

Name: _____

(4 points)

Chemistry 114 Third Hour Exam

Remember- Show all work for partial credit

All questions are worth 12 points

1. Define the following terms

A. Order (of a reaction)

In a rate equation the power or exponent associated with a concentration.

B. Elementary step

In a reaction mechanism, a reaction whose rate law can be written directly from the molecularity of the reaction.

C. Steric factor

A factor that tells how many collisions have the correct orientation for a reaction to occur.

D. Homogeneous catalyst

A substance that increases the rate of a reaction, but is not consumed in the reaction and is in the same phase as the product and reactants of the reaction.

E. Heterogeneous equilibria

An equilibrium that involves substances that are in different phases.

F. Reaction quotient (Q)

A calculation that is set up like an equilibrium expression, but the current concentrations of products and reactants are used rather than the equilibrium concentrations.

2. The reaction $A + B + C \rightarrow D$ is first order with respect to A and second order with respect to B and zero order with respect to C. Use that information to fill in the following table:

Rate (mole/sec)	[A](mole/l)	[B](mole/l)	[C](mole/l)
.0075	.005	.005	.005
.03 _____	.005	.01	.005
.015 _____	.01	.005	.005
.0075 _____	.005	.005	.01
.02058 _____	.007	.007	.007

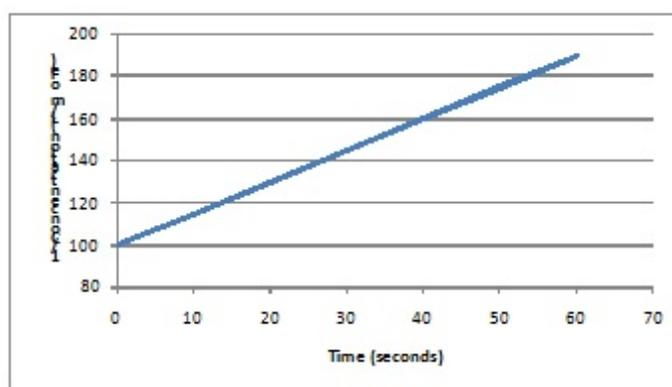
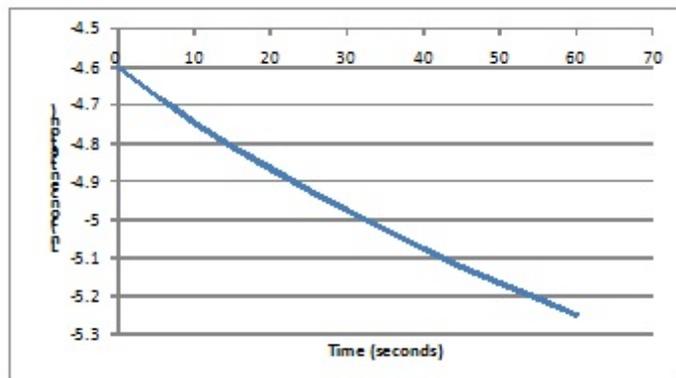
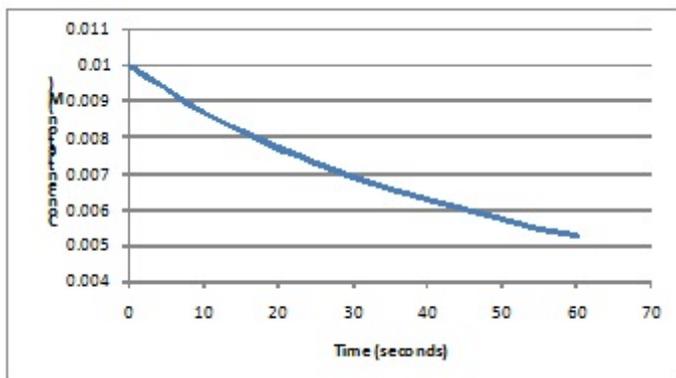
In all equations, $\text{rate} = k [A]^1 [B]^2 [C]^0$ There are a couple different ways to solve this, but the easiest is to solve the first equation for k;

$$.0075 = k [.005]^1 [.005]^2 [.005]^0$$

$$k = .0075 / (.005)(.005)^2; k = 60,000$$

Now plug this k and the other concentrations and solve for all other equations

3. Below are three plots depicting kinetic data for the disappearance of B in the reaction $B \rightarrow C$. From these plots determine the order of the reaction with respect to B, and the rate constant k of the reaction.

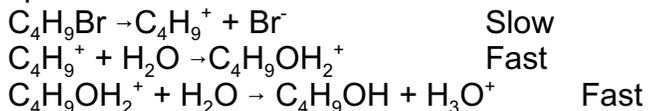


Of the above plots only the plot of $1/\text{concentration}$ vs time is linear, so the reaction is **second order**.

Note if you determined a different rate from each plot, and did not identify the order from the single plot that was linear I took 3 points off.

In this plot $k = \text{slope} = \Delta y / \Delta x = (190 - 100) / (60 - 0) =$
rate **1.5 liter/mol·sec**

4. A proposed mechanism for a reaction is:



A. Write the rate law expected for this mechanism.

$$\text{Rate} = k [\text{C}_4\text{H}_9\text{Br}]$$

B. What is the overall balanced equation for this reaction?



C. What are the intermediates in the proposed mechanism



5. Ozone is destroyed in the upper atmosphere by the reaction

$$\text{O}_3(\text{g}) + \text{O}(\text{g}) \rightarrow 2\text{O}_2(\text{g})$$

The uncatalyzed reaction has an activation energy of 14 kJ/mol. Freon can act as a catalyst for this reaction, and when this happens, the activation energy drops to 2.1 kJ/mol. How many times faster is the catalyzed reaction than the uncatalyzed reaction at 25°C.

$$\begin{aligned} \text{Rate catalyzed reaction} &= A e^{-2100/(8.314 \cdot 298)} \\ \text{Rate uncatalyzed reaction} &= A e^{-14000/(8.314 \cdot 298)} \end{aligned}$$

Dividing the catalyzed rate by the uncatalyzed rate

$$\begin{aligned} \text{Ratio} &= A e^{-2100/(8.314 \cdot 298)} / A e^{-14000/(8.314 \cdot 298)} \\ &\text{The A term drops out and you are left with} \\ &e^{-2100/(8.314 \cdot 298)} / e^{-14000/(8.314 \cdot 298)} = e^{-848} / e^{-5.65} = .428 / .00352 = 122 \end{aligned}$$

So catalyzed reaction is 122 x faster than uncatalyzed reaction.

6. At a certain temperature the reaction:



What is K_C for this reaction at 25° C?

$$\begin{aligned} K_p &= K_C(RT)^{\Delta n} \\ \Delta n &= 2-3 = -1 \\ .25 &= K_C(.08206 \cdot 298)^{-1}; \quad .25 = K_C / (.08206 \cdot 298) \\ K_C &= .25(.08206 \cdot 298) \\ &= 6.113 \end{aligned}$$

What is K_p for the reaction $2\text{SO}_3(\text{g}) \rightleftharpoons 2\text{SO}_2(\text{g}) + \text{O}_2(\text{g})$

$$\begin{aligned} &\text{This is reverse of first equation so } K_{\text{new}} = 1/K_{\text{original}} \\ K_p \text{ reverse} &= 1/K_p \text{ forward} \\ &= 1/.25 = 4 \end{aligned}$$

What is K_C for the reaction $\text{SO}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightleftharpoons \text{SO}_3(\text{g})$

$$\begin{aligned} \text{New equation} &= (\text{original equation}) \times \frac{1}{2} \\ \text{So } K \text{ new} &= (K \text{ original})^{\frac{1}{2}} \\ \text{This is a } K_C \text{ so I will use the answer from my first question} \\ K_C \text{ new} &= (6.113)^{\frac{1}{2}} = 2.4724 \end{aligned}$$

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

For which sets of conditions is the system at equilibrium? For those that are not at equilibrium, in which direction will the equilibrium shift?

A. $P_{\text{NO}} = .010 \text{ atm}$ $P_{\text{N}_2} = .11 \text{ atm}$ $P_{\text{O}_2} = 2.0 \text{ atm}$

$Q = (.11)(2.0)/(.01)^2 = 2200$; $Q < K$; not enough products reaction moves \rightarrow

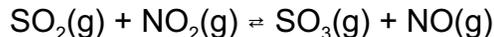
B. $P_{\text{NO}} = .0078 \text{ atm}$ $P_{\text{N}_2} = .36 \text{ atm}$ $P_{\text{O}_2} = 0.67 \text{ atm}$

$Q = (.36)(.67)/(.0078)^2 = 3960$; $Q > K$; too many products reaction moves \leftarrow

C. $P_{\text{NO}} = .0062 \text{ atm}$ $P_{\text{N}_2} = .51 \text{ atm}$ $P_{\text{O}_2} = .18 \text{ atm}$

$Q = (.51)(.18)/(.0062)^2 = 2380$; $Q \approx K$; at equilibrium

8. At a particular temperature $K = 3.75$ for the reaction



If all four gases have an initial concentration of .8M, calculate the equilibrium concentrations for each gas.

$Q = 1(1)/1(1) = 1$, $Q < K$, reaction will make products

	$\text{SO}_2(\text{g}) +$	$\text{NO}_2(\text{g}) \rightleftharpoons$	$\text{SO}_3(\text{g}) +$	$\text{NO}(\text{g})$
Initial	.8	.8	.8	.8
Change	-X	-X	+X	+X
Equilibrium	.8-X	.8-X	.8+X	.8+X

$$3.75 = (.8+X)(.8+X)/(.8-X)(.8-X) = (.8+X)^2/(\.8-X)^2$$

$$\text{sqrt}(3.75) = (.8+X)/(\.8-X)$$

$$1.936 = (.8+X)/(\.8-X)$$

$$1.936(.8-X) = .8+X$$

$$1.5488 - 1.936X = .8 + X$$

$$1.5488 - .8 = X + 1.936 X$$

$$0.7488 = 2.936 X$$

$$X = 0.7488/2.936 = .255$$

$$[\text{SO}_3] = [\text{NO}] = .8 + .255 = 1.055\text{M}$$

$$[\text{SO}_2] = [\text{NO}_2] = .8 - .255 = .545 \text{ M}$$