$
$$\frac{\text{mol}}{\text{L} \cdot \text{s}} = \frac{\text{mol}}{\text{L} \cdot \text{s}}$$

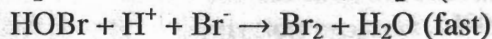
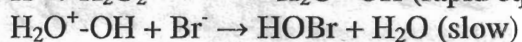
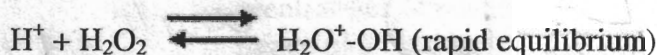
Key

8. A reaction has an activation energy of 195.0 kJ/mol. When the temperature is increased from 200. °C to 220. °C, the rate constant will increase by a factor of

- A. 1.1  
B.  $4.3 \times 10^4$   
C. 3.2  
 D. 7.5  
E. none of these choices is correct

$$\ln \frac{k_2}{k_1} = \frac{195.0 \times 10^3 \text{ J/mol}}{8.314 \frac{\text{J}}{\text{mol K}}} \left( \frac{1}{473} - \frac{1}{493} \right)$$
$$\ln \frac{k_2}{k_1} = 2.011 \quad e^{2.011} = 7.5$$

9. Consider the following mechanism for the oxidation of bromide ions by hydrogen peroxide in aqueous acid solution.



Which of the following rate laws is consistent with the mechanism?

- A. Rate =  $k[\text{H}_2\text{O}_2][\text{H}^+]^2[\text{Br}^-]$   
B. Rate =  $k[\text{H}_2\text{O}^+ \text{-OH}][\text{Br}^-]$   
 C. Rate =  $k[\text{H}_2\text{O}_2][\text{H}^+][\text{Br}^-]$   
D. Rate =  $k[\text{HOBr}][\text{H}^+][\text{Br}^-][\text{H}_2\text{O}_2]$   
E. Rate =  $k[\text{Br}^-]$

From slow: rate =  $k_2 [\text{H}_2\text{O}^+ \text{-OH}^\ddagger] [\text{Br}^-]$

From eq.  $[\text{H}_2\text{O}^+ \text{-OH}^\ddagger] = \frac{k_1}{k_{-1}} [\text{H}^+] [\text{H}_2\text{O}_2]$

Substitute:

$$\text{rate} = \frac{k_1 k_2}{k_{-1}} [\text{H}^+] [\text{Br}^-] [\text{H}_2\text{O}_2]$$

10. When a catalyst is added to a reaction mixture, it

- A. increases the rate of collisions between reactant molecules.  
B. provides reactant molecules with more energy.  
C. slows down the rate of the back reaction.  
 D. provides a new pathway (mechanism) for the reaction.  
E. does none of these.

11. Consider the following reaction



If [C] is increasing at the rate of  $4.0 \text{ mol L}^{-1} \text{ s}^{-1}$ , at what rate is [B] changing?

- A.  $-0.40 \text{ mol L}^{-1} \text{ s}^{-1}$   
 B.  $-2.5 \text{ mol L}^{-1} \text{ s}^{-1}$   
C.  $-4.0 \text{ mol L}^{-1} \text{ s}^{-1}$   
D.  $-6.4 \text{ mol L}^{-1} \text{ s}^{-1}$   
E. none of these choices is correct, since its rate of change must be positive

$$\frac{-\Delta[\text{B}]}{5 \Delta t} = \frac{\Delta[\text{C}]}{8 \Delta t} = \frac{4.0}{8} = 0.5$$
$$\Delta[\text{B}] = -\frac{5}{8} \Delta[\text{C}] = -\frac{5}{8} \cdot 4 = -2.5$$

12. When the reaction  $\text{A} \rightarrow \text{B} + \text{C}$  is studied, a plot of  $\ln[\text{A}]$ , vs. time gives a straight line with a negative slope. What is the order of the reaction?

- A. zero  
 B. first  
C. second  
D. third  
E. More information is needed to determine the order.

# Key

13. A reactant R is being consumed in a first-order reaction. What fraction of the initial R is consumed in 4.0 half-lives?

- A. 0.94
- B. 0.87
- C. 0.75
- D. 0.13
- E. 0.063

$\frac{1}{2^4}$  gives what remains ↗

so:  $1 - \frac{1}{2^4} = 0.94$

14. The rate constant for a reaction is  $4.65 \text{ L mol}^{-1} \text{ s}^{-1}$ . What is the overall order of the reaction?

- A. zero
- B. first
- C. second
- D. third
- E. More information is needed to determine the overall order.

rate =  $k[A]^x$   
 $\frac{\Delta M}{\Delta t} = \frac{\text{mol}}{\text{L} \cdot \text{s}} = \frac{\text{L}}{\text{mol} \cdot \text{s}} [ ]^2$  ← to give units that match

15. The decomposition of hydrogen peroxide is a first-order process with a rate constant of  $1.06 \times 10^{-3} \text{ min}^{-1}$ . How long will it take for the concentration of  $\text{H}_2\text{O}_2$  to drop from  $0.0200 \text{ M}$  to  $0.0120 \text{ M}$ ?

- A. < 1 min
- B. 7.55 min
- C. 481 min
- D. 4550 min
- E. 31,400 min

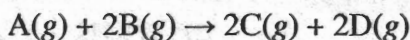
$\ln[A]_t = -kt + \ln[A]_0$   
 $\ln[0.0120] = -1.06 \times 10^{-3} t + \ln[0.0200]$   
 t = 481 min

16. Carbon-14 is a radioactive isotope which decays with a half-life of 5730 years. What is the first-order rate constant for its decay, in units of  $\text{years}^{-1}$ ?

- A.  $5.25 \times 10^{-5} \text{ years}^{-1}$
- B.  $1.21 \times 10^{-4} \text{ years}^{-1}$
- C.  $1.75 \times 10^{-4} \text{ years}^{-1}$
- D.  $3.49 \times 10^{-4} \text{ years}^{-1}$
- E.  $3.97 \times 10^3 \text{ years}^{-1}$

$\ln[A]_t = -kt + \ln[A]_0$   
 $\ln\left[\frac{[A]_t}{[A]_0}\right] = -kt$   
 at  $t = \frac{1}{2}$   $\ln \frac{1}{2} = -k(5730)$

17. For the reaction



the following data was collected at constant temperature. Determine the correct rate law for this reaction.

Trial	Initial [A] (mol/L)	Initial [B] (mol/L)	Initial Rate (mol/L·min)
1	0.125	0.200	7.25
2	0.375	0.200	21.75
3	0.250	0.400	14.50
4	0.375	0.400	21.75

← for B no change so 0 order

- A. Rate =  $k[A][B]$
- B. Rate =  $k[A]^2[B]$
- C. Rate =  $k[A][B]^2$
- D. Rate =  $k[A]$
- E. Rate =  $k[A]^3$

for A:  $\frac{21.75}{7.25} = \left(\frac{0.375}{0.125}\right)^x$   
 try 1st  $3 = (3)^x$   
 $x = 1$

Key

18. Ammonium cyanate ( $\text{NH}_4\text{CNO}$ ) reacts to form urea ( $\text{NH}_2\text{CONH}_2$ ). At  $65^\circ\text{C}$  the rate constant,  $k$ , is  $3.60 \text{ L mol}^{-1}\text{s}^{-1}$ . What is the rate law for this reaction?

- A. Rate =  $3.60 \text{ L mol}^{-1}\text{s}^{-1}[\text{NH}_4\text{CNO}]$
- B. Rate =  $3.60 \text{ L mol}^{-1}\text{s}^{-1}[\text{NH}_4\text{CNO}]^2$
- C. Rate =  $0.28 \text{ mol L}^{-1}\text{s}^{-1}[\text{NH}_4\text{CNO}]$
- D. Rate =  $0.28 \text{ mol L}^{-1}\text{s}^{-1}[\text{NH}_4\text{CNO}]^2$
- E. Rate =  $3.60 \text{ L mol}^{-1}\text{s}^{-1}[\text{NH}_2\text{CONH}_2]^{-1}$

$$\text{rate} = kC \quad \downarrow \times \text{so} \quad \frac{\text{mol}^2}{\text{L}} \\ \frac{\text{mol}}{\text{L}\cdot\text{s}} = \frac{3.60 \text{ L}}{\text{mol}\cdot\text{s}}$$

19. Which of the following affects the activation energy of a reaction?

- A. temperature of the reactants
- B. concentrations of reactants
- C. presence of a catalyst
- D. surface area of reactants
- E. reaction progress

20. The gas-phase reaction  $\text{CH}_3\text{NC} \rightarrow \text{CH}_3\text{CN}$  has been studied in a closed vessel, and the rate equation was found to be:  $\text{rate} = -\Delta[\text{CH}_3\text{NC}]/\Delta t = k[\text{CH}_3\text{NC}]$ . Which one of the following actions is least likely to cause a change in the rate of the reaction?

- A. lowering the temperature
- B. adding a catalyst
- C. using a larger initial amount of  $\text{CH}_3\text{NC}$  in the same vessel
- D. using a bigger vessel, but the same initial amount of  $\text{CH}_3\text{NC}$
- E. continuously removing  $\text{CH}_3\text{CN}$  as it is formed

21. Of the following, only \_\_\_\_\_ are not valid units for a reaction rate.

- a) mol/L
- b) M/s
- c) mol/hr
- d) g/s
- e) mol/L-hr

must have time

22. Which of the following does not play a part in determining the rate of a reaction?

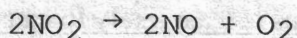
- a) temperature
- b) the concentrations of reactants
- c) the presence of a catalyst
- d) the surface area of solid or liquid reactants
- e) the equilibrium constant

23. The kinetics of the reaction below were studied and it was determined that the reaction rate did not change when the concentration of B was tripled. The reaction is \_\_\_\_\_ order in B.



- a) zero
- b) first
- c) second
- d) third
- e) one-half

24. The reaction



follows second-order kinetics. At 300°C,  $[\text{NO}_2]$  drops from 0.0100 to 0.00650 M in 100 s. Use of the integrated form of the proper rate equation affords a value of \_\_\_\_\_ L/mol·s for the rate constant  $k$ .

- a) 0.096
- b) 0.65
- c) 0.81
- d) 1.2
- e) 0.54

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$$
$$\frac{1}{0.00650} = k(100) + \frac{1}{0.0100}$$

25. One difference between first and second-order reactions is that

- a) the half-life of a first-order reaction does not depend on  $[A]_0$ ; the half-life of a second-order reaction does depend on  $[A]_0$
- b) the rate of a first-order reaction does not depend on reactant concentrations; the rate of a second-order reaction does depend on reactant concentrations
- c) the rate of a first-order reaction depends on reactant concentrations; the rate of a second-order reaction does not depend on reactant concentrations
- d) a first-order reaction can be catalyzed; a second-order reaction cannot be catalyzed
- e) the half-life of a first-order reaction depends on  $[A]_0$ ; the half-life of a second-order reaction does not depend on  $[A]_0$