

JOE PERFECT

Practice Test Chapter 14 Chemical Kinetics

Part I: Circle the letter of the best answer(s). No calculators are allowed on this portion of the practice test!

Target 1: I can list and explain the 6 factors which determine the speed of a chemical reaction.

- Which of the following does not play a part in determining the rate of a reaction?
 - temperature
 - the concentrations of reactants
 - the presence of a catalyst
 - the surface area of solid or liquid reactants
 - the equilibrium constant
- Why does increasing the temperature of a reaction speed up most chemical reactions? (Choose the best 2 answers!)
 - Particles collide more often at higher temperatures.
 - Concentration of reactants increase with increasing temperatures.
 - The activation energy is raised as a result of increased temperatures.
 - The activation energy is lowered as a result of increased temperatures.
 - Each collision is more energetic as temperature is increased.
- Increasing the surface area of reactants usually speeds up a reaction because . . .
 - you increase the total number of moles of the particles.
 - you increase the total number of grams of the particles.
 - an increase in surface area is often accompanied by an increase in temperature.
 - the concentration of the reactants is increased with increasing surface area.
 - more particles can collide as soon as the reactants are mixed.
- A burning splint will burn more vigorously in pure oxygen than in air because . . .
 - oxygen is a reactant in combustion and the concentration of oxygen is higher in pure oxygen than it is in air.
 - oxygen is a catalyst for combustion.
 - oxygen is a product of combustion.
 - nitrogen is a product of combustion and the system reaches equilibrium at a lower temperature.
 - nitrogen is a reactant in combustion and its low concentration in pure oxygen catalyzes the combustion.

Target 2: I can express the rate of a given reaction in terms of decreasing concentration of reactants and increasing concentration of products.

5. Based on the table of data shown below, the average rate of the reaction between 10 seconds and 30 seconds is _____ M/s.

A + B → C

Time (s)	[A] mol/L
0.0	0.124
10.0	0.110
20.0	0.088
30.0	0.073
40.0	0.054

$$\frac{\Delta[A]}{\Delta t} = \frac{0.073 \text{ M} - 0.110 \text{ M}}{20 \text{ sec}} = -0.00185 \frac{\text{M}}{\text{s}}$$

negative sign as [A] is decreasing w/ time.

- a. $+2.4 \times 10^{-3}$ **b.** -1.8×10^{-3} c. $+1.8 \times 10^{-3}$ d. -0.0012 e. $+0.0012$

6. Consider the following reaction: $3 \text{ A} \rightarrow 2 \text{ B}$

The average rate of appearance of B is given by $\Delta[B]/\Delta t$. How is the average rate of appearance of B related to the average rate of disappearance of A?

- a. $+2\Delta[A]/\Delta t$ **b.** $-2\Delta[A]/3\Delta t$ c. $-3\Delta[A]/2\Delta t$ d. $-\Delta[A]/\Delta t$ e. $+3\Delta[A]/2\Delta t$

7. Reaction rates usually ...
- a. increase with time. $-\frac{\Delta[A]}{3\Delta t} = \frac{\Delta[B]}{2\Delta t}$; $\frac{\Delta[B]}{\Delta t} = -\frac{2}{3} \frac{\Delta[A]}{\Delta t}$ } $2/3 \times$ as fast
- b.** decrease with time. *due to the reactants becoming less concentrated.* d. double with time.
- c. multiply with time. e. remain constant.

8. For the gas phase reaction, $3 \text{ H}_2 + \text{ N}_2 \rightarrow 2 \text{ NH}_3$, how does the rate of disappearance of H_2 compare to the rate of production of NH_3 ?
- a. The initial rates are equal.
- b. The rate of disappearance of H_2 is 1/2 the rate of appearance of NH_3 .
- c.** The rate of disappearance of H_2 is 3/2 the rate of appearance of NH_3 .
- d. The rate of disappearance of H_2 is 2/3 the rate of appearance of NH_3 .
- e. The rate of disappearance of H_2 is 1/3 the rate of appearance of NH_3 .

Target 3: I can calculate the instantaneous rates from a graph of reactant or product concentrations as a function of time.

9. Instantaneous rate is defined as ...
- a. the rate of a reaction at the start of the reaction.
- b. the rate of a reaction at the end of the reaction.
- c. the rate of a reaction in the middle of the reaction.
- d. the average rate of a reaction.
- e.** the rate of the reaction at some particular moment in time.



10. The instantaneous rate of a reaction can be determined by
- calculating the slope of a tangent line to the concentration versus time graph at a specific time.
 - calculating the slope of a tangent line to the natural log of concentration versus time graph at a specific time.
 - calculating the slope of a tangent line to the 1/concentration versus time graph at a specific time.
 - dividing the slope of the tangent of a line by 2 in a concentration versus time graph.
 - averaging the change in time and the change in concentration and then dividing the 2 quantities.

Target 4: I can determine the rate law from experimental results that show how concentration affects rate. I can calculate rate, rate constants, or reaction concentration, given two of these together with the rate law.

11. The exponents in a rate law for a multi-step reaction are determined by . . .
- the coefficients in the balanced equation
 - experimentation
 - the physical states of the reactants and products

a. 1 only b. 2 only c. 3 only d. 1 and 2 only e. 1, 2 and 3

12. For the reaction $2A + 2B \rightarrow 3C$, it was determined that the reaction is third order overall. The rate law for this reaction might be:

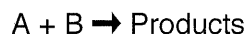
- rate = $k[A]^2[B]^2$
- rate = $k[A][B]$
- rate = $k[A][B]^2$
- rate = $k[A][B]^3$
- rate = $k[C]^3$

⇒ "C" is a product... rate laws consist of reactants not products

13. Which **one** of the following rate laws denotes a reaction that is third order overall?

- rate = $k[NO]^3 [O_2]$
- rate = $k[C_6H_{20}]^2 [H_2]$
- rate = $k[CH_4] [O_2]$
- rate = $k[A]^{-1}[B]^2$
- rate = $3k[NO]$

14. The kinetics of the reaction below were studied and it was determined that the reaction rate increased by a factor of 9 when the concentration of B was tripled. The reaction is _____ order in B.



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

$(\Delta \text{conc})^{\text{order}} = \Delta \text{rate}$

$3^x = 9$ $x = 2$

15. Use the information below to determine the rate law for the following reaction:



Experiment Number	[A] (M)	[B] (M)	Initial Rate (M/s)
1	0.273	0.763	2.83
2	0.273	1.526	2.83
3	0.819	0.763	25.47

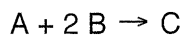
- a. rate = $k[A][B]^2$
 b. rate = $k[B]^2$
 c. rate = $k[A]^2[B]^1$
 d. rate = $k[A]^2[B]^2$
 e. rate = $k[A]^2$

Trials 1 & 2 = zero order in "B"

Trials 1 & 3 =

$$\left(\frac{0.819}{0.273}\right)^x = \frac{25.47}{2.83} ; 3^x = 9 \quad x = 2$$

16. The table below contains concentration and rate data for the reaction:



The rate law for this reaction is _____.

Experiment Number	[A] /M	[B] /M	Initial Rate (M/s)
1	0.23	0.17	0.33
2	0.46	0.17	0.66
3	0.23	0.51	0.33

- a. Rate = $k[A]^4[B]$
 b. Rate = $k[A]^2[B]$
 c. Rate = $k[A]^4$
 d. Rate = $k[A]^1$
 e. Rate = $k[A]^3$

Rate = $k[A]^1[B]^0$
 Rate = $k[A]$

17. Consider the reaction: $2 \text{ClO}_2 + 2 \text{OH}^- \rightarrow \text{ClO}_3^- + \text{ClO}_2^- + \text{H}_2\text{O}$

Number	[ClO ₂]	[OH ⁻]	Initial Rate (M/s)
1	0.060	0.030	0.0248
2	0.020	0.030	0.00276
3	0.020	0.090	0.00828

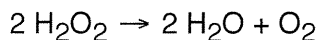
What is the order of the reaction with respect to ClO₂?

- a. 1 b. 0 c. 2 d. 3 e. 4

$$\left(\frac{1}{3}\right)^x = \frac{1}{9} \quad x = 2$$

Target 5: I can use the integrated first and second order rate law equations to solve for concentration or time.

18. The reaction below is first order in $[H_2O_2]$.



M
 0.600 } one 1/2 life
 0.300 } two 1/2 lives
 0.150 } three 1/2 lives
 0.075

A solution originally at 0.600 M H_2O_2 is found to be 0.075 M after 54 min. The half-life for this reaction is approximately _____ minutes.

- a. 6.8 **(b. 18)** c. 14 d. 28 e. 54

$$\frac{54}{3} = 18 \text{ min}$$

19. The half-life of a first-order reaction

- a. is the time necessary for the reactant concentration to drop to half its original value ☺
 b. is constant — only for first order
 c. can be calculated from the reaction rate constant — $t_{1/2} = \frac{0.693}{k}$
 d. does not depend on the initial reactant concentration
(e) all of these are correct

only for 1st order

20. What is the numerical value of the rate constant for a second order reaction if the reactant concentration drops from 0.657 M to 0.0981 M in 17.0 seconds? (Calculator OK for this one!)

- a. 8.67 **(b)** 0.510 c. 0.112 d. 0.332 e. 114

$$\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt ; \quad \frac{1}{0.0981} - \frac{1}{0.657} = k \cdot 17.0 \text{ sec} ; \quad k = 0.510$$

21. What is the half-life of a second-order reaction with the rate constant $0.0271 \text{ M}^{-1}\text{s}^{-1}$ and the initial reactant concentration 0.902 M? (Calculator OK for this one!)

- (a)** 40.9 s
 b. 25.6 s
 c. not enough information to determine
 d. 0.451 s
 e. 0.625 s

$$t_{1/2} = \frac{1}{k[A]_0} = \frac{1}{(0.0271 \frac{1}{\text{M}\cdot\text{s}})(0.902 \text{ M})} = 40.9 \text{ sec}$$

No need to memorize this equation!
 It will be given to you!

22. The initial concentration of reactant in a first-order reaction is 0.27 M. The rate constant for the reaction is 0.75 s^{-1} . What is the concentration (mol/L) of reactant after 1.5 s? (Calculator OK for this one!)

- a. 3.8 b. 1.7 **(c)** 8.8×10^{-2} d. 2.0×10^{-2} e. 0.135

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

$$\ln [A]_t - \ln (0.27) = -(0.75 \text{ s}^{-1})(1.5 \text{ sec})$$

$$\ln [A]_t = -2.434$$

$$[A]_t = 0.0877 \text{ M}$$

23. The reaction $A \rightarrow B$ is first order in $[A]$. Using the data below, the rate constant for this reaction is ____ s^{-1} . (Calculator OK for this one!)

time (s)	$[A]$ (M)
0.0	1.60
5.0	0.80
10.0	0.40
15.0	0.20
20.0	0.10

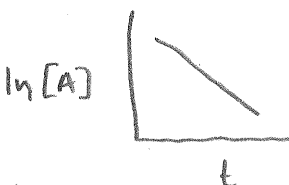
$t_{1/2} = 5.0 \text{ sec}$ $k = \frac{0.693}{5.0 \text{ sec}} = 0.1386 \text{ s}^{-1}$

- a. 0.013 b. 0.030 **c. 0.14** d. 3.0 e. 3.1×10^{-3}

Target 6: I can determine whether or not a reaction is first order or second order by analyzing graphs of experimental data.

24. Which one of the following pairs would you graph in order to produce a straight line for a first order reaction?

- a.** $\ln[A]$ versus time
 b. $1/\ln[A]$ versus time
 c. $1/[A]$ versus time
 d. $\ln 1/[A]$ versus time
 e. $\log 1/\ln 1$ versus E_a/RT



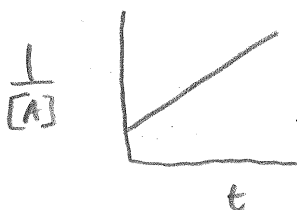
$$\ln[A]_t - \ln[A]_0 = -kt$$

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

$$y = mx + b$$

25. Which one of the following pairs would you graph in order to produce a straight line for a second order reaction?

- a. $\ln[A]$ versus time
 b. $1/\ln[A]$ versus time
c. $1/[A]$ versus time
 d. $\ln 1/[A]$ versus time
 e. $\log 1/\ln 1$ versus E_a/RT



$$\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt$$

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$$

$$y = mx + b$$

26. In analyzing a linear graph of $\ln[A]$ versus time for a particular reaction, you could determine the value of the rate constant by . . .

- a. taking the reciprocal of the slope of the line
b. multiplying the slope of the line by -1
 c. dividing 0.693 by the slope of the line
 d. multiplying 0.693 by the slope of the line
 e. dividing E_a by RT

FIRST ORDER . . .

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

$$y = mx + b$$

$(m = -k)$

→ 2nd order

$$\frac{1}{[A]}_t = kt + \frac{1}{[A]_0}$$
$$y = mx + b$$

27. In analyzing a linear graph of $1/[A]$ versus time for a particular reaction, the value of the rate constant is . . .
- determined by multiplying the slope of the line by -1
 - determined by calculating the tangent to the curve at any point on the curve.
 - the same as E_a/RT .
 - the same as $\ln E_a/RT$.
 - the same as the slope of the line.

Target 7: I can explain the concept of activation energy and how it relates to the variation of reaction rate with temperature.

28. Which 2 statements are TRUE regarding the activation energy of a reaction?
- The activation energy of a reaction is higher at lower temperatures.
 - The activation energy of a reaction is lower at lower temperatures.
 - The activation energy of a reaction is independent of temperature.
 - The activation energy of a reaction is lowered by adding a catalyst.
 - The higher the activation energy of a reaction, the faster the reaction.

29. The rate constant for a particular reaction is $1.3 \times 10^{-4} \text{ M}^{-1}\text{s}^{-1}$ at 150°C , and $1.1 \times 10^{-3} \text{ M}^{-1}\text{s}^{-1}$ at 200°C . What is the energy of activation (in kJ/mol) for this reaction at 250°C ? (Calculator OK for this one!)

Since E_a is independent of temp., you do not need 250° ... this is just extra info!

- 132
- 56
- 99
- 71
- 22

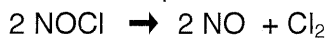
$$\ln \frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

Always use KELVIN!

$$\ln \frac{1.3 \times 10^{-4}}{1.1 \times 10^{-3}} = \frac{E_a}{8.314} \left(\frac{1}{473} - \frac{1}{423} \right)$$

$E_a = 71,047 \text{ J/mol}$
 $E_a = 71 \text{ kJ/mol}$

30. What is the slope of an Arrhenius plot for the following reaction? (Calc's OK)



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

Use 2 of the temp's to find E_a . Then, from Arrhenius equation... $m = -E_a/R$

$$\ln \frac{6.6 \times 10^{-4}}{2.9 \times 10^{-1}} = \frac{E_a}{8.314} \left(\frac{1}{500} - \frac{1}{400} \right)$$

$$E_a = 101,187 \text{ J/mole}$$

- 8.18×10^{-6}
- 5.06×10^{-2}
- -1.21×10^4
- -1.96×10^7
- 1.00

$$m = \frac{-E_a}{R} = \frac{-101,187 \text{ J/mole}}{8.314 \text{ J/mol}\cdot\text{K}} = -1.2 \times 10^4$$

31. A reaction with activation energy of 123 kJ/mol has a rate constant of 0.200 s^{-1} at 311 K. At a temperature of _____ K, the rate constant will be double that at 311 K. (Calculator OK for this one!)

- 304
- 316
- 622
- 349
- 246

$$\ln \frac{1}{2} = \frac{123,000 \text{ J/mol}}{8.314 \text{ J/mol}\cdot\text{K}} \left(\frac{1}{k_2} - \frac{1}{311} \right); k_2 = 316$$

Intermediates can be stable and can therefore sometimes be identified. An activated complex (a transition state) are always UNSTABLE and can never be identified.

Target 8: I can explain what is meant by the mechanism of a reaction using the terms elementary steps, rate-determining step, and intermediate.

Consider the following reaction mechanism for #32 and 33.



32. The catalyst is B and the intermediate is C.

- a. A, B b. B, D c. C, A d. C, B **e. B, C**

33. The rate determining step is _____.

- a. step 1 **b. step 2** c. step 1 and step 2 d. neither step 1 or step 2

34. Which of the following statements is/are true?

- a.** If we know a reaction is an elementary reaction, then we know its rate law. *From stoich of the equation.*
b. The rate-determining step in a mechanism is the slowest step.
 c. An intermediate is identical to an activated complex.
d. A reaction mechanism is often difficult to prove. \rightarrow We can only make reasonable guesses.
 e. It is impossible to determine the rate law for a reaction.

35. A proposed mechanism for a reaction is

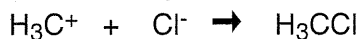


What is the overall reaction conforming to this proposed mechanism?

- a. $CH_3Cl + H_2O \rightarrow CH_3OH + HCl$
b. $CH_3OH + HCl \rightarrow CH_3Cl + H_2O$
 c. $CH_3Cl \rightarrow CH_3^+ + Cl^-$
 d. $CH_3^+ + Cl^- \rightarrow CH_3Cl$
 e. $HCl \rightarrow H^+ + Cl^-$

Target 9: I can derive the rate law for a reaction that has a rate-determining step, given the elementary steps and their speeds; or, conversely, choose a plausible mechanism for a reaction given the rate law.

36. What is the molecularity of the elementary chemical step:



- a. unimolecular **b. bimolecular** c. termolecular d. equimolecular

37. Which of the following rate expressions is to be expected for the reaction,

$A + 2B \rightarrow E$, given the mechanism below:

Step 1: $A + B \rightarrow C$ slow step

Step 2: $C \rightarrow D$ fast step

Step 3: $D + B \rightarrow E$ fast step

- a. rate = $k [C]^2$
- b. rate = $k [A]^2 [B]^3$
- c. rate = $k [A]^2 [B]^2$
- d. rate = $k [A] [B]$
- e. rate = $k [A] [B] [C]$

\Rightarrow Rate = $k [A][B]$; rate law is always derived from the slow step!

38. The experimentally determined rate law for the reaction $A + 2B \rightarrow C + D$ is

rate = $k[A]^2$.

Which of the following mechanisms is plausible for this reaction?

- a. step 1: $A \rightarrow E + D + 2B$
step 2: $E + 2B \rightarrow C + A$

(slow) Rate = $k [A]$
(fast)

- b. step 1: $A + B \rightarrow E + C$
step 2: $E + B \rightarrow D$

(slow) Rate = $k [A][B]$
(fast)

- c. step 1: $A + A \rightarrow E + D$
step 2: $E + 2B \rightarrow C + A$
 $A + 2B \rightarrow C + D$

(slow) Rate = $k [A]^2$
(fast)

- d. step 1: $A + A + B \rightarrow E + C$
step 2: $E + C \rightarrow D + F$
step 3: $F \rightarrow C$

(slow) Rate = $k [A]^2 [B]$
(fast)
(fast)

- e. None of these mechanisms are plausible for this reaction.

39. A possible mechanism for the stoichiometric reaction $Br_2 + 2NO \rightarrow 2NOBr$ is shown below. The rate law is _____.

$2NO \rightleftharpoons N_2O_2$ (fast, equilibrium) (1, -1)

$N_2O_2 + Br_2 \rightarrow 2NOBr$ (slow) (2)

From slow step...

$$\text{Rate} = k_2 [N_2O_2] [Br_2]$$

a. $k_1 [NO]^2$

b. $k_1 [Br_2]^2$

c. $\left(\frac{k_2 k_1}{k^{-1}}\right) [NO]^2 [Br_2]$

d. $\left(\frac{k_1}{k^{-1}}\right) [NO]^2$

e. $\left(\frac{k_2 k_1}{k^{-1}}\right) [NO] [Br_2]^2$

From equilibrium step...

$$k_1 [NO]^2 = k_{-1} [N_2O_2]$$

$$[N_2O_2] = \frac{k_1}{k_{-1}} [NO]^2$$

Plug into rate law for slow step!

$$\text{Rate} = \frac{k_2 k_1}{k_{-1}} [NO]^2 [Br_2]$$

$$\text{Rate} = k [NO]^2 [Br_2]$$

Target 10: I can describe the effects of a catalyst on the energy requirements for a reaction.

40. A catalyst . . .

a. increases the rate of the forward reaction, but does not alter the rate of the reverse reaction.

b. alters the reaction mechanism. \rightarrow The E_a of new mechanism is lower than

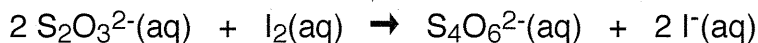
c. increases the average KE of the reactant particles. E_a of old mechanism!

d. increases the concentration of the reactants.

e. raises the activation energy of a reaction.

Part II: Solve each of the following problems. Label all of your answers with correct units and circle your final answers.

1. The thiosulfate ion is oxidized by iodine according to the following reaction:



If, in a certain experiment, 0.0080 mol $\text{S}_2\text{O}_3^{2-}$ is consumed in 1.0 L of solution each second,

a) what is the rate of consumption of I_2 ?

$$\frac{1}{2} (0.0080 \text{ M/s}) = \boxed{0.0040 \text{ M/s}}$$

or -0.0040 M/s

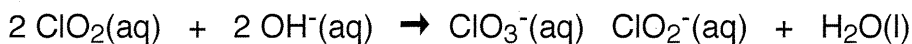
b) at what rate is $\text{S}_4\text{O}_6^{2-}$ produced?

$$\boxed{0.0040 \text{ M/s}}$$
 same as I_2 used up!

c) at what rate is I^- produced?

$$\boxed{0.0080 \text{ M/s}}$$
 same as $\text{S}_2\text{O}_3^{2-}$ used up!

2. The following rate data were obtained for the reaction:



where the rate = $-\Delta[\text{ClO}_2]/\Delta t$

$[\text{ClO}_2], \text{M}$	$[\text{OH}^-], \text{M}$	Initial rate (M/s)
0.0500	0.100	5.75×10^{-2}
0.100	0.100	2.30×10^{-1}
0.100	0.050	1.15×10^{-1}

Determine the rate law and the value for the rate constant.

From trials 1 & 2 ... 2nd order in ClO_2 .

From trials 2 & 3 ... 1st order in OH^- .

$$\boxed{\text{Rate} = k [\text{ClO}_2]^2 [\text{OH}^-]}$$

FROM TRIAL #1 ... $5.75 \times 10^{-2} \text{ M/s} = k (0.0500 \text{ M})^2 (0.100 \text{ M})$

$$\boxed{k = 230 \text{ M}^{-2} \text{ s}^{-1}}$$

3. It took 143 seconds for 50.0% of a particular substance to decompose. If the initial concentration was 0.060 M and the decomposition reaction follows second-order kinetics, what is the value of the rate constant?

$$t_{1/2} = \frac{1}{k[A]_0} ; 143 \text{ sec} = \frac{1}{k(0.060 \text{ M})} ; k = 0.12 \text{ M}^{-1}\text{s}^{-1}$$

4. At 25°C the first order rate constant for a reaction is $2.0 \times 10^3 \text{ s}^{-1}$. The activation energy is 15.0 kJ/mol. What is the value of the rate constant at 75°C?

$$T_1 = 298 \text{ K}$$

$$T_2 = 348 \text{ K}$$

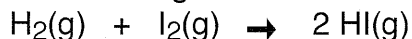
$$\ln \frac{2.0 \times 10^3}{k_2} = \frac{15,000 \text{ J/mol}}{8.314 \text{ J/mol}\cdot\text{K}} \left(\frac{1}{348} - \frac{1}{298} \right)$$

$$\ln 2.0 \times 10^3 - \ln k_2 = -0.8699 ; \ln k_2 = 8.471 ; k_2 = 4774$$

$$k_2 = 4800 \text{ s}^{-1}$$

2 sig fig's

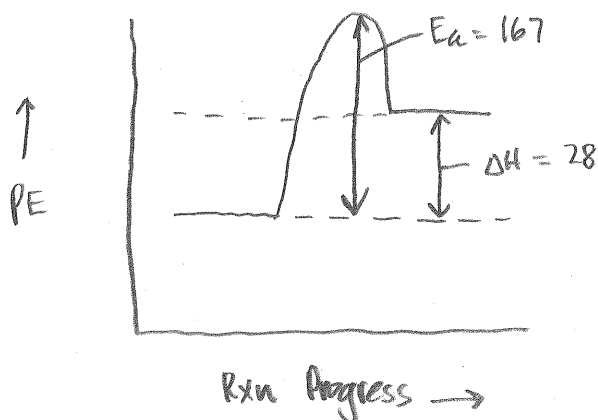
5. Consider the following reaction:



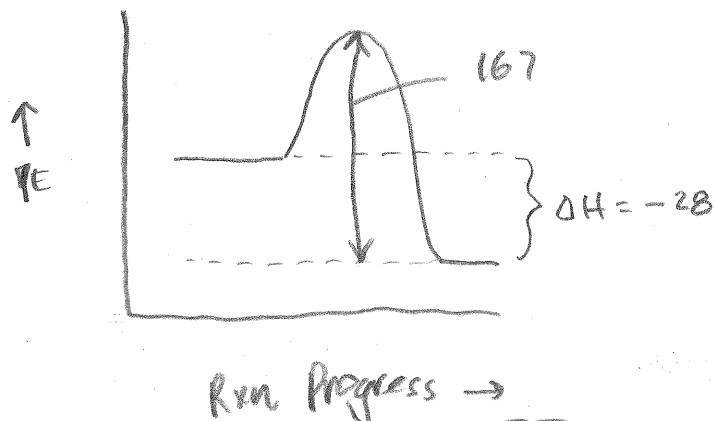
The activation energy for the reaction is 167 kJ/mol and ΔH for the reaction is +28 kJ/mol. What is the activation energy for the decomposition of HI? Explain how you got your answer.

↳ reverse reaction!

For given reaction...
(ENDOTHERMIC)



For reverse reaction...
(EXOTHERMIC)

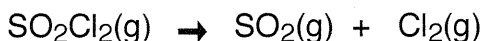


$$E_a = 167 - 28 = 139 \text{ kJ/mole}$$

6. Sulfuryl chloride, SO_2Cl_2 , decomposes in the gas phase into sulfur dioxide, SO_2 , and chlorine, Cl_2 . The partial pressure of the SO_2Cl_2 is monitored over time. It is found that the plot of $\ln[\text{SO}_2\text{Cl}_2]$ versus time is linear, and that in 66.0 seconds the partial pressure of SO_2Cl_2 decreased from 135 torr to 98 torr.

1st order

a) What is the rate constant for the reaction:



$$\ln [A]_t - \ln [A]_0 = -kt ; \ln(98) - \ln(135) = -k(66 \text{ sec}) ; k = 0.00485 \text{ s}^{-1}$$

b) What is the half-life of the reaction?

$$t_{1/2} = \frac{0.693}{0.00485 \text{ s}^{-1}} = 143 \text{ sec}$$

7. In an experiment, all the air in a rigid 2.0 L flask is pumped out. Then some liquid ethanol is injected into the sealed flask, which is held at 35°C . The amount of liquid ethanol initially decreases, but after five minutes the amount of liquid ethanol in the flask remains constant. Ethanol has a boiling point of 78.5°C and an equilibrium vapor pressure of 100 torr at 35°C .

(a) When the amount of liquid ethanol in the flask is constant, is the pressure in the flask greater than, less than, or equal to 100 torr? Justify your answer.

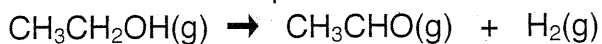
The pressure would be EQUAL to 100 torr. Because the quantity of liquid ethanol is not changing, the gas & liquid phase have reached equilibrium. Therefore, the pressure of ethanol in gas phase equals vapor pressure.

(b) The flask is then heated to 45°C , and the pressure in the flask increases. In terms of kinetic molecular theory, provide TWO reasons that the pressure in the flask is greater at 45°C than at 35°C .

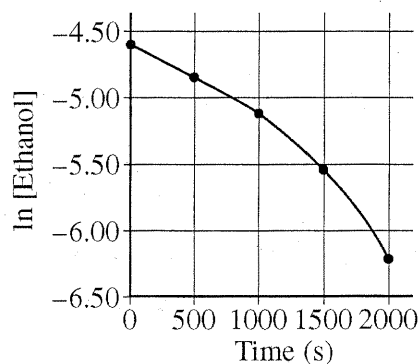
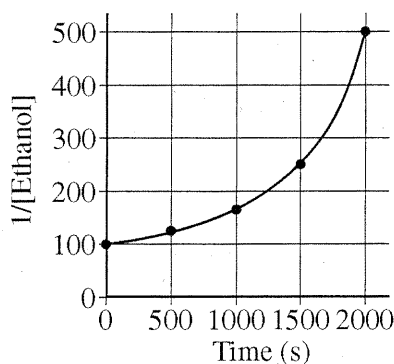
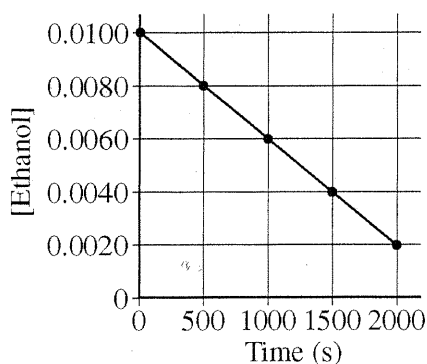
3 POSSIBLE ANSWERS

- ① At higher temperature, there are more ethanol gas molecules, so there are more collisions with flask walls resulting in greater pressure.
- ② At higher temperature, molecules move faster thus colliding with walls more frequently resulting in greater pressure.
- ③ Because molecules are moving faster, their collisions with walls will exert more force resulting in greater pressure.

In a second experiment, which is performed at a much higher temperature, a sample of ethanol gas and a copper catalyst are placed in a rigid, empty 1.0 L flask. The temperature of the flask is held constant, and the initial concentration of the ethanol gas is 0.0100 M. The ethanol begins to decompose according to the chemical reaction represented below.



The concentration of ethanol gas over time is used to create the three graphs below.



(c) Given that the reaction order is zero, one, or two, use the information in the graphs to respond to the following.

(i) Determine the order of the reaction with respect to ethanol. Justify your answer. *The order of reaction is ZERO! The graph on the left is a straight line indicating the rate of decrease of ethanol is constant. The rate of the reaction does NOT depend on concentration.*

(ii) Write the rate law for the reaction.

$$\text{Rate} = k$$

(iii) Determine the rate constant for the reaction, including units.

$$\text{rate} = k = - \frac{\Delta [\text{ethanol}]}{\Delta t} = - \frac{(0.0020 - 0.0100) \text{ M}}{2000 \text{ sec}} = -4.0 \times 10^{-6} \text{ M/s}$$

(d) The pressure in the flask at the beginning of the experiment is 0.40 atm. If the ethanol completely decomposes, what is the final pressure in the flask?

The final pressure is 0.80 atm. This is twice the original pressure. The reaction produces 2 moles of gas for every 1 mole consumed!