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:

<table>
<thead>
<tr>
<th>Rate (mole/sec)</th>
<th><a href="mole/l">A</a></th>
<th><a href="mole/l">B</a></th>
<th><a href="mole/l">C</a></th>
</tr>
</thead>
<tbody>
<tr>
<td>.0075</td>
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<tr>
<td>.03</td>
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<tr>
<td>.02058</td>
<td>.007</td>
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<td>.007</td>
</tr>
</tbody>
</table>

   In all equations, 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 → 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/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 \( k = \text{slope} = \Delta y/\Delta x = (190-100)/(60-0) = \text{rate 1.5 liter/mol·sec} \)

4. A proposed mechanism for a reaction is:

\[
\begin{align*}
\text{C}_4\text{H}_9\text{Br} & \text{ → C}_4\text{H}_9^+ + \text{Br}^- \quad \text{Slow} \\
\text{C}_4\text{H}_9^+ + \text{H}_2\text{O} & \text{ → C}_4\text{H}_9\text{OH}_2^+ \quad \text{Fast} \\
\text{C}_4\text{H}_9\text{OH}_2^+ + \text{H}_2\text{O} & \text{ → C}_4\text{H}_9\text{OH} + \text{H}_3\text{O}^+ \quad \text{Fast}
\end{align*}
\]

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?

\[
\text{C}_4\text{H}_9\text{Br} + 2 \text{H}_2\text{O} \text{ → C}_4\text{H}_9\text{OH} + \text{H}_3\text{O}^+
\]

C. What are the intermediates in the proposed mechanism

\[
\text{C}_4\text{H}_9^+, \text{ and C}_4\text{H}_9\text{OH}_2^+
\]
5. Ozone is destroyed in the upper atmosphere by the reaction
\[ \text{O}_3(g) + \text{O}(g) \rightarrow 2\text{O}_2(g) \]

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?

\[ \text{Rate catalyzed reaction} = A e^{-\frac{2100}{(8.314 \times 298)}} \]
\[ \text{Rate uncatalyzed reaction} = A e^{-\frac{14000}{(8.314 \times 298)}} \]

Dividing the catalyzed rate by the uncatalyzed rate

\[ \text{Ratio} = \frac{A e^{-\frac{2100}{(8.314 \times 298)}}}{A e^{-\frac{14000}{(8.314 \times 298)}}} \]

The \( A \) term drops out and you are left with

\[ e^{-\frac{2100}{(8.314 \times 298)}} / e^{-\frac{14000}{(8.314 \times 298)}} = e^{-5.65} / e^{-8.48} = 0.428 / 0.00352 = 122 \]

So catalyzed reaction is 122 x faster than uncatalyzed reaction.

6. At a certain temperature the reaction:
\[ 2\text{SO}_2 (g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g) \] has a \( K_p \) of .25

What is \( K_c \) for this reaction at 25°C?

\[ K_p = K_c (RT)^{\Delta n} \]
\[ \Delta n = 2-3 = -1 \]
\[ .25 = K_c (.08206 \times 298)^{-1} ; \quad .25 = K_c /(.08206 \times 298) \]
\[ K_c = .25 (0.8206 \times 298) \]
\[ = 6.113 \]

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

This is reverse of first equation so \( K_{new} = 1/K_{original} \)

\[ K_p \text{ reverse} = 1/K_p \text{ forward} \]
\[ = 1/.25 = 4 \]

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

New equation = (original equation) x \( \frac{1}{2} \)

So \( K_{new} = (K \text{ original})^{\frac{1}{2}} \)

This is a \( K_c \) so I will use the answer from my first question

\[ K_c \text{ new} = (6.113)^{\frac{1}{2}} = 2.4724 \]
7. The reaction $2\text{NO}(g) \rightarrow \text{N}_2(g) + \text{O}_2(g)$ has a $K_p$ of $2.4 \times 10^3$

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. $P_{\text{NO}} = 0.010 \text{ atm}$  $P_{\text{N}_2} = 0.11 \text{ atm}$  $P_{\text{O}_2} = 2.0 \text{ atm}$

$Q = (0.11)(2.0)/(0.01)^2 = 2200; Q<K; \text{not enough products} \text{ reaction moves } -$ 

B. $P_{\text{NO}} = 0.0078 \text{ atm}$  $P_{\text{N}_2} = 0.36 \text{ atm}$  $P_{\text{O}_2} = 0.67 \text{ atm}$

$Q = (0.36)(0.67)/(0.0078)^2 = 3960; Q>K; \text{too many products} \text{ reaction moves } -$ 

C. $P_{\text{NO}} = 0.0062 \text{ atm}$  $P_{\text{N}_2} = 0.51 \text{ atm}$  $P_{\text{O}_2} = 0.18 \text{ atm}$

$Q = (0.51)(0.18)/(0.0062)^2 = 2380; Q=K; \text{at equilibrium}$

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

$$\text{SO}_2(g) + \text{NO}_2(g) \rightarrow \text{SO}_3(g) + \text{NO}(g)$$

If all four gases have an initial concentration of $0.8\text{M}$, calculate the equilibrium concentrations for each gas.

$Q = 1(1)/1(1) = 1, Q<K, \text{reaction will make products}$

<table>
<thead>
<tr>
<th></th>
<th>$\text{SO}_2(g)$+</th>
<th>$\text{NO}_2(g)$ =</th>
<th>$\text{SO}_3(g)$ +</th>
<th>$\text{NO}(g)$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>.8</td>
<td>.8</td>
<td>.8</td>
<td>.8</td>
</tr>
<tr>
<td>Change</td>
<td>-X</td>
<td>-X</td>
<td>+X</td>
<td>+X</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>.8-X</td>
<td>.8-X</td>
<td>.8+X</td>
<td>.8+X</td>
</tr>
</tbody>
</table>

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

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

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

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

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

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

$$0.7488=2.936 X$$

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

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

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