

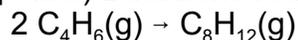
Name: \_\_\_\_\_

(4 points)

## Chemistry 114 Third Hour Exam

**Remember- Show all work for partial credit**

1 (12 points) Butadiene react to form dimers according to the equation:



The following data were collected for this reaction at 25°C:

[C <sub>4</sub> H <sub>6</sub> ] (mol/L)	Time (s)
.01	0
.00476	1800
.00313	3600
.00241	5200

A. If you thought this was a second order reaction, what would you plot (X and Y) to prove it was second order? How would you get the k of the reaction from this plot?

Plot  $1/[ ]$  as Y vs. t as X; the slope of the line = k

B. If you thought this was a first order reaction, what would you plot (X and Y) to prove it was first order? How would you get the k of the reaction from this plot?

Plot  $\ln[ ]$  as Y vs. t as X; the slope of the line = -k

C. If you thought this was a zero order reaction, what would you plot (on X and Y) to prove it was zero order? How would you get the k of the reaction from this plot?

Plot  $[ ]$  as Y vs. t as X; the slope of the line = -k

2. (12 points) I have a second order reaction that has a half-life of 20 minutes when my reactant concentration is 0.400 M.

A. What does the term half-life mean?

The half-time of the reaction is how long it takes the concentration of the reactant to decrease by 1/2.

B. What is the k of this reaction?

For a second order reaction  $t_{1/2} = 1/[A]_0 k$ ;  $20 \text{ min} = 1/0.4(X)$ ;  $X = 1/20(.40)$   
 $X = .125 \text{ l/mol}\cdot\text{min}$

C. If the initial reactant concentration is 0.400M, how long will it take until the reaction is 95% complete

if 95% complete, then only 5% of reactant remains

$[\text{reactant}]_{95\%} = 0.4(.05) = .02 \text{ Mol/l}$

For second order  $1/[ ]_t = 1/[ ]_0 + kt$

$$1/.02 = 1/.4 + .125X$$

$$50 = 2.5 + .125X$$

$$50 - 2.5 = .125X$$

$$47.5 = .125X$$

$$47.5/.125 = X; X = 380 \text{ min}$$

3. (12 points) In the collision model for reaction kinetics we relate the  $k$  of a reaction to several other parameters in the equation:

$$k = zpe^{-E_a/RT}$$

Briefly state what the collision model is, then tell me what each symbol in the above equation is and how it relates to the collision theory.

The collision theory states that the rate of a reaction ( $k$ ) is proportional to the number of collisions that take place times a steric factor, times a factor that accounts for the energy in a collision

$k$  = the rate of a reaction

$z$  = number of collisions

$p$  = orientation of collision or steric factor

$e$  = the exponential function

$E_a$  is the activation energy or minimum  $E$  needed for a reaction

$R = 8.314 \text{ J/K}\cdot\text{mol}$  (The gas constant)

$T$  = temperature in K

4. (12 points) Define the following terms:

Inhibitor - A material that slows the rate of a reaction.

Homogeneous equilibrium-

An equilibrium in which all substance are in the same phase.

Frequency factor

The factor  $A$  in the Arrhenius equation:  $k=Ae^{-E_a/RT}$ . With our current collision model of kinetics it is equal to the number of collisions times the steric factor

A bimolecular step

An elementary step in a reaction in which two compounds or elements must collide for a reaction to occur.

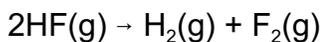
A pseudo-second-order reaction

A rate constant that appears to be second order because one or more other reactants were at such a high concentration that their concentrations did not change significantly over the course of the reaction, so they were essential constants in the reaction rate calculation.

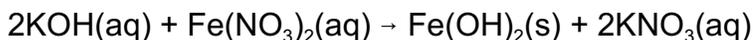
The Reaction quotient ( $Q$ )

The quotient you get when you plug your current set of conditions (that may not be at equilibrium) into the equilibrium expression.

5. (12 points) Write equilibrium expressions for the following reactions:



$$K = \frac{[\text{H}_2][\text{F}_2]}{[\text{HF}]^2}$$



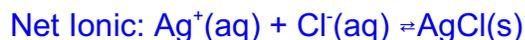
$$\text{If you use the above equation: } \frac{[\text{KNO}_3]^2}{[\text{KOH}]^2[\text{Fe}(\text{NO}_3)_2]}$$

Its actually more correct to go to the net ionic equation so you eliminate the spectator ions that don't contribute to the equilibrium



$$\text{Now } K = [\text{Fe}^{2+}][\text{OH}^-]^2$$

The precipitation reaction of silver nitrate with magnesium chloride.



$$K = [\text{Ag}^+][\text{Cl}^-]$$

6. (12 points) When the reaction:  $\text{N}_2(\text{g}) + 3\text{Cl}_2(\text{g}) \rightleftharpoons 2\text{NCl}_3(\text{g})$  is at equilibrium, the concentration of  $\text{NCl}_3$  is .19M, the concentration of  $\text{N}_2$  is .014M and the concentration of  $\text{Cl}_2$  is .0043M .

A What is the  $K_c$  for this reaction?

$$K_c = \frac{[\text{NCl}_3]^2}{[\text{N}_2][\text{Cl}_2]^3} = \frac{.19^2}{(.014)(.0043)^3} = 3.24 \times 10^7$$

B. Assuming the temperature is 350 K, what is  $K_p$  for this reaction?

$$K_p = K_c (\text{RT})^{\Delta n}; \Delta n = 2 - (3 + 1) = -2$$

$$K_p = 3.24 \times 10^7 (0.08206 \cdot 350)^{-2} = 3.92 \times 10^4$$

C. What is  $K_c$  for the reaction  $2\text{N}_2(\text{g}) + 6\text{Cl}_2(\text{g}) \rightleftharpoons 4\text{NCl}_3(\text{g})$ ?



$$K_{\text{new}} = K_{\text{old}}^2 = (3.24 \times 10^7)^2 = 1.05 \times 10^{15}$$

7. (12 points) The reaction  $2\text{NO}(\text{g}) \rightleftharpoons \text{N}_2(\text{g}) + \text{O}_2(\text{g})$  has a  $K_c$  of  $2.4 \times 10^3$ .

A. Which of the following systems are at equilibrium?

B. For any system that is not at equilibrium tell me if more products will form as it comes to equilibrium, or if product will disappear as it comes to equilibrium.

[NO]	[N <sub>2</sub> ]	[O <sub>2</sub> ]	At Equilibrium? (Y/N)	Product will <b>form</b> or <b>disappear</b> ?
.016	.310	2.00	Y	
.020	.800	.570	N	Form
.024	2.00	2.60	N	Disappear

(Use this space for calculations)

$$Q_1 = 2 \cdot .310 / (.016)^2 = 2,422 ; Q \approx K; \text{at equilibrium}$$

$$Q_2 = .570 \cdot .310 / (.02)^2 = 1,140 ; Q < K; \text{not enough products; products will form}$$

$$Q_3 = 2.00 \cdot 2.60 / (.024)^2 = 9028 ; Q > K; \text{too many products; products will disappear}$$

8. (12 points) At a particular temperature 8.0 mol of  $\text{SO}_3$  is placed in a 1.0 L container where the  $\text{SO}_3$  dissociates by the reaction



At equilibrium 2.0 mol of  $\text{SO}_2$  is present. Calculate K for this reaction.

	$2\text{SO}_3(\text{g})$	$\rightleftharpoons$	$2\text{SO}_2(\text{g})$	$+$	$\text{O}_2(\text{g})$
Initial	8		0		0
Change	?		2		?

From stoichiometry

$$2 \text{ mole } \text{SO}_2 \times (1 \text{ mole } \text{O}_2) / 2 \text{ mole } \text{SO}_2 = 1 \text{ mole } \text{O}_2$$

$$2 \text{ mole } \text{SO}_2 \times (2 \text{ mole } \text{SO}_3) / 2 \text{ mole } \text{SO}_2 = 2 \text{ mole } \text{SO}_3$$

	$2\text{SO}_3(\text{g})$	$\rightleftharpoons$	$2\text{SO}_2(\text{g})$	$+$	$\text{O}_2(\text{g})$
Initial	8		0		0
Change	-2		+2		+1
Equilibrium	6		2		1

$$K = 2^2 \cdot 1 / 6^2 = .111$$