Diebolt CHM 152 Exam I Key Fall 17(100 pts)

Part 1. Multiple Choice. (40 points; 4 points each)

1. Consider the reaction $2Cr(OH)_3 + BrO_3^- + 4OH^- \rightarrow Br^- + 2CrO_4^{2-} + 5H_2O$. If the rate of disappearance of $Cr(OH)_3$ is 3.0×10^{-3} M/s, what is the rate of appearance of H_2O?
   A. 1.2×10^{-3} M/s B. 1.5×10^{-3} M/s C. 3.0×10^{-3} M/s D. 1.5×10^{-2} M/s **E. 7.5 x 10^{-3} M/s**

   $\frac{3.0\times10^{-3}\text{ mol Cr(OH)}_3}{\text{L}\cdot\text{s}} \times \frac{5\text{ mol H}_2\text{O}}{2\text{ mol Cr(OH)}_3}$ = 7.5×10^{-3} M/s H_2O

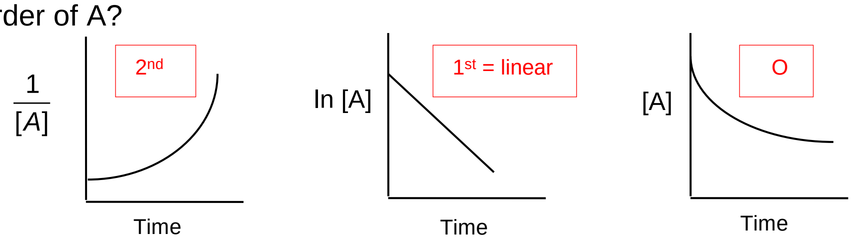
2. The rate law for a given reaction is **Rate = k[A][B]^3**. If the concentration of A is quadrupled and the concentration of B is tripled, the reaction rate would increase by a factor of ___.
   A. 12 B. 27 C. 36 **D. 108** E. 192

   Rate = k[4]^1[3]^3 = k×4×27 = k×108

3. The half-life is 115 minutes for a first order reaction. The rate constant, k, in s^{-1}, is
   A. 1.00 × 10^{-4} s^{-1} B. 6.03×10^{-3} s^{-1} C. 0.362 s^{-1} D. 4.78×10^{3} s^{-1} E. 9.96×10^{3} s^{-1}

   $115\,\text{min}\times\frac{60\,\text{s}}{1\,\text{min}} = 6900\,\text{s}$ $k = \frac{0.693}{6900\text{ s}} = 1.00\times10^{-4}\ \text{s}^{-1}$

4. The following graphs were prepared from experimental data for a reactant, A. What is the correct order of A?



   A. zero order **B. first order** C. second order D. insufficient information provided

5. A mechanism for a naturally occurring reaction that destroys ozone is:
   Step 1: $O_3(g) + HO(g) \rightarrow HO_2(g) + O_2(g)$
   Step 2: $HO_2(g) + O(g) \rightarrow HO(g) + O_2(g)$
   Which statement is **not true** for this mechanism
   A. There is not enough information to determine which step is slower. **T**
   B. The overall reaction is: $O_3(g) + O(g) \rightarrow 2\,O_2(g)$ **T**
   C. The overall rate law must be rate = k[O$_3$][O]. F
   D. Both steps are bimolecular. **T**
   E. HO is a catalyst and and HO_2 is an intermediate. **T**

6. For the following reaction, $2H_2(g) + O_2(g) \leftrightarrows 2H_2O(g)$, $K_c = 2.5 \times 10^4$.
   What is K_c for the reaction $2H_2O(g) \leftrightarrows 2H_2(g) + O_2(g)$, K_c = ?
   A. 4.0×10^{-5} B. -2.5×10^4 C. 2.5×10^4 D. 1.6×10^2 E. 6.3×10^{-3}

   Reaction is reversed, so K is inverse: K = $\frac{1}{2.5\times10^4} = 4.0\times10^{-5}$

7. The rate law for the reaction $NO_2 + CO \rightarrow NO + CO_2$ is **Rate = k[NO$_2$]^2**. Which one of the following mechanisms is consistent with this experimental rate law?

   A. $NO_2 + CO \rightarrow N + CO_2$ *slow*
      $N + NO_2 \rightarrow NO$ *fast*

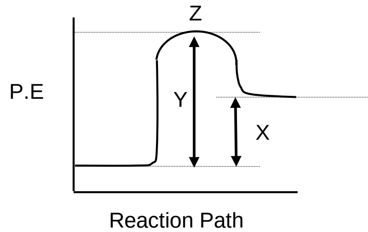
   B. $NO_2 + 2CO \rightarrow N + 2CO_2$ *slow*
      $N + NO_2 \rightarrow 2NO$ *fast*

   C. $NO_2 + NO_2 \rightarrow NO_3 + NO$ *fast*
      $NO_3 + CO \rightarrow NO_2 + CO_2$ *slow*

   D. $NO_2 + NO_2 \rightarrow NO_3 + NO$ *slow*
      $NO_3 + CO \rightarrow NO_2 + CO_2$ *fast*

   Rate = k[NO$_2$]^2 so 2 NO$_2$ molecules are reactants in slow step

8. For the following diagram, which statement is not correct?



A. X is the heat of reaction. **T**
B. **The products are lower in energy than the reactants.** **F - the product level is higher**
C. Z is the transition state. **T**
D. Y is the activation energy for the reaction. **T**
E. The reaction is endothermic. **T**

9. All the following statements are true **EXCEPT**
A. in a series of stepwise reactions, the rate-determining step is the slow one. T
B. the rate constant for a reaction changes when temperature is changed. T
C. **a catalyst increases the rate of reaction by decreasing the heat of reaction, ΔH.** **F- catalyst lowers E_a but ΔH stays the same**
D. the rates of most chemical reactions change with time. T
E. the rate constant does not depend on the reactant concentrations. T; **Rate depends on concentration but the rate constant, k, is a constant – k only varies with T!**

10. Arrange the following reactions in order of increasing tendency to go to completion.

1) $2NOCl(g) \leftrightarrows 2NO(g) + Cl_2(g)$	$K_p = 1.7 \times 10^{-2}$ second largest
2) $N_2O_4(g) \leftrightarrows 2NO_2(g)$	$K_p = 1.5 \times 10^{3}$ largest
3) $2SO_3(g) \leftrightarrows 2SO_2(g) + O_2(g)$	$K_p = 1.3 \times 10^{-5}$ smallest (1.3 is smaller than 5.9!)
4) $2NO_2(g) \leftrightarrows 2NO(g) + O_2(g)$	$K_p = 5.9 \times 10^{-5}$ second smallest

From **least to most** complete:

A. 2 < 1 < 4 < 3 B. 3 < 1 < 4 < 2 C. 4 < 3 < 1 < 2 D. 4 < 3 < 2 < 1 **E. 3 < 4 < 1 < 2**

Part 2. Short answer (18 points)

1. Consider the following equilibrium reaction:

Heat + $CO_2(g) + 2H_2O(g) + 4Cu(s) \leftrightarrows 4CuO(s) + CH_4(g)$ **ΔH = +559 KJ = endo**

a) Write the K_c expression for this equilibrium reaction. (3 pts)

Omit solids and liquids! $K_c = \frac{[CH_4]}{[CO_2][H_2O]^2}$

For b-e, predict the effect of the following changes on the equilibrium position (15 pts):

b) Adding CO_2 (left, right, no change) **Right**

c) Increasing the volume (left, right, no change) **left**

d) Removing CH_4 (left, right, no change) **Right**

e) Decreasing the temperature (left, right, no change) **left**

f) What happens to the value of K_c (increase, decrease or stay the same), if we decrease Consider the following equilibrium reaction: **decreases**

Part Three. Numerical Problems. You must **SHOW YOUR WORK** to receive full credit! Make sure you circle your **final answer**, and express your final answer with the proper number of sig figs and the proper units! (42 points)

1. Given the initial reaction rate data for the following reaction: $3A + B \rightarrow C$

Trial	[A] (M)	[B] (M)	Rate $\left(\frac{M}{s}\right)$
1	0.200	0.250	1.1×10^{-4}
2	0.800	0.250	4.4×10^{-4}
3	0.200	0.750	9.9×10^{-4}

a) What trials should be used to find the order for A? (1 pt) 1 & 2

Pick trials where A changes, B same:

b) Order for A = 1. (2 pts)

$$\left(\frac{0.800}{0.200}\right)^x = \left(\frac{4.4\times10^{-4}}{1.1\times10^{-4}}\right); \; 4^x = 4 \; ; \; x = 1$$

c) What trials should be used to find the order for B? (1 pt) 1 & 3

Pick trials where B changes, A same:

d) Order for B = 2. (2 pts)

$$\left(\frac{0.750}{0.250}\right)^y = \left(\frac{9.9\times10^{-4}}{1.1\times10^{-4}}\right); \; 3^y = 9; \; y = 2$$

e) Write the rate law for this reaction. Make sure to use the orders found above! (2 pts)

Rate = k[A][B]^2

d) Calculate the rate constant, k, and include the proper units. Must show set-up! (4 pts)

$$k = \frac{Rate}{[A][B]^2} = \frac{1.1x10^{-4}\ M/s}{(0.200\ M)(0.250\ M)^2}$$

$k = 8.8x10^{-3}\ M^{-2}s^{-1}$

2. The decomposition of hydrogen peroxide is described by the equation: $2H_2O_2 \rightarrow 2H_2O + O_2$. The reaction is first order in H_2O_2 and the rate constant is $1.8 \times 10^{-5}\ s^{-1}$ at a certain temperature. The initial concentration of H_2O_2 is 1.45 M. What will the concentration of H_2O_2 be after 58 hours? (8 pts)

$[H_2O_2]_0 = 1.45$ M $\quad k = 1.8 \times 10^{-5}\ s^{-1} \quad t = 58\text{ hours}\left(\frac{60\ min}{1\,hr}\right)\left(\frac{60\ s}{1\ min}\right) = 2.088 \times 10^5$ s

$$\ln\frac{[A]_t}{[A]_0} = -kt \quad \Rightarrow \quad \ln\frac{[H_2O_2]_t}{1.45\ M} = -1.8 \times 10^{-5}\ s^{-1}\ (2.088 \times 10^5\ s)$$

$$\ln\frac{[H_2O_2]_t}{1.45\ M} = -3.7584 \quad \Rightarrow \quad \frac{[H_2O_2]_t}{1.45\ M} = e^{-3.7584} = 2.332x10^{-2}$$

$[H_2O_2]_t = (1.45)(2.332x10^{-2}) =$ **$3.4x10^{-2}$ M**

or $n = 58\ h/10.69\ h = 5.42 \Rightarrow$ **$[A]_t = (0.5)^n[A]_0 = (0.5)^{5.42}\cdot(1.45\ M) = 3.4x10^{-2}$ M**

3. A mixture of 0.415 M $I_2(g)$ and 0.415 M $Br_2(g)$ is placed in a container and undergoes the following reaction. Calculate the concentrations of I_2, Br_2, and IBr after the system has reached equilibrium. (10 pts)

$I_2(g) + Br_2(g) \leftrightarrows 2IBr(g)$ $K_c = 121$

	$I_2(g)$ +	$Br_2(g)$ $\leftrightarrows$	$2IBr(g)$
Initial, M	0.415	0.415	0
Change, M	-x	-x	+2x
Equilibrium, M	0.415 - x	0.415 - x	2x

$K_c = 121$

$$K_c = \frac{[IBr]^2}{[I_2][Br_2]} \qquad 121 = \frac{(2x)^2}{(0.415-x)(0.415-x)} = \frac{(2x)^2}{(0.415-x)^2} \Leftarrow \text{perfect square}$$

Take square root both sides: $11 = \frac{2x}{(0.415-x)}$ $\Rightarrow 11(0.415 - x) = 2x \Rightarrow$

$4.565 - 11x = 2x \Rightarrow 4.565 = 13x$ thus, $x = \frac{4.565}{13} = 0.351$ (3 sf)

Equilibrium Concs: $[I_2] = [Br_2] = 0.415 - 0.351 =$ **0.064 M**; $[IBr] = 2(0.351) =$ **0.702 M**

$[I_2] = [Br_2] = 0.064$ M $[IBr] = 0.702$ M

4. For the following reaction, $K_c = 55$ at 280 °C.

$N_2(g) + 3 H_2(g) \leftrightarrows 2 NH_3(g)$

a) Calculate Q when 0.50 moles of N_2, 1.0 moles of H_2 and 4.0 moles of NH_3 are placed in a 2.0 L flask: (5 pts)

$[N_2] = \frac{0.50\,\text{mol}}{2.0\,\text{L}} = 0.25$ M; $[H_2] = \frac{1.0\,\text{mol}}{2.0\,\text{L}} = 0.50$ M; $[NH_3] = \frac{5.0\,\text{mol}}{2.0\,\text{L}} = 2.0$ M

$Q = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(2.0)^2}{(0.25)(0.5)^3} = \frac{4}{0.03125}$ **= 128** **Q = 130**

b) Comparing your calculated Q to K_c, which statement is true? (Circle one) (2 pts)

i. the reaction is at equilibrium

ii. the reaction will shift right to attain equilibrium

iii. the reaction will shift left to attain equilibrium

Q is larger than K so there are too many products present

c) What is the value of K_p at 280 °C? (5 pts)

$K_p = 55(0.08206*553)^{-2} = 0.0267$ **$K_p = 0.027$**

	pts possible
Multiple choice	40
Page 2	18
Page 3	20
Page 4	22
Total Pts	100

5 pts extra credit added for review from 1st week of class