$\qquad$

Directions: (1) Put your name on PART 1 and your name and signature on PART 2 of the exam where indicated.
(2) Sign the Aggie Code on PART 2 of this exam.
(3) Each multiple choice question is actually 2 questions on your scanning sheet. If you are sure of an answer, put the same answer down for both questions for 5 pts. If you cannot decide between two answers, put your best answer down for the first (odd) question and the other answer down for the second (even) question. If you get the first one correct you'll get 3 pts; if you get the second one correct you'll get 2 pts. If there is an ambiguous multiple choice question, use the last page to explain your answer.
(4) Do NOT write on the envelope.
(5) When finished, put everything in the envelope and wait to be excused. At the table, take everything out of the envelope. You can pick up the multiple choice part with the answers outside my office after 3pm.
(6) There are a total of 35 questions (19 actual questions).

## PART 1

1\&2. The rate law for the chemical reaction, $2 \mathrm{ClO}_{2}(\mathrm{aq})+2 \mathrm{OH}^{-}(\mathrm{aq}) \rightarrow \mathrm{ClO}_{3}^{-}(\mathrm{aq})+\mathrm{ClO}_{2}^{-}(\mathrm{aq})+\mathrm{H}_{2} \mathrm{O}(\ell)$ has been determined experimentally to be:

$$
\text { Rate }=\mathrm{k}\left[\mathrm{ClO}_{2}\right]^{2}\left[\mathrm{OH}^{-}\right]
$$

The reaction order with respect to the hydroxide ion is:
(a) $\mathrm{k}\left[\mathrm{OH}^{-}\right]^{2}$
(b) $\mathrm{k}\left[\mathrm{ClO}_{2}\right]^{2}\left[\mathrm{OH}^{-}\right]$
(c) 3
(d) 2
(e) 1

3\&4. Which is the correct $K_{\mathrm{c}}$ expression for the equilibrium: $2 \mathrm{C}(\mathrm{s})+\mathrm{O}_{2}(\mathrm{~g}) \underset{\leftarrow 2 \mathrm{CO}(g) \text { ? }}{\rightleftarrows}$
(a) $K_{\mathrm{c}}=\frac{2[\mathrm{CO}]}{2[\mathrm{C}]\left[\mathrm{O}_{2}\right]}$
(b) $K_{\mathrm{c}}=\frac{[\mathrm{C}]^{2}\left[\mathrm{O}_{2}\right]}{[\mathrm{CO}]^{2}}$
(c) $K_{\mathrm{c}}=\frac{[\mathrm{CO}]^{2}}{[\mathrm{C}]^{2}\left[\mathrm{O}_{2}\right]}$
(d) $K_{\mathrm{c}}=\frac{[\mathrm{CO}]^{2}}{\left[\mathrm{O}_{2}\right]}$
(e) $K_{\mathrm{c}}=\frac{\left[\mathrm{O}_{2}\right]}{[\mathrm{CO}]^{2}}$

5\&6. In general, as temperature increases, the rate of a chemical reaction
(a) decreases due to fewer collisions with proper molecular orientation.
(b) increases for exothermic reactions, but decreases for endothermic reactions.
(c) increases due to a greater number of effective collisions.
(d) decreases due to an increase in the activation energy.
(e) remains the same since the activation energy for the forward and the reverse reaction are lowered by the same amount.

7\&8. The units of the specific rate constant for a reaction that is second order overall is $\qquad$ .
(a) $\mathrm{M}^{-1} \cdot \mathrm{~s}^{-1}$
(b) $\mathrm{M}^{-2} \cdot \mathrm{~s}^{-1}$
(c) $\mathrm{M}^{-3} \cdot \mathrm{~s}^{-1}$
(d) $\cdot \mathrm{s}^{-1}$
(e) $M \cdot \mathrm{~s}^{-1}$

9\&10. Which statement or statements are TRUE about catalysts?
(1) Catalysts change the mechanism of a reaction.
(2) A catalyst cannot appear as a reactant in a rate law expression.
(3) A catalyst lowers the activation energy of both the forward and reverse reaction.
(a) $1 \& 2$
(b) $2 \& 3$
(c) $1 \& 3$
(d) only 1
(e) only 3

11\&12. Consider the reaction: $A+B \rightarrow C$. When the concentration of $A$ doubles and the concentration of $B$ is halved, the rate decreases by a factor of 2 . The appropriate rate law expression is:
(a) Rate $=\mathrm{k}[\mathrm{A}]$
(b) Rate $=k[A][B]^{2}$
(c) Rate $=\mathrm{k}[\mathrm{A}]^{2}[B]^{2}$
(d) Rate $=k[A]^{2}[B]$
(e) Rate $=\mathrm{k}[\mathrm{A}][\mathrm{B}]$

13\&14. The dissolution of ammonium nitrate occurs spontaneously in water at $25^{\circ} \mathrm{C}$. As the solid dissolves, the temperature of the water decreases. What are the signs of $\Delta \mathrm{H}, \Delta \mathrm{S}$ and $\Delta \mathrm{G}$ for this process?
(a) $\Delta \mathrm{H}>0, \Delta \mathrm{~S}<0$ and $\Delta \mathrm{G}>0$
(b) $\Delta \mathrm{H}>0, \Delta \mathrm{~S}>0$ and $\Delta \mathrm{G}<0$
(c) $\Delta \mathrm{H}>0, \Delta \mathrm{~S}<0$ and $\Delta \mathrm{G}<0$
(d) $\Delta \mathrm{H}<0, \Delta \mathrm{~S}<0$ and $\Delta \mathrm{G}<0$
(e) $\Delta H<0, \Delta S>0$ and $\Delta G>0$

15\&16. Determine the activation energy for the forward reaction if the $\Delta H=\Delta E$ for the reaction is +50 kJ and the activation energy for the reverse reaction is 70 kJ . (It may be useful to draw the potential energy diagram.)
(a) +120 kJ
(b) -120 kJ
(c) +70 kJ
(d) -20 kJ
(e) +20 kJ

17\&18. Consider the equilibrium, $2 \mathrm{~A} \underset{\leftarrow}{\rightleftarrows} \mathrm{~A}_{2}$ with $\mathrm{Kc} \gg 1$. Which picture represents this system?
(a)

(b)

(c)

(d)

(e)


19\&20. At what temperatures will a reaction be spontaneous if $\Delta H^{\circ}=+45 \mathrm{~kJ} / \mathrm{mol} \mathrm{rxn}$ and $\Delta S^{\circ}=+312 \mathrm{~J} / \mathrm{mol}$ rxn•K.
(a) Temperatures between 45 K and 312 K
(b) The reaction will be spontaneous at any temperature.
(c) The reaction will never be spontaneous.
(d) All temperatures below 144 K
(e) All temperatures above 144 K

21\&22. Consider the gas-phase equilibrium at a certain temperature: $A \rightleftarrows 2 B$
A reaction begins with 2.0 moles of A in a 1.0 L container. When the system reaches equilibrium, there is 1.0 mole of $A$ present. What is the value of the equilibrium constant, $\mathrm{K}_{\mathrm{c}}$ ?
(a) 1
(b) 2
(c) $1 / 2$
(d) 4
(e) $1 / 4$

23\&24. Rate data were collected for the following reaction at a particular temperature. What is the correct rate law expression?

$$
3 X+Y+2 Z \rightarrow \text { products }
$$

| Experiment | $[\mathrm{X}]_{\text {jinitial }}$ | $[\mathrm{Y}]_{\text {initial }}$ | $[Z]_{\text {initial }}$ | Initial Rate of Reaction |
| :---: | :---: | :---: | :---: | :---: |
| 1 | 0.30 M | 0.20 M | 0.10 M | $6.4 \times 10^{-3} \mathrm{M} / \mathrm{s}$ |
| 2 | 0.30 M | 0.20 M | 0.30 M | $5.76 \times 10^{-2} \mathrm{M} / \mathrm{s}$ |
| 3 | 0.60 M | 0.40 M | 0.10 M | $1.28 \times 10^{-2} \mathrm{M} / \mathrm{s}$ |
| 4 | 0.90 M | 0.20 M | 0.60 M | $2.3 \times 10^{-1} \mathrm{M} / \mathrm{s}$ |
| 5 | 0.80 M | 0.40 M | 0.10 M | $1.28 \times 10^{-2} \mathrm{M} / \mathrm{s}$ |

(a) Rate $=k[X][Y][Z]$
(b) Rate $=\mathrm{k}[\mathrm{Y}]^{2}[\mathrm{Z}]$
(c) Rate $=\mathrm{k}[Y][Z]^{2}$
(d) Rate $=k[X]^{2}[Z]^{2}$
(e) Rate $=\mathrm{k}[\mathrm{Y}]^{2}$

25\&26. Consider the following equilibrium reaction: $\mathrm{N}_{2}(\mathrm{~g})+\mathrm{O}_{2}(\mathrm{~g}) \underset{\mathrm{g}}{\rightleftarrows} \mathrm{NO}(\mathrm{g}) \quad K_{\mathrm{c}}=4.0 \times 10^{-4}$ at 2010 K If the initial concentration of NO in a closed container is 3.00 M , what will be the concentration of NO after the system finally reaches equilibrium at 2010 K ?
(a) 0.49 M
(b) 0.10 M
(c) 0.007 M
(d) 0.030 M
(e) 1.30 M

27\&28. Hydrogen peroxide decomposes into water and oxygen in a first order process:

$$
\mathrm{H}_{2} \mathrm{O}_{2}(\mathrm{aq}) \rightarrow \mathrm{H}_{2} \mathrm{O}(\ell)+1 / 2 \mathrm{O}_{2}(\mathrm{~g})
$$

At $20^{\circ} \mathrm{C}$, the half-life for the reaction is $3.92 \times 10^{4}$ seconds. If the initial concentration of $\mathrm{H}_{2} \mathrm{O}_{2}$ is 0.32 M , how much time (in minutes) must pass to reduce the concentration to $82 \%$ of its original concentration?
(a) 190 min
(b) $50 . \mathrm{min}$
(c) 310 min
(d) 11 min
(e) 94 min

29\&30. Calculate the $\Delta \mathrm{G}$ for the reaction at $25^{\circ} \mathrm{C}: \mathrm{Fe}_{3} \mathrm{O}_{4}(\mathrm{~s}) \rightarrow 3 \mathrm{Fe}(\mathrm{s})+2 \mathrm{O}_{2}(\mathrm{~g})$

| Species | $\Delta H_{\mathrm{f} 298}^{\circ}(\mathrm{kJ} / \mathrm{mol})$ | $S^{\circ}(\mathrm{J} / \mathrm{mol} \cdot \mathrm{K})$ |
| :---: | :---: | :---: |
| $\mathrm{Fe}_{3} \mathrm{O}_{4}(\mathrm{~s})$ | -1118.4 | 146.4 |
| $\mathrm{Fe}(\mathrm{s})$ | 0 | 27.8 |
| $\mathrm{O}_{2}(\mathrm{~g})$ | 0 | 205.1 |

(a) -1221.9 kJ
(b) -1014.9 kJ
(c) -771.2 kJ
(d) +771.2 kJ
(e) +1014.9 kJ

31\&32. The activation energy for a reaction was determined to be $145 \mathrm{~kJ} / \mathrm{mol} \mathrm{rxn}$. If the rate constant is $3.0 \times 10^{-3} \mathrm{~s}^{-1}$ at $25^{\circ} \mathrm{C}$, what would the rate constant be at when the temperature was raised to $50^{\circ} \mathrm{C}$ ? $\ln \left(\frac{k_{2}}{k_{1}}\right)=-\frac{E_{a}}{R}\left(\frac{1}{T_{2}}-\frac{1}{T_{1}}\right)$ or $\ln \left(\frac{k_{2}}{k_{1}}\right)=\frac{E_{a}}{R}\left(\frac{T_{2}-T_{1}}{T_{1} T_{2}}\right)$ or $\ln \left(\frac{k_{2}}{k_{1}}\right)=\frac{E_{a}}{R}\left(\frac{1}{T_{1}}-\frac{1}{T_{2}}\right) \quad R=8.314 \mathrm{~J} / \mathrm{mol} \cdot \mathrm{K}$
(a) $387 \mathrm{~s}^{-1}$
(b) $2.11 \mathrm{~s}^{-1}$
(c) $0.278 \mathrm{~s}^{-1}$
(d) $0.0415 \mathrm{~s}^{-1}$
(e) $6.0 \times 10^{-3} \mathrm{~s}^{-1}$

CHEMISTRY 102
EXAM 2
Form C
FALL 2010 NAME $\qquad$
Section 502

## PART 2

Please read and sign: "On my honor, as an Aggie, I have neither given nor received unauthorized aid on this exam."
33. Consider the following gas phase reaction: $X+2 Y \rightarrow Z$.

The experimentally-derived rate-law expression is: Rate $=k[X]^{2}$.
For each mechanism below, write the rate-law expression that fits it AND state whether or not that mechanism could be the correct one. Show your work for Part (c) - I want to know that you understand how the equilibrium constant works into the problem.
(2 pts)
$\begin{array}{lll}\text { (a) } \begin{array}{lll}\text { Step 1 } & X+Y \rightarrow C & \text { slow } \\ \text { Step 2 } & C+Y \rightarrow D & \text { fast } \\ & \text { Step 3 } & D \rightarrow Z\end{array} & \text { fast } \\ & \text { Overall } & X+2 Y \rightarrow Z\end{array}$
(2 pts)
(b) Step $1 \quad \mathrm{X}+\mathrm{X} \rightarrow \mathrm{X}_{2} \quad$ slow

Step $2 \quad X_{2}+Y \rightarrow C+X \quad$ fast
Step $2 \quad \mathrm{C}+\mathrm{Y} \rightarrow \mathrm{Z}$ fast
Overall $X+2 Y \rightarrow Z$
(4 pts)
(c) Step $1 \quad \mathrm{X}+\mathrm{D} \underset{\leftarrow}{\rightleftarrows} \quad$ fast equilibrium

Step $2 \quad \mathrm{E}+\mathrm{Y} \rightarrow \mathrm{F}$ slow
Step $3 \quad F+Y \rightarrow Z+D$ fast
Overall $\quad X+2 Y \rightarrow Z$
(1 pt) (d) In mechanism (c), what is/are the intermediate(s) (if any): $\qquad$
(1 pt) In mechanism (c), what is the catalyst (if any): $\qquad$
OVER $\Rightarrow$
(3 pts) 34. Consider the equilibrium: $2 \mathrm{X}(\mathrm{g}) \underset{\mathrm{y}}{\rightleftarrows} \mathrm{Y}(\mathrm{g})$ with $\mathrm{K}_{\mathrm{c}}=16$.
Calculate the equilibrium constant for $2 \mathrm{Y}(\mathrm{g}) \stackrel{\mathrm{X}(\mathrm{g})}{\rightleftarrows}$
(4 pts) 35. Consider the equilibrium: $\mathrm{A}(\mathrm{g}) \underset{\mathrm{B}}{\rightleftarrows} \mathrm{g})+2 \mathrm{C}(\mathrm{g})$
The reaction as written is strongly exothermic. Predict how the following changes will affect the moles of C at equilibrium,
The possible answers are: increase (I), decrease (D), or remain unchanged (U)

|  | Moles of $\mathbf{C}$ |
| :--- | :--- |
| (a) The concentration of B is increased. |  |
| (b) The temperature is increased. |  |
| (c) A catalyst is introduced. |  |
| (d) The volume of the container is doubled. |  |

(3 pt) How does temperature affect the value of $\mathrm{K}_{\mathrm{c}}$ for this reaction? Explain briefly.

## SCRAP PAPER OR COMMENTS ON EXAM

## CHEMISTRY 102

EXAM 2 Form C

FALL 2010 NAME
S 502

