Chemistry 104 – Dr. Batterton
Practice Problems that may help you prepare for the 2nd exam
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Chapter 15 - Kinetics

1. Which of the following will affect the value of the rate constant?
   I. Changing the temperature of the reaction.
   II. Doubling the concentration of the product.
   III. Doubling the concentration of a reactant that is part of the rate law.

a. I b. III c. I and III d. I and II e. II and III

2. For the reaction, $CS_2(g) + 4H_2(g) \rightarrow CH_4(g) + 2H_2S(g)$, if the rate of consumption of H_2 was 0.1M/min, what would be the rate of production of H_2S?

3. The rate of a reaction is dependent on what factors?

4. For the reaction, $3C_3H_8O + Cr_2O_7^{-2} + 8H^+ \rightarrow 3C_3H_6O + 2Cr^{+3} + 7H_2$, if the rate of consumption of H^+ , rate = $-\Delta[H^+]/\Delta t$, was 0.50M/min, what would be the rate of consumption for C_3H_8O, rate = $-\Delta[C_3H_8O]/\Delta t$?

5. In general, when the ***temperature increases***, the rate of the reaction will
   I. decreases because there are fewer collisions with the proper orientation.
   II. decrease for exothermic reactions.
   III. increase because the activation energy will decrease.
   IV. increase because the number of collisions with sufficient energy is greater.

a. I b. II c. III d. IV e. none

6. a. The hydrolysis of ethyl propanoate ester, $CH_3CH_2COOCH_2CH_3$, is performed:

$$CH_3CH_2COOCH_2CH_3 + H_2O \rightarrow CH_3CH_2COOH + CH_3CH_2OH$$

The concentration of the ester as it hydrolyzes is observed and its value is recorded.

time (s)	**[ester]**
0	1.00
200	0.88
400	0.78
600	0.69
800	0.61
1200	0.48
1400	0.43
1800	0.34
2000	0.30

a. What is the value of the initial rate of this reaction?
b. What is the concentration of the ester at 600sec?
c. What is the value of the rate of ester consumption at 600sec?
d. Assume the order of the H_2O can be ignored, and the order of the ester is found to be first order. With this information, write the rate expression.
e. What is the value of the rate constant? Include units.
f. What is the half-life of this reaction?
g. What is the concentration of the ester after 3 half lives?

(watch some TV)

7. In the given reaction, $A + 2B + 3C \rightarrow 2BC + AC$ with a rate law of rate = $k[A]^2[B]^0[C]$, when the concentrations of all reactants are doubled, the rate will

a. increase by a factor of 2 b. increase by a factor of 6 c. increase by a factor of 8
d. increase by a factor of 16 e. decrease by a factor of 4

8. An experiment involving NO and O_2 is completed and the data collected is shown below.

$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

Experiment	**$[NO]_0$**	**$[O_2]_0$**	**Initial Rate O_2 (M/s)**
1	1.25×10^{-2}	2.50×10^{-2}	2.80×10^{-2}
2	2.50×10^{-2}	2.50×10^{-2}	1.12×10^{-1}
3	1.25×10^{-2}	5.00×10^{-2}	5.60×10^{-2}
4	5.00×10^{-2}	5.00×10^{-2}	?

a. Given the data above, determine the rate law.
b. What is the reaction order?
c. Determine the rate constant and include units.
d. What is the initial rate for Experiment 4?

9. Given the reaction below and the initial rate data, answer the following questions.

$NH_4^+(aq) + NO_2^-(aq) \rightarrow N_2(g) + 2H_2O(l)$

Experiment	**$[NH_4^+]_0$**	**$[NO_2^-]_0$**	**Initial Rate $[NH_4^+]$ (M/s)**
1	0.100	0.0050	1.35×10^{-7}
2	0.100	0.010	2.70×10^{-7}
3	0.200	0.010	5.40×10^{-7}

a. What is the order of NH_4^+ in this reaction?
b. What is the order of NO_2^- in this reaction?
c. Write the rate law.
d. What is the value of the rate constant? Include units.
e. If $[NH_4^+] = 0.15M$ and $[NO_2^-] = 0.015M$, what would be the initial rate? Include units.

10. A graph of ln[A] versus time yields a straight line with a slope of -1.55×10^{-4} M/min
a. What is the order with respect to A?
b. What is the value of the rate constant?
c. What is the half-life of A?
d. If the initial concentration of A was 2.75×10^{-2}M, how much of A remains after 60 minutes?

11. Sulfuryl chloride, SO_2Cl_2, decomposes when heated:

$SO_2Cl_2(g) \rightarrow SO_2(g) + Cl_2(g)$

In an experiment, the initial concentration of SO_2Cl_2 was 0.0248 M. If the rate constant is 2.2×10^{-5}/s, what is the concentration of SO_2Cl_2 after 4.5 hours? The reaction is first order. (Hint: Convert hours into seconds.)

12. How long (in years) does it take for a sample of ^{226}Ra ($t_{1/2}$ = 1602yr) to decay by 10%? (Ra decays as a 1st order reaction.)
(take a nap)

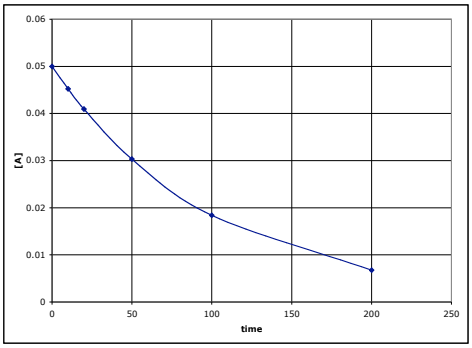
13. A boat was found at the top of a mountain in a receding glacier. The ^{14}C activity of 8.75 dpm/g (disintegrations per minute per gram). What is the approximate age in years of the boat given that the half-life of ^{14}C is 5.73×10^3 years and the activity of ^{14}C found in living materials is 12.6 dpm/g? (Hint: Treat 12.6dpm/g as initial value, and radioactive decays are first order.)

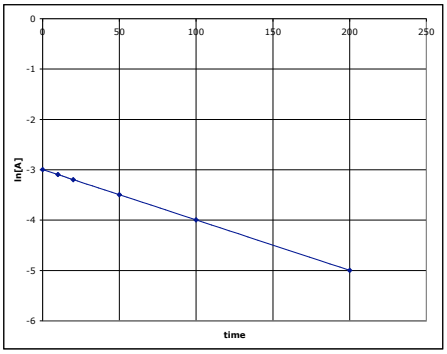
14. Which of the following statements is ***incorrect*** for reactions that follow first order kinetics?
a. As the rate constant increases, the half-life decreases.
b. If the rate increases, the rate constant also must increase.
c. If the temperature increases, the half-life usually decreases.
d. As the reactant concentration increases, the half-life remains unchanged.

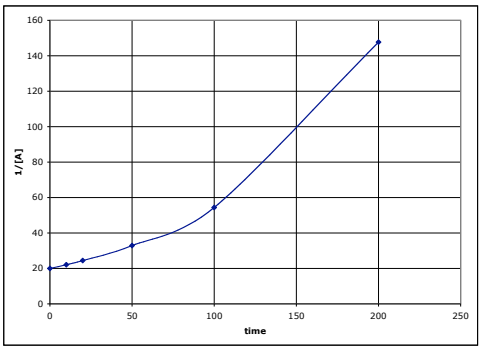
15. For the reaction, $2BrNO(g) \rightarrow 2NO(g) + Br_2(g)$, a plot of ln[BrNO] versus times yields a straight line with a slope = $-2.0 \times 10^{-5}\ s^{-1}$. What is the half-life of the reaction?

16. For the second order reaction $A \rightarrow B$, with an initial concentration of $A = 5.00 \times 10^{-4}$ M, and a rate constant = $150.0\ s^{-1}$, what will the concentration of A be after 2.00 half-lives? b. How long in seconds did it take for the first half-life?

17. The reaction, $2A \rightarrow 2C + B$ was analyzed and kinetic data was plotted below.
a. What is the rate law for this reaction? b. What is the value of the rate constant?
c. What was the initial concentration of A? d. What is the concentration of A after 120s?







18. Which of the following statements is ***incorrect*** for catalyzed versus uncatalyzed reactions?
I. The energy of the reactants and products are the same.
II. The overall reaction is unchanged.
III. The rate of the forward and reverse reactions in the catalyzed reaction are faster.
IV. The activation energy for the uncatalyzed reaction is lower than the catalyzed reaction.
V. The reaction coordinate diagram for the catalyzed and uncatalyzed reactions are different.

a. I b. II c. III d. IV e. V

19. Use the following mechanism to answer the questions below.

$NO + {}^1/_2O_2 \rightarrow NO_2$	slow
$NO_2 \rightarrow NO + O$	fast
$O_2 + O \rightarrow O_3$	fast

a. What catalyst, if any, is present in the above mechanism?
b. What intermediate(s), if any, are present in the above mechanism?
c. What are the product(s) in the overall reaction?
d. What is the rate law as determined from the above mechanism?
(eat some pizza)

20. An intermediate, in general, is a species that
I. speeds up a reaction by lowering the activation energy
II. slows down a reaction by decreasing the value of the rate constant.
III. is produced in one step of a mechanism and then consumed in a later step.
IV. increases the number of molecules with the correct orientation and thereby increases the reaction rate.

a. I b. II c. III d. IV e. none

21. A 2-step mechanism is postulated for the reaction $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$. The initial step is a fast slightly endothermic step, the second step is a slow, exothermic step, and the overall reaction is exothermic. Draw a reaction coordinate diagram labeling the axes, the activation energy for each forward step and the change in energy for the overall reaction.

22. The reaction between molecule A and molecule B is found to go through an intermediate, AB. Would the mechanism for the reaction be a single step mechanism or a multi-step mechanism?

23. The reaction $2NO + O_2 \rightarrow 2NO_2$ obeys the rate law, rate = $k[NO]^2[O_2]$. Which mechanism is consistent with this reaction and rate law?

a. $NO + NO \rightarrow N_2O_2$ (slow)
$N_2O_2 + O_2 \rightarrow 2NO_2$ (fast)

b. $NO + O_2 \rightarrow NO_3$ (fast, equilibrium)
$NO_3 + NO \rightarrow 2NO_2$ (slow)

c. $2NO \rightarrow N_2O_2$ (fast, equilibrium)
$N_2O_2 \rightarrow NO_2 + N$ (slow)
$N + O_2 \rightarrow NO_2$ (fast)

24. In collisional theory, which of the following statements is ***incorrect*** for a reaction to occur?
I. the molecule must collide
II. the molecules must have sufficient energy to overcome the activation energy of the reaction.
III. no intermediate can be formed since these are unstable molecules.
IV. the proper orientation of the molecules colliding is required.

a. I b. II c. III d. IV e. none

25. Which of the following statements in ***incorrect***?
I. As the temperature is increased, the activation energy increases.
II. In the Arrhenius equation, a plot of lnk versus 1/T has a slope of $-k_{obs}$.
III. If the concentration of a reactant is double, the rate will at least double.

a. I b. II c. I and II d. I and III e. I, II, and III

26. For the reaction, $A + B \rightarrow C$, the rate constant increases from $1100s^{-1}$ to $12{,}000s^{-1}$ when the temperature is increased from 25°C to 60°C. What is the activation energy for the reaction in J/mol?

a. 5.6×10^4 b. 5.6×10^2 c. 2.6×10^5 d. 5.7×10^{-1} e. 2.39

27. For the reaction, $2A(g) + B(g) \rightarrow 2C(g)$, the rate constant at 373K was $1.0 \times 10^{-2}\ s^{-1}$. At what temperature would the reaction rate double if the activation energy was 1×10^4J?
(another nap)

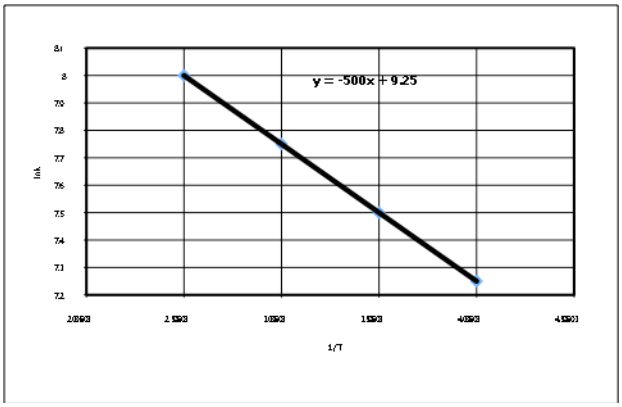
28. The reaction $2NO_2 \rightarrow 2NO + O_2$ has a rate law = $1.4 \times 10^{-2}[NO_2]^2$ at 500K. What is the value of the rate constant at 298K if the activation energy is 80kJ/mol?

a. 3.0×10^{-8} b. 6.5×10^{3} c. 1.3×10^{-15} d. 4.8×10^{-5} e. 4.0×10^{-26}

29. Consider the plot below. Which conclusion is correct?

a. The reaction is second order b. A is about 8 c. $E_a = 4200J$

d. $E_a = 500J$ e. The rate constant increases with decreasing temperature.



Chapter 16 - Equilibrium

30. Which of the following statements is ***incorrect***?

a. When equilibrium is reached, the concentrations of the reactants and products are constant.

b. Individual molecules of reactants can form products and individual molecules of products can form reactants at equilibrium giving rise to a dynamic equilibrium.

c. The substances that are solids or pure liquids are not included in the equilibrium expression.

d. At equilibrium, the forward rate constant and reverse rate constant are equal.

e. All of the above statements are true.

31. If the equilibrium constant for the reaction $H_2(g) + I_2(g) \rightarrow 2HI(g)$ is 50.3, what is the equilibrium constant for $HI(g) \rightarrow {}^1/_2H_2(g) + {}^1/_2I_2(g)$?

a. 50.3 b. 0.0199 c. 7.09 d. 0.141 e. 0.0398

32. a. Given the following reactions with their respective equilibrium constants,

Rxn I: $C(s) + 2H_2(g) \rightarrow CH_4(g)$ $K = K_1$

Rxn II: $C(s) + 2Cl_2 \rightarrow CCl_4(l)$ $K = K_2$

Rxn III: $\frac{1}{2} H_2(g) + \frac{1}{2} Cl_2(g) \rightarrow HCl(g)$ $K = K_3$

What is the equilibrium constant for: $CH_4(g) + 4Cl_2(g) \rightarrow CCl_4(l) + 4HCl(g)$ is

33. For the reaction, $N_2(g) + O_2(g) \rightarrow 2NO(g)$, it can be stated that

a. $K_p > K_c$ b. $K_p < K_c$ c. $K_p = K_c$

d. Which is larger, K_c or K_p, is determined by the temperature.

34. The value of K_p is 3.02×10^{-5} for the following reaction at 25°C. What is the value of K_c for this reaction also at 25°C?

$CS_2(g) + 4H_2(g) \rightarrow CH_4(g) + 2H_2S(g)$

(a quick trip to Florida)

35. A 5.00-L vessel contained 0.015 mol phosphorus trichloride, 0.018 mol phosphorus pentachloride, and 0.087 mol chlorine at 230 ˚C in an equilibrium mixture. Calculate the value of K for the reaction

$PCl_3(g) + Cl_2(g) \rightarrow PCl_5(g)$

36. Consider the reaction $CO(g) + H_2O(g) \rightarrow CO_2(g) + H_2(g)$

When 1.0mol of CO is mixed with 0.50mol H_2O in a 1.0-L flask, it is found that 0.42 mol of CO_2 is created at equilibrium at 500K. What is the value of the equilibrium constant?

37. Initially 1.5 moles of A are added to a 3.0 L reaction vessel and the system is allowed to come to equilibrium. At equilibrium, the amount of A has decreased by 35%.

$2A \rightarrow 2B + C$

a. What is the concentration of the reactants and products at equilibrium? b. What is the value of K?

38. Nitrogen and oxygen react to give nitric oxide, NO.

$N_2(g) + O_2(g) \rightarrow 2NO(g)$ $\qquad K = 1.00 \times 10^{-2}$

What is the equilibrium concentration of N_2, O_2, and NO in a reaction vessel that initially contained 0.106M each of N_2 and O_2?

39. The equilibrium constant for $N_2(g) + O_2(g) \rightarrow 2NO(g)$ is 0.00250 at 2127 ˚C. Calculate the equilibrium concentrations when 2.00M of N_2 and 3.00M of O_2 are mixed.

40. What are the equilibrium concentrations of NO_2 and N_2O_4 if initially there was 0.00150 mol of N_2O_4 in a 5.00 L flask, and the K_c at 25 ˚C is 0.125? (Hint: Use the quadratic equation to solve it.)

$N_2O_4(g) \rightarrow 2NO_2(g)$

41. For the following reaction, K = 10.5.

$CO(g) + 2H_2(g) \rightarrow CH_3OH(g)$

If the initial concentrations of reactants and products are as follows: [CO] = 0.00135 M, $[H_2]$ = 0.00226 M, $[CH_3OH]$ = 1.33 M, is the system at equilibrium? If not, in which direction will the reaction proceed to reach equilibrium?

42. A 2.00-L vessel contains 2.0 mol N_2, 1.0 mol H_2, and 3.0 mol NH_3. The following reaction with a K = 0.51 occurs: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

a. Calculate the reaction quotient, Q for the reaction.

b. The concentrations of which chemical(s) will increase and which chemical(s) will decrease as the reaction proceeds to equilibrium?

(yea done)

ANSWERS

1. a {changing the T changes the rate constant as per the Arrhenius reaction}

2. 0.05M/min $\left\{\frac{\Delta[H_2S]}{\Delta t} = \frac{1}{2}\left(-\frac{\Delta[H_2]}{\Delta t}\right) = (^1/_2)(0.1) = 0.05\text{M/min}\right\}$

3. Temperature, concentration, catalyst, and particle size

4. 0.19 M/min $\left\{\frac{1}{3}\left(-\frac{\Delta[C_3H_8O]}{\Delta t}\right) = \frac{1}{8}\left(-\frac{\Delta[H^+]}{\Delta t}\right);\ \frac{\Delta[C_3H_8O]}{\Delta t} = \frac{3}{8}\frac{\Delta[H^+]}{\Delta t} = (^3/_8)(0.5) = 0.1875\right\}$

5. d {rate increases so "a" and "b" are not correct; a catalyst lowers E_a not T; energy of molecules > E_a increases → "d"}

6. a. 6. x 10^{-4} M/s {rate = $-\Delta[]/\Delta t = -(1.00 - 0.88)/(0 - 200) = 6.0 \times 10^{-4}$ M/s}

b. 0.69M {read right off the table, 0.69M}

c. 4.25 x 10^{-4} M/s {rate = $-\Delta[]/\Delta t = (0.78 - 0.61)/(400 - 800) = 4.25 \times 10^{-4}$ M/s}

d. rate = $k[\text{ester}]^1 = k[CH_3CH_2COOCH_2CH_3]^1$

e. k = 6.00 x 10^{-4} s^{-1} or k = 6.16 x 10^{-4} s^{-1} or average, k = 6.28 x 10^{-4} s^{-1} {rate = $k[\text{ester}]^1 = k[CH_3CH_2COOCH_2CH_3]^1$; use the initial rate or 600 sec rate just calculated; initial rate: 6.0 x 10^{-4} M/s = k(1.0); k = 6.0 x 10^{-4} s^{-1}; 600sec rate: 4.25 x 10^{-4} M/s = k(0.69); k = 6.16 x 10^{-4} s^{-1}; can also avg these to yield, k = 6.15 x 10^{-4} s^{-1}}

f. 1083 sec or 1125 sec, or 1104 sec (avg = 1104) {$t_{1/2} = 0.693/k$; 1083sec for k = 6.40 x 10^{-4}; 1125sec for k = 6.16 x 10^{-4};or 1104sec for average k = 6.28 x10^{-4}}

g. 0.125M {1.00 → 0.500 → 0.250 → 0.125M}

7. c {rate = $(1)(1)^2(1)^0(1) = 1$; now double the concentrations: rate = $(1)(2)^2(2)^0(2) = 8$}

8. a. Rate = $k[NO]^2[O_2]$ {use Exp 1 & 3 to determine order of O_2; use 1 & 2 for order of NO}

b. third order

c. 7170$M^{-2}s^{-1}$ {use any experiment; Experiment 1: 2.8 x 10^{-2} = k(1.25 x $10^{-2})^2$(2.5 x 10^{-2}); solve for k; k = 7168$M^{-2}s^{-1}$}

d. 8.96 x 10^{-1}M/s {rate = $k[NO]^2[O_2]$ = (7170)(5 x $10^{-2})^2$(5 x 10^{-2}) = 8.96 x 10^{-1}M/s}

9. a. 1st order {$\text{rate}_a/\text{rate}_b = ([NH_4^+]_a/[NH_4^+]_b)^x$; 5.4 x 10^{-7}/2.7 x 10^{-7} = $(0.2/0.1)^x$; x = 1}

b. 1st order {$\text{rate}_a/\text{rate}_b = ([NO_2^-]_a/[NO_2^-]_b)^y$; 2.7 x 10^{-7}/1.35 x 10^{-7} = $(0.01/0.005)^y$; y = 1}

c. rate = $k[NH_4^+][NO_2^-]$

d. 2.7 x 10^{-4} 1/(Ms) {rate = $k[NH_4^+][NO_2^-]$; 1.35 x 10^{-7} = k(0.1)(0.005); k = 2.7 x 10^{-4} 1/(Ms)}

e. 6.1 x 10^{-7} M/s {rate = $k[NH_4^+][NO_2^-]$; rate = 2.7 x 10^{-4}(0.15)(0.015); rate = 6.075 x 10^{-7} M/s}

10. a. 1st order {ln[] vs t yielding a line indicates first order}

b. 1.55 x 10^{-4} min^{-1} {slope = -k for zero order so k = -slope = -(-1.55 x 10^{-4}) = 1.55 x $10^{-4}min^{-1}$}

c. 4470min {$t_{1/2} = 0.693/(1.55 \times 10^{-4})$ = 4471 min}

d. 2.72 x 10^{-2}M {$\ln[A] = -1.55 \times 10^{-4}(60) + \ln(2.75 \times 10^{-2})$}

11. 0.0174 M {$\ln[SO_2Cl_2] = -(2.2 \times 10^{-5})$(4.5hrs x 60min/1hr x 60s/1min) + ln(0.0248); $\ln[SO_2Cl_2]$ = -4.053; $[SO_2Cl_2]$ = 0.0174M}

12. 244yr {k = 0.693/1602yr = 4.326 x 10^{-4}/yr; $\ln(A_t/A_o) = -kt$ → (A_t/A_o) x 100% = %remaining so A_t/A_o = 0.90; solve for t; ln(0.9) = -(4.326 x 10^{-4}/yr)t; t = 243.6yr}

13. 3020 years {k = 0.693/5730 = 1.209 x 10^{-4}/yr; $\ln(A_t/A_o) = -kt$ → solve for t; ln(8.75/12.6) = -(1.209 x 10^{-4})t; t = 3016 years}

14. b {The rate can increase also by increasing concentrations of substances in the rate law.}

15. 3.5 x 10^4s {slope = -k → k = 2.0 x 10^{-5}; use $t_{1/2} = 0.693/k$ to find $t_{1/2}$}

16. a. 1.25 x 10^{-4} M {5 x $10^{-4}(^1/_2)^2$ = 1.25 x 10^{-4}M} b. 13.3s {$t_{1/2} = 1/k[A]_o = 1/[150(5 \times 10^{-4})]$ = 13.33}

17. a. rate = k[A] {since the plot of ln[A] vs t yields a line → 1st order}

b. 0.010 s^{-1} {slope = -k; slope = $\Delta y/\Delta x = [-3 - (-5)]/[0 - 200] = -0.01$; k = 0.010}

c. 0.050M {$\ln[A]_0 = -3$; $[A]_0$ = 0.0498M}

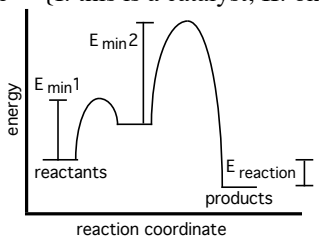
d. 0.015M {$\ln[A]_t = -kt + \ln[A]_0 = -(0.01)(120) + (-3) = -4.2$; $[A]_t = e^{-4.2}$ = 0.0149M}

18. d {IV: catalyzed reaction has lower activation energy.}

19. a. NO {first appears as a reactant, then a product; does not appear in the overall reaction} b. NO_2 and O

c. O_3 d. rate = $k_1[NO][O_2]^{1/2}$ {catalysts can be in a rate law; only intermediates can not be part of the rate law}

20. c {I: this is a catalyst; II: only T changes k; IV: doesn't change molecular orientation}



21.

22. multi-step mechanism; since an intermediate is formed there must be at least 2 steps; for instance: A + B → AB followed by AB → more chemistry not described in the problem

23. b {a: rate = $k_1[NO]^2$; c: rate = $k[NO]^2$; d: rate = $k_1[O_2]^2$}

24. c {III: intermediates are not part of collisional theory}

25. e {I: T increases → rate and k increase but E_a unchanged; II: Slope = $-E_a/R$; III: If reactant is 0 order it won't affect the rate.}

26. a {$\ln(k_2/k_1) = -E_a/R(1/T_2 - 1/T_1)$; $\ln(12000/1100) = -[E_a/(8.314)][1/333 - 1/298]$; $2.390 = -[E_a/8.314][-3.527 \times 10^{-4}]$; $E_a = 56{,}338 = 5.6 \times 10^4$ J/mol}

27. 475 K {use Arrhenius equation: $\ln(k_2/k_1) = -(E_a/R)(^1/T_2 - {}^1/T_1)$ with R = 8.314; for reaction to double with everything constant except T means that the new k must be twice as large → $k_2 = 2.0 \times 10^{-2}$; $\ln(2.0 \times 10^{-2}/1.0 \times 10^{-2}) = -(1 \times 10^4/8.314)(^1/T_2 - {}^1/373)$ → solve for T_2)

28. a {$\ln(k_2/k_1) = -(E_a/R)(^1/T_2 - {}^1/T_1)$ with R = 8.314; $k_1 = 1.4 \times 10^{-2}$; $T_1 = 500$; $T_2 = 298$; $E_a = 80{,}000$J/mol; solve for k_2; $\ln(k_2/1.4 \times 10^{-2}) = -(80{,}000)/(8.314)[1/298 - 1/500]$; $\ln(k_2/1.4 \times 10^{-2}) = -13.04$; $k_2/1.4 \times 10^{-2} = 2.161 \times 10^{-6}$; $k_2 = 3.025 \times 10^{-8}$}

29. c {slope = $-500 = -E_a/R = -E_a/8.314$; $E_a = (500)(8.314) = 4157$J}

30. d {the rates are equal, not the rate constants}

31. d {the new reaction is the reverse and times 1/2 of the previous reaction; change K by taking reciprocal, 1/K, and by raising to the power of 1/2; $(1/50.3)^{1/2} = 1.4 \times 10^{-1}$}

32. $[K_2(K_3)^4]/K_1$ {reverse rxn I, K' = $1/K_1$; rxn II unchaged, K' = K_2; multiply rxn III by 4, K' = $(K_3)^4$; add rxns and multiply K's}

33. c {$K_p = K_c(RT)^{\Delta n}$ and since $\Delta n = 0$, $K_p = K_c(RT)^0$ or $K_p = K_c$}

34. 1.81×10^{-2} {$K_p = K_c\,(RT)^{\Delta n}$; R = 0.0821; T = 298, $\Delta n = 3 - 5 = -2$, $K_c = K_p/[(0.0821)(298)]^{-2} = (3.02 \times 10^{-5})(24.47)^2 = 1.81 \times 10^{-2}$}

35. 69 {$K = [PCl_5]/[PCl_3][Cl_2] = (0.018/5)/[(0.015/5)(0.087/5)] = 68.97$}

36. K = 3.80 {set up ICE table; I: 1M, 0.5M, 0, 0; C: -0.42, -0.42, +0.42, +0.42; E: 0.58M CO; 0.08M H_2O, 0.42M CO_2, 0.42M H_2; $K = [CO_2][H_2]/[CO][H_2O] = (0.42)(0.42)/(0.58)(0.08) = 3.80$}

37. a. [A] = 0.325 mol/l, [B] = 0.175 mol/l, [C] = 0.0875 mol/l {[A] = 1.5mol/3L = 0.5M; 35% decrease = (0.35)0.5) = 0.175; it'll be negative for A (decreasing); ICE table: I: 0.5M, 0, 0; C: -0.175, +0.175, +0.0875; E: 0.325, 0.175, 0.0875}

b. $K = 2.5 \times 10^{-2}$ {K = [B]^2[C]/[A]^2 = (0.175)^2(0.0875)/(0.325)^2 = 0.025}

38. 0.101 M N_2 and O_2; 1.01×10^{-2} M NO {ICE table; I: 0.106M, 0.106M, 0; C: -x, -x, +2x; E: 0.106-x, 0.106-x, 2x; $K = [NO]^2/[O_2][N_2] = (2x)^2/(0.106\text{-}x)(0.106\text{-}x) = (2x)^2/(0.106\text{-}x)^2 = 1 \times 10^{-2}$; take square root of both sides: $(2x)/(0.106\text{-}x) = 1 \times 10^{-1}$; $2x = 0.0106 - 0.1x$; $2.1x = 0.0106$; $x = 0.00504$; $[NO] = 2x = 2(0.00504) = 0.0101$M; $[N_2] = [O_2] = 0.106 - 0.005 = 0.101$M}

39. $[N_2] = 1.94$M, $[O_2] = 2.94$M, $[NO] = 0.122$M {do ICE table; $K = [NO]^2/[N_2][O_2] = (2x)^2/[(2\text{-}x)(3\text{-}x)] \approx (2x)^2/[(2)(3)] = 0.0025$; $4x^2 = 0.015$; $x = 0.06123$; $[N_2] = 2\text{-}0.06123 = 1.94$M; $[O_2] = 3\text{-}0.06123 = 2.94$M; $[NO] = 2(0.06123 = 0.122$M}

40. $[NO_2] = 5.94 \times 10^{-4}$ M; $[N_2O_4] = 3.00 \times 10^{-6}$M {set up ICE table; $[N_2O_4] = 0.0015mol/5L = 0.0003M$; I: 0.0003, 0; C: -x, +2x; E: 0.0003-x, 2x; $K = [NO_2]^2/[N_2O_4]$; $0.125 = (2x)^2/(0.0003-x)$; $4x^2 + 0.125x - 3.75 \times 10^{-5} = 0$; use quadratic to solve for x; $x = 0.000297$; $[N_2O_4] = 0.0003-0.000297 = 3.00 \times 10^{-6}$; $[NO_2] = 2(0.000297) = 5.94 \times 10^{-4}$}

41. No, left {$Q = 1.9 \times 10^8$; Q > K so reaction goes to the left}

42. a. Q = 18 {$Q = [NH_3]^2/[N_2][H_2]^3 = (1.5)^2/(1)(0.5)^3$}

b. $[N_2]$ and $[H_2]$ will increase while the $[NH_3]$ will decrease {since Q > K, reaction goes left}