Exam 2A
Chem 1142
Spring 2017

Name: KEY

MULTIPLE CHOICE. [3 pts ea.] Record the best response on the scantron sheet. [**45 pts total.**]

Q1. Which version of the exam do you have?

a) 2A
b) 2B

Q2. Which of the following expressions is equal to the **rate of reaction** for the following chemical equation?

$2A + B \rightarrow 3C + D$

a) $-\frac{1}{2}\Delta[A]/\Delta t$
b) $+\Delta[B]/\Delta t$
c) $3\Delta[C]/\Delta t$
d) $-\Delta[D]/\Delta t$

Q3. For the overall stoichiometric equation: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, the rate law can be written as:

a) rate = $k[N_2][H_2]$
b) rate = $k[N_2][H_2]^3$
c) rate = $k[N_2][H_2]^2/[NH_3]^2$
d) impossible to say for sure
e) zeroth order with respect to nitrogen

Q4. The units of a rate constant for a second order overall reaction are:

a) M/s
b) s^{-1}
c) $M^{-1}cm^{-1}$
d) $M^{-1}s^{-1}$
e) M^2

$$\underbrace{rate}_{M \cdot s^{-1}} = \overset{M^{-1}s^{-1}}{K}\underbrace{[A]^2}_{M^2}$$

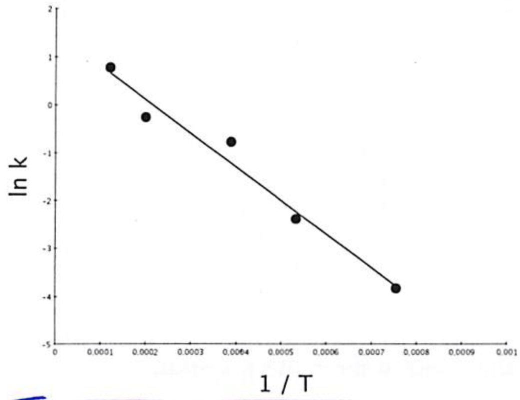
Q5. The rate constant for a reaction with the rate law: *rate* = k[A], is found to have a value of 0.012 s^{-1} at a temperature of 17 °C. How long will it take for the concentration of A to change from 1.2 M to 0.60 M?

$\times \frac{1}{2}$

a) 58 s
b) 50. s
c) 1.0×10^2 s
d) 83 s
e) 85 s

$t_{\frac{1}{2}} = \frac{0.693}{K} = 57.75 s$

Q6. What information can be determined from the following type of graph:



a) The activation energy of the reaction
b) Whether the reaction is zeroth order
c) Whether the reaction is first order
d) Whether the reaction is second order
e) Whether the reaction is an elementary reaction

Q7. The half-life for a first order reaction is 15 s. What fraction of **reactants** will remain after 30. s have passed?
a) 75 %
b) 50 %
c) 25 %
d) 0 %
e) Not enough information to determine

Q8. What type of plot can be used to determine whether a reaction involving a reactant, A, is second order?
a) ln k vs. $1 / T$
b) [A] vs t
c) ln [A] vs. t
d) 1 / [A] vs. t
e) k vs. t

Q9. A reaction is thought to proceed via a mechanism:

$A + B \rightarrow C$	k_1 SLOW
$C + B \rightarrow D$	k_2 FAST

the reaction's **rate law** will be given by the expression:

a) $rate = k_1[A][B]$
b) $rate = k_1[C]/[A][B]$
c) $rate = k_2[B][C]$
d) $rate = k_2[D]/[C][B]$
e) $rate = k_2[A][B]^2[C]$

Q10. A chemical reaction has an equilibrium constant of 4,000 at 25 °C. This means that at equilibrium:
a) There will mainly be reactants present
b) There will mainly be products present
c) There will be similar amounts of reactants and products

Q11. Which of the following chemical equations will have the same value for both K_c and K_p

$K_p = K_c (RT)^{\Delta n_g}$

if $\Delta n_g = 0$, $K_p = K_c$

a) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta n_g = -2$
b) $CH_4(g) + 2O_2(g) \rightleftharpoons CO_2(g) + 2H_2O(g)$ 0
c) $2C(s) + O_2(g) \rightleftharpoons 2CO(g)$ $+1$
d) $N_2O_4(g) \rightleftharpoons 2NO_2(g)$ $+1$
e) $2BH_3(g) \rightleftharpoons B_2H_6(g)$ -1

Q12. A chemical equilibrium: A $\rightleftharpoons$ B has an equilibrium constant of 12.4 at 25 °C. If we re-write the chemical equilibrium as: 2B $\rightleftharpoons$ 2A, what will the equilibrium constant for this chemical equilibrium become?

reverse and double

$12.4 \rightarrow \frac{1}{12.4} \rightarrow \left(\frac{1}{12.4}\right)^2 = 6.5 \times 10^{-3}$

a) 12.4
b) 6.2
c) 2.48×10^2
d) 6.5×10^{-3}
e) 4.0×10^{-2}

Q13. A chemical equilibrium has a reaction quotient of 34 and an equilibrium constant of 24. This information tells us that:

$Q_c > K_c$

$Q_c \sim \frac{P \downarrow}{R \uparrow}$

a) The reaction is already at equilibrium
b) The reaction will form more products
c) The reaction will form more reactants
d) Unable to tell anything from this information
e) The reaction will reach equilibrium in 10. s

Q14. A small amount of NaOH is added to the following equilibrium, resulting in neutralization of some of the H^+ ions. How will this affect the position of equilibrium?

lower conc

$$\underset{\longleftarrow}{H^+(aq) + HCO_3^-(aq) \rightleftharpoons H_2CO_3(aq)}$$

Le Chatelier: shift to make more H^+

a) No affect
b) There will be a shift to the left-hand-side (LHS)
c) There will be a shift to the right-hand-side (RHS)
d) There will be an increase in the value of the equilibrium constant
e) There will be a decrease in the value of the equilibrium constant

Q15. Consider the following chemical equilibrium:

$$2A(aq) + B(s) \rightleftharpoons 3C(g) \; \text{+heat}; \Delta H^\circ = -34.0 \text{ kJ/mol}$$ exothermic

What will happen if the temperature of the equilibrium is raised?

$T \uparrow$, "add heat".
⇒ shift to LHS to "remove heat"

$K = \frac{P \downarrow}{R \uparrow} \Rightarrow K \downarrow$

a) No affect
b) The equilibrium constant will increase
c) The equilibrium constant will decrease

Short Response.

Show ALL work to receive credit.

Q16. [11 pts.] The chemical reaction between hydrazine and dinitrogen tetroxide was initially investigated for travel to the moon in the 1960s. The chemical equation for this reaction is given below:

$$2N_2H_4 + N_2O_4 \rightarrow 3N_2 + 4H_2O$$

a) Write out four expressions for the reaction rate (not the rate law!) below:

i) $-\frac{1}{2}\frac{\Delta[N_2H_4]}{\Delta t}$	ii) $-\frac{\Delta[N_2O_4]}{\Delta t}$
iii) $+\frac{1}{3}\frac{\Delta[N_2]}{\Delta t}$	iv) $+\frac{1}{4}\frac{\Delta[H_2O]}{\Delta t}$

b) If a set of three experiments was carried out measuring how the initial rate depended upon the initial concentration of each reactant, determine the rate law for this reaction. As part of your answer determine the value of the rate constant, *k*. <u>SHOW ALL WORK.</u>

Experiment #	$[N_2H_4]_0$ / M	$[N_2O_4]_0$ / M	Initial rate / $M{\cdot}s^{-1}$
1	0.010	0.010	1.80
2	0.020	0.010	3.60
3	0.010	0.025	11.25

$\text{rate} = k\,[N_2H_4]^x[N_2O_4]^y$

$$\frac{\text{rate}(2)}{\text{rate}(1)} = \frac{3.60\,\cancel{M{\cdot}s^{-1}}}{1.80\,\cancel{M{\cdot}s^{-1}}} = \frac{\cancel{k}[0.020M]^x[\cancel{0.010M}]^y}{\cancel{k}[0.010M]^x[\cancel{0.010M}]^y} \Rightarrow 2.00 = 2.0^x$$

$\Rightarrow x = 1$ (by inspection)

or $\log(2.00) = \log(2.0^x) = x\cdot\log(2.0)$

$\Rightarrow x = \log(2.00)/\log(2.0) = 1$

$$\frac{\text{rate}(3)}{\text{rate}(1)} = \frac{11.25\,\cancel{M{\cdot}s^{-1}}}{1.80\,\cancel{M{\cdot}s^{-1}}} = \frac{\cancel{k}[\cancel{0.010M}]^x[0.025M]^y}{\cancel{k}[\cancel{0.010M}]^x[0.010M]^y}$$

$\Rightarrow 6.25 = 2.5^y \Rightarrow y = 2$ (by inspection)

<u>OR</u> $\log(6.25) = \log(2.5^y) = y\cdot\log(2.5)$

$\Rightarrow y = \log(6.25)/\log(2.5) = 2$

$\boxed{\text{rate} = k\,[N_2H_4][N_2O_4]^2}$

$$k = \frac{\text{rate}}{[N_2H_4][N_2O_4]^2} = \frac{1.80\,M{\cdot}s^{-1}}{0.010M \times (0.010M)^2} = 1.8\times10^6\ M^{-2}s^{-1}$$

Q17. [11 pts.] The following chemical equation:

$$A \rightarrow B$$

has a rate law of:

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

with a rate constant equal to 0.098 $M^{-1} \cdot s^{-1}$ at 310 K.

$-\frac{1}{2}$ units, sf.

a) What is the half-life for this reaction? The initial concentration of A is 2.0 M.

$$t_{\frac{1}{2}} = \frac{1}{[A]_0 K} = \frac{1}{2.0M \times 0.098 M^{-1}s^{-1}} = 5.1s$$

$\boxed{+3}$

1 1 1

b) How long will it take for the concentration of A to change from 2.0 M to 0.25 M?

$$\frac{1}{[A]_t} = Kt + \frac{1}{[A]_0} \Rightarrow \frac{1}{[A]_t} - \frac{1}{[A]_0} = Kt \Rightarrow t = \frac{\frac{1}{[A]_t} - \frac{1}{[A]_0}}{K}$$

1 3

$$\Rightarrow t = \frac{\frac{1}{0.25M} - \frac{1}{2.0M}}{0.098M^{-1}s^{-1}}$$

1

$$= 36s$$

1

$\boxed{+6}$

(−3pts if use $t_{\frac{1}{2}}$ in (a) + say: 15.3s)

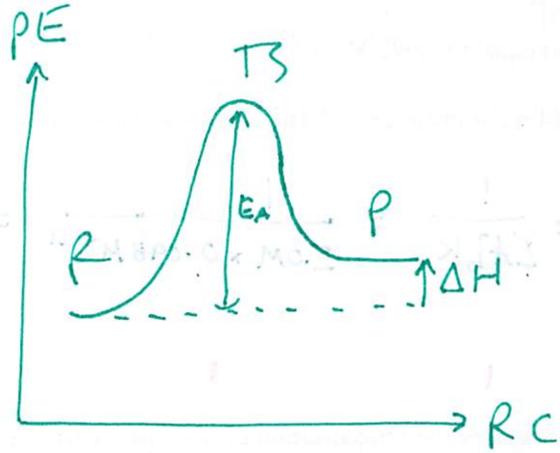
c) If you measured how the concentration of A varied with time, what kind of plot would be expected to give a straight line?

$1/[A]$ vs. t

$\boxed{+2}$

11

Q18. [11 pts.] a) Sketch the potential energy vs. reaction coordinate diagram for an endothermic reaction. Label the axes and carefully indicate on your graph the following pieces of information: ΔH, the activation energy, the transition state, reactants, and products.



+6 pts

-1 pt each error.

302K 309K

b) The rate constant for this reaction doubles as the temperature changes from 29 °C to 36 °C. What is the activation energy of the reaction?

$$\ln\left(\frac{k_2}{k_1}\right) = \frac{E_A}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right) \quad \Big\} +1$$

+5 pts.

$$E_A = \frac{R \cdot \ln\left(\frac{k_2}{k_1}\right)}{\left(\frac{1}{T_1} - \frac{1}{T_2}\right)} \quad \Big\} +1$$

$$= \frac{8.3145\ \mathrm{J/mol\cdot K} \times \ln(2)}{\left(\frac{1}{302\,\mathrm{K}} - \frac{1}{309\,\mathrm{K}}\right)} \quad \Big\} +2$$

$$= 76{,}800\ \mathrm{J/mol} \quad \text{or} \quad 76.8\ \mathrm{kJ/mol}. \quad \Big\} +1$$

$-\frac{1}{2}$ units, sf, etc.

Q19. [11 pts.] Given the chemical equilibrium:

$PbCl^{+}(aq) + 2Cl^{-}(aq) \rightleftharpoons PbCl_3^{-}(aq)$; $K_c = 0.45$ (25 °C)

-1.5 if do not square [Cl⁻]

-1 errors

a) Write out the expression for the equilibrium constant, K_c:

$$K_c = \frac{[PbCl_3^-]}{[PbCl^+][Cl^-]^2_{eq}}$$

+3

b) Show how to, and then calculate the reaction quotient if the initial concentrations are:
$[PbCl^+]_0 = 0.10$ M, $[Cl^-]_0 = 1.5$ M, $[PbCl_3^-]_0 = 0.20$ M

$$Q_c = \underbrace{\frac{[PbCl_3^-]}{[PbCl^+][Cl^-]^2_i}}_{+1} = \underbrace{\frac{0.20}{0.10 \times 1.5^2}}_{+1} = \underbrace{0.89}_{+1}$$

+3

c) What will happen to the value of the reaction quotient as time passes?

$\underbrace{Q_c \text{ will decrease}}_{+1}$, $\underbrace{\text{until it equals } K_c}_{+1}$

+2

d) Will the concentration of chloride ion increase , decrease, or stay the same over time? Explain.

$\overbrace{\text{It will increase}}^{+1}$, $\overbrace{\text{since more reactants have to be made to reach equilibrium.}}^{+2}$

$$\downarrow Q_c \sim \frac{P\downarrow}{R\uparrow}$$

+3

Q20. [11 pts.] 0.25 moles of Cl_2, H_2, and HCl are introduced into a 5.0-L flask. Given the chemical equation:

$$Cl_2(g) + H_2(g) \rightleftharpoons 2HCl(g) \qquad ; K_c = 140$$

Calculate the equilibrium concentrations of Cl_2, H_2, and HCl. Be sure to show **all** work.

$[Cl_2]_0 = [H_2]_0 = [HCl]_0 = \frac{0.25 mol}{5.0 L} = 0.050M$

$Q_c = \frac{[HCl]^2}{[Cl_2][H_2]_0} = \frac{0.050^2}{0.050 \times 0.050} = 1.0$

$Q_c < K_c \Rightarrow$ shift to RHS!

	$Cl_2(g)$	$+ H_2(g)$	$\rightleftharpoons 2HCl(g)$
I	0.050	0.050	0.050
C	$-x$	$-x$	$+2x$
E	$(0.050-x)$	$(0.050-x)$	$(0.050+2x)$

$K_c = \frac{[HCl]^2}{[Cl_2][H_2]_{eq}} = 140 = \frac{(0.050+2x)^2}{(0.050-x)^2}$

perfect square! $\Rightarrow \sqrt{140} = 11.83 = \frac{0.050+2x}{(0.050-x)}$

$\Rightarrow 11.83(0.050-x) = 0.050+2x$

$\Rightarrow 0.5916 - 11.83x = 0.050+2x$

$\Rightarrow 13.83x = 0.5416$

$\Rightarrow x = \frac{0.5416}{13.83} = 0.03916$

$\Rightarrow [Cl_2]_{eq} = [H_2]_{eq} = 0.050 - x = 0.0108M$
$= 0.011M$ (3d.p.)

$[HCl]_{eq} = 0.050 + 2x$
$= 0.128M$

BONUS Question:

What is the name given to the molecularity of an elementary reaction where the following number of molecules collide?

+1 ea (all/nothing)

one = unimolecular; two = bimolecular;

three = termolecular.

Useful Information

Periodic Table of the Elements

IA 1	IIA 2	3	4	5	6	7	8	9	10	11	12	IIIA 13	IVA 14	VA 15	VIA 16	VIIA 17	VIIIA 18
1 H 1.01																	2 He 4.00
3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.61	33 As 74.92160	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc [98]	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba* 137.33	71 Lu 174.97	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.20	83 Bi 208.98	84 Po [210]	85 At [210]	86 Rn [222]
87 Fr [223]	88 Ra** [226]	103 Lr [262]	104 Rf [261]	105 Db [262]	106 Sg [266]	107 Bh [264]	108 Hs [265]	109 Mt [268]	110 [269]	111 [272]	112 [277]	113	114 [285]	115	116 [289]	117	118 [293]

*	57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm [145]	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04
**	89 Ac [227]	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np [237]	94 Pu [244]	95 Am [243]	96 Cm [247]	97 Bk [247]	98 Cf [251]	99 Es [252]	100 Fm [257]	101 Md [258]	102 No [259]

$$R = 8.3145\frac{\text{J}}{\text{mol}\cdot\text{K}} = 0.08206\frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}}$$

$$k = Ae^{-\frac{E_A}{RT}} \qquad \ln k = -\frac{E_A}{R}\cdot\frac{1}{T} + \ln A \qquad \ln\left(\frac{k_2}{k_1}\right) = \frac{E_a}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

• 0-order: $[\text{A}]_t = -kt + [\text{A}]_0$ $\qquad t_{1/2} = \dfrac{[\text{A}]_0}{2k}$

• 1-order: $\ln[\text{A}]_t = -kt + \ln[\text{A}]_0$ $\qquad \ln\left(\dfrac{[A]_t}{[A]_0}\right) = -kt$ $\qquad t_{1/2} = \dfrac{0.693}{k}$

• 2-order: $\dfrac{1}{[\text{A}]_t} = kt + \dfrac{1}{[\text{A}]_0}$ $\qquad t_{1/2} = \dfrac{1}{[\text{A}]_0 k}$

$K_p = K_c(RT)^{\Delta n_g}$

Given: $ax^2 + bx + c = 0$, then $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$