STATE UNIVERSITY OF NEW YORK COLLEGE AT ONEONTA,
DEPARTMENT OF CHEMISTRY & BIOCHEMISTRY
CHEMISTRY 112, GENERAL CHEMISTRY II, EXAM 2, APRIL 11, 2012

You have 80 minutes to complete this exam. Fill in the appropriate selection on your scantron form. You may keep the exam.

1. For the hypothetical reaction \(2A + B \rightarrow 2C + D\), the initial rate of disappearance of \(A\) is \(2.0 \times 10^{-2}\) mol/(L \cdot s). What is the initial rate of disappearance of \(B\)?
   a) \(1.4 \times 10^{-1}\) mol/(L \cdot s)
   b) \(8.0 \times 10^{-2}\) mol/(L \cdot s)
   c) \(1.0 \times 10^{-2}\) mol/(L \cdot s)
   d) \(1.4 \times 10^{-2}\) mol/(L \cdot s)
   e) \(4.0 \times 10^{-4}\) mol/(L \cdot s)

   Rate of disappearance of \(B = \frac{1}{2}\) Rate of disappearance of \(A\).
   \[= \frac{1}{2} \times 2.0 \times 10^{-2}\] \text{ mol/(L \cdot s)} = \(1.0 \times 10^{-2}\) \text{ mol/(L \cdot s)}

2. Which of the following statements is true concerning the reaction given below?
   \(2H_2S(g) + O_2(g) \rightarrow 2S(s) + 2H_2O(g)\)
   a) The reaction is second-order in \(H_2S(g)\) and first-order in \(O_2(g)\).
   b) The reaction is first-order in \(H_2S(g)\) and second-order in \(O_2(g)\).
   c) The rate law is \(\text{Rate} = k[H_2S]^2[O_2]\).
   d) The rate law is \(\text{Rate} = k[H_2S][O_2]\).
   e) The rate law may be determined only by experiment.

3. Ozone reacts with nitrogen dioxide to produce oxygen and dinitrogen pentoxide according to the following chemical equation:
   \(O_3(g) + 2NO_2(g) \rightarrow O_2(g) + N_2O_5(g)\)

   The rate law for this reaction is \(\text{Rate} = k[O_3][NO_2]\). If concentration is measured in moles per liter and time is measured in seconds, what are the units of \(k\)?
   a) \(\text{mol} \cdot \text{L}^{-1} \cdot \text{s}^{-1}\)
   b) \(\text{L} \cdot \text{mol}^{-1} \cdot \text{s}^{-1}\)
   c) \(\text{mol}^2 \cdot \text{L}^{-2} \cdot \text{s}^{-1}\)
   d) \(\text{L} \cdot \text{mol}^{-1} \cdot \text{s}^{-1}\)
   e) \(\text{L}^2 \cdot \text{mol}^{-2} \cdot \text{s}^{-1}\)

4. Nitrosyl chloride is produced from the reaction of nitrogen monoxide and chlorine:
   \(2NO(g) + Cl_2(g) \rightarrow 2NOCl(g)\)

   The following initial rates at a given temperature were obtained for the concentrations listed below.

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Initial Rate (\text{mol} \cdot \text{L}^{-1} \cdot \text{h}^{-1})</th>
<th>([\text{NO}]_0) (\text{mol} \cdot \text{L}^{-1})</th>
<th>([\text{Cl}_2]_0) (\text{mol} \cdot \text{L}^{-2})</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>2.21</td>
<td>0.25</td>
<td>0.25</td>
</tr>
<tr>
<td>2</td>
<td>8.34</td>
<td>0.50</td>
<td>0.25</td>
</tr>
<tr>
<td>3</td>
<td>4.42</td>
<td>0.25</td>
<td>0.50</td>
</tr>
</tbody>
</table>

   From the data, what is the experimental rate law?

   Order w.r.t. \([\text{Cl}_2]\) consider exp 1 and 2

   \[
   \left(\frac{0.5}{0.25}\right)^n = \left(\frac{8.34}{2.21}\right) = \left(\frac{4.24}{2.21}\right) = \left(\frac{0.5}{0.25}\right)^2 \Rightarrow n = 2.
   \]

   Order w.r.t. \([\text{NO}]_0\) consider exp 1 and 3

   \[
   \left(\frac{0.25}{0.25}\right)^n = \left(\frac{4.42}{2.21}\right) = \left(\frac{8.34}{2.21}\right)^2 \Rightarrow n = 1.
   \]
5. The reaction between selenous acid and the iodide ion in acid solution is

$$H_2SeO_3(aq) + 6I^-(aq) + 4H^+(aq) \rightarrow Se(s) + 2I_3^-(aq) + 3H_2O(l)$$

The data in the following table were measured at 0°C.

<table>
<thead>
<tr>
<th>Experiment</th>
<th>([H_2SeO_3]_0 (M))</th>
<th>([H^+]_0 (M))</th>
<th>([I^-]_0 (M))</th>
<th>Initial Rate [mol/(L \cdot s)]</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1.00 \times 10^{-4}</td>
<td>2.00 \times 10^{-2}</td>
<td>3.00 \times 10^{-2}</td>
<td>5.30 \times 10^{-7}</td>
</tr>
<tr>
<td>2</td>
<td>2.00 \times 10^{-4}</td>
<td>2.00 \times 10^{-2}</td>
<td>3.00 \times 10^{-2}</td>
<td>1.06 \times 10^{-6}</td>
</tr>
<tr>
<td>3</td>
<td>3.00 \times 10^{-4}</td>
<td>4.00 \times 10^{-2}</td>
<td>3.00 \times 10^{-2}</td>
<td>6.36 \times 10^{-6}</td>
</tr>
<tr>
<td>4</td>
<td>3.00 \times 10^{-4}</td>
<td>8.00 \times 10^{-2}</td>
<td>3.00 \times 10^{-2}</td>
<td>2.54 \times 10^{-5}</td>
</tr>
<tr>
<td>5</td>
<td>3.00 \times 10^{-4}</td>
<td>8.00 \times 10^{-2}</td>
<td>6.00 \times 10^{-2}</td>
<td>2.04 \times 10^{-4}</td>
</tr>
<tr>
<td>6</td>
<td>2.00 \times 10^{-4}</td>
<td>2.00 \times 10^{-2}</td>
<td>6.00 \times 10^{-2}</td>
<td>8.48 \times 10^{-6}</td>
</tr>
</tbody>
</table>

The overall order of this reaction is

$$\text{Rate} = K \left[ H_2SeO_3 \right] \left[ I^- \right]^2 \left[ H^+ \right]^3$$

overall order = (1 + 2 + 3) = 6

6. The radioactive nuclide \(^{63}\text{Ni}\) decays by a first-order process via the emission of a beta particle. The \(^{63}\text{Ni}\) nuclide has a half-life of 100 years. How long will it take for 93% of \(^{63}\text{Ni}\) to decay?

a) 10 years  
\[ t = \frac{\ln \left( \frac{100}{7} \right)}{-0.869} = 380 \text{ years} \]

b) 1.8 years  
\[ t = \frac{\ln \left( 0.27 \right)}{-0.869} = 380 \text{ years} \]

c) 170 years  
\[ t = \frac{\ln \left( 0.27 \right)}{-0.869} = 380 \text{ years} \]

d) 4.5 years  
\[ t = \frac{\ln \left( 0.27 \right)}{-0.869} = 380 \text{ years} \]

e) 380 years

7. The catalyzed pathway in a reaction mechanism has a ____ activation energy and thus causes a ______ reaction rate.

a) higher, lower  
b) higher, higher  
c) lower, higher  
d) lower, steady  
e) higher, steady

![Graph showing catalyzed and uncatalyzed reaction pathways]
8. A reaction that is second-order in one reactant has a rate constant of \(2.0 \times 10^{-2} \text{ L/(mol}\cdot\text{s})\). If the initial concentration of the reactant is 0.340 mol/L, how long will it take for the concentration to become 0.170 mol/L?

a) 35 s
b) 74 s
c) 150 s
d) 1500 s
e) 350 s

\[ \frac{1}{[A]_t} = k + \frac{1}{[A]_0} \]

\[ \frac{1}{0.170} = 2.0 \times 10^{-2} \cdot t + \frac{1}{0.340} \]

Solve for \(t\):

\[ t = 147 \text{ s} > 150 \text{ seconds} \]

9. The half-life of a reaction is

a) how long the reaction can run before stopping.

b) the time it takes for the reactant concentration to decrease to one-half of its initial value.

c) the time it takes for the amount of product formed to equal half the initial amount of reactant.

d) one-half of the time the reaction will take to go to completion.

e) twice as long for a second-order reaction as it is for a first-order reaction.

10. Which of the following corresponds to the correct integrated expression for a first-order reaction?

a) \( \frac{1}{A} - \frac{1}{B} = kt \)

b) \( \frac{1}{A} - \frac{1}{A_0} = kt \)

c) \( \frac{1}{A_0} - \ln \frac{A}{A_0} = kt \)

d) \( \ln A - \ln A_0 = -kt \)

e) \( t_{1/2} = \frac{0.693}{k} \)

11. The reaction \( A \rightarrow \text{ products} \) is first-order in \( A \). If the concentration of \( A \) is cut in half, the half-life of the reaction will

a) double.

b) decrease by a factor of 1/2.

c) decrease by a factor of 1/4.

d) remain constant.

e) quadruple.

12. the following reaction producing 1 mol of oxygen gas at a particular temperature, \( \Delta H = -200 \text{ kJ} \).

\[
\text{NO(g)} + \text{O}_3(g) \rightarrow \text{NO}_2(g) + \text{O}_2(g)
\]

The activation energy is 11 kJ/mol. What is the activation energy for the reverse reaction?
13. The rate constant for a first-order reaction is $1.8 \times 10^{-2} \text{ s}^{-1}$ at 690 K and $3.6 \times 10^{-2} \text{ s}^{-1}$ at 876 K. What is the activation energy?

a) 8.1 kJ/mol
b) 188 kJ/mol
c) 200 kJ/mol
d) 211 kJ/mol
e) 222 kJ/mol

B 19 kJ/mol

\[ k_1 = 1.8 \times 10^{-2} \quad T_1 = 690 \text{ K} \]
\[ k_2 = 3.6 \times 10^{-2} \quad T_2 = 876 \text{ K} \]

\[ \ln \left( \frac{k_2}{k_1} \right) = \frac{E_a}{R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right) \]

Solve for $E_a = 18727.32 \text{ J/mol} = 18.7 \text{ kJ/mol}$

14. At 400 K, an equilibrium mixture of $H_2$, $I_2$, and $H_I$ consists of 0.054 mol $H_2$, 0.019 mol $I_2$, and 0.059 mol $HI$ in a 1.00-L flask. What is the value of $K_p$ for the following equilibrium? \( R = 0.08211 \text{ L} \cdot \text{atm}/(\text{mol} \cdot \text{K}) \)

\[ 2HI(g) \rightleftharpoons H_2(g) + I_2(g) \]

a) 3.4
b) 21
c) 0.29

d) 0.017

e) 58

C 0.29

\[ K_p = \frac{[H_2][I_2]}{[HI]^2} = \frac{(0.054)(0.019)}{(0.059)^2} = 0.2947 \]

15. For the reaction $NO(g) + \frac{3}{2}O_2(g) \rightleftharpoons NO_2(g)$ at 750°C, what is the relationship between $K_c$ and $K_p$?

a) $K_c = K_p = 1.0$

b) $K_c = K_p$

c) $K_c = K_p \times (RT)^{\frac{3}{2}}$

d) $K_c = K_p \times (RT)^{\frac{3}{2}}$

e) $K_c = K_p \times (RT)^{\frac{3}{2}}$

E

\[ \Delta n = \left( 1 - \frac{3}{2} \right) = -\frac{1}{2} \]

\[ K_c = \frac{K_p}{(RT)^{\frac{3}{2}}} \]

16. At 298 K, the value of $K_c$ for the reaction $H_2(g) + Br_2(g) \rightleftharpoons 2HBr(g)$ is $2.0 \times 10^{-19}$. What is $K_c$ for $HBr(g) \rightleftharpoons \frac{1}{2}H_2(g) + \frac{1}{2}Br_2(g)$?

a) $-2.0 \times 10^{19}$

b) $5.0 \times 10^{20}$

c) $2.2 \times 10^{10}$

d) $4.0 \times 10^{38}$

e) $1.0 \times 10^{19}$

C

\[ K_c = \left( \frac{1}{2.0 \times 10^{19}} \right)^{\frac{1}{2}} = 2.236 \times 10^{-10} \]
17. What is the $K_c$ equilibrium-constant expression for the following equilibrium?

$$\text{NiO}(s) + \text{H}_2(g) \rightleftharpoons \text{Ni}(s) + \text{H}_2\text{O}(g)$$

\[
\begin{align*}
\text{a)} & \quad \frac{[\text{NiO}][\text{H}_2]}{[\text{Ni}][\text{H}_2\text{O}]} \\
\text{b)} & \quad \frac{[\text{Ni}][\text{H}_2\text{O}]}{[\text{NiO}][\text{H}_2]} \\
\text{c)} & \quad \frac{[\text{Ni}][\text{H}_2\text{O}]}{[\text{H}_2]} \\
\text{d)} & \quad \frac{[\text{H}_2]}{[\text{H}_2\text{O}]} \\
\text{e)} & \quad \frac{[\text{H}_2\text{O}]}{[\text{H}_2]} 
\end{align*}
\]

18. Consider the following reaction:

$$2\text{HF}(g) \rightleftharpoons \text{H}_2(g) + \text{F}_2(g) \quad (K = 1.00 \times 10^{-2})$$

Given that 1.00 mol of HF(g), 0.360 mol of H$_2$(g), and 0.750 mol of F$_2$(g) are mixed in a 5.00-L flask, determine the reaction quotient, $Q$.

\[
\begin{align*}
a) & \quad Q = 0.0540 \\
b) & \quad Q = 0.270 \\
c) & \quad Q = 0.0675 \\
d) & \quad Q = 2.11 \\
e) & \quad \text{none of these} \\
\end{align*}
\]

19. For the following reaction system at equilibrium, which one of the changes below would cause the equilibrium to shift to the right?

$$2\text{Br}_2(g) + 2\text{NO}(g) \rightleftharpoons 4\text{NOBr}(g); \Delta H^\circ = -30 \text{ kJ}$$

\[
\begin{align*}
a) & \quad \text{Increase the volume of the reaction vessel.} \\
b) & \quad \text{Remove some NO.} \\
c) & \quad \text{Add some NOBr.} \\
d) & \quad \text{Remove some Br}_2. \\
e) & \quad \text{Decrease the temperature.}
\end{align*}
\]

Exhibit summary 18.
20. For which of the following systems at equilibrium and at constant temperature will decreasing the volume cause the equilibrium to shift to the right?
   a) \( N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \)
   b) \( 2H_2O(g) \rightleftharpoons 2H_2(g) + O_2(g) \)
   c) \( 2NO(g) \rightleftharpoons 2NO(g) + O_2(g) \)
   d) \( NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g) \)
   e) \( H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g) \)

21. Which of the following statements is incorrect?
   a) An Arrhenius base is an electron-pair acceptor.
   b) An Arrhenius acid increases the concentration of hydronium ion.
   c) A Brønsted–Lowry base is a proton acceptor.
   d) A Brønsted–Lowry acid is a proton donor.
   e) Acids tend to be sour, and bases tend to be bitter.

22. What is a conjugate acid–base pair for the following equilibrium?
   \( H_2O(l) + HPO_4^{2-}(aq) \rightleftharpoons H_2PO_4^{-}(aq) + OH^-(aq) \)
   a) \( H_2O \) is an acid and \( OH^- \) is its conjugate base.
   b) \( H_2O \) is an acid and \( HPO_4^{2-} \) is its conjugate base.
   c) \( HPO_4^{2-} \) is an acid and \( OH^- \) is its conjugate base.
   d) \( HPO_4^{2-} \) is an acid and \( H_2PO_4^- \) is its conjugate base.
   e) \( HPO_4^{2-} \) is an acid and \( H_2O \) is its conjugate base.

23. Which of the following species cannot act as a Lewis base?
   a) \( Be^{2+} \)
   b) \( O^{2-} \)
   c) \( OH^- \)
   d) \( H_2O \)
   e) \( H_2O_3 \)

24. Rank \( H_3PO_4, H_2PO_4^-, \) and \( HPO_4^{2-} \) in order of increasing acid strength.
   a) \( H_3PO_4 < H_2PO_4^- < HPO_4^{2-} \)
   b) \( H_2PO_4^- < HPO_4^{2-} < H_3PO_4 \)
   c) \( HPO_4^{2-} < H_2PO_4^- < H_3PO_4 \)
   d) \( H_2PO_4^- < H_3PO_4 < HPO_4^{2-} \)
   e) \( HPO_4^{2-} < H_2PO_4^- < H_3PO_4 \)
25. What is the pH of a 0.0041 M Ba(OH)_2 solution?
   a) 2.09
   b) 11.61
   c) 2.39
   d) 9.20
   e) 11.91

26. What is the pOH of a solution prepared by dissolving 0.578 g of KOH(s) in 6.00 L of water?
   a) 2.765
   b) 12.013
   c) 1.987
   d) 7.000
   e) 11.235

27. The pH of a solution is 5.30. What is its hydronium-ion concentration?
   a) 5.30 M
   b) 5.0 x 10^{-6} M
   c) 2.0 x 10^{-9} M
   d) 5.0 x 10^{-3} M
   e) 2.0 x 10^{-9} M

28. What is the conjugate acid of H_2PO_4^{-}(aq)?
   a) H_3O^+
   b) H_2PO_4
   c) HPO_4^{2-}
   d) H_3P
   e) PO_4^{3-}

29. The acid strength decreases in the series HBr > HSO_4^{-} > CH_3COOH > HCN > HCO_3^{-}. Which of the following is the strongest base?
   a) CO_3^{2-}
   b) CN^{-}
   c) CH_3COO^{-}
   d) SO_4^{2-}
   e) Br^{-}

30. A solution in which the pOH is 12.5 would be described as
   a) very acidic.
   b) slightly acidic.
   c) neutral.
   d) very basic.
   e) slightly basic.

\[
\text{pH} = (14 - 12.5) = 1.5
\]

\[\text{Strongly Acidic}\]
KEY EQUATIONS

\[ \text{pH} = - \log ([\text{H}_3\text{O}^+]) \quad \text{pH} + \text{pOH} = 14 \]

\[ [\text{H}_3\text{O}^+] = 10^{-\text{pH}} \]

\[ K_v = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.00 \times 10^{-14} \text{ (at 25°C)} \]

\[ K_a K_b = K_v \]

Henderson-Hasselbalch equation \[ \text{pH} = pK_a + \log \frac{[\text{A}^-]}{[\text{HA}]} \text{ or simply as} \]

\[ \text{pH} = - \log K_a + \log \frac{[\text{buff. base}]}{[\text{buff. acid}]} \]

Gas law constant, \( R \): 0.0821 L atm mol\(^{-1}\) K\(^{-1}\)

8.314 J mol\(^{-1}\) K\(^{-1}\)

Ideal gas law: \( PV = nRT \)

\[ K_p = K_v (RT)^\frac{n}{\delta} \]

Integrated rate laws:

\[ \ln \frac{[A]_t}{[A]_o} = -kt \]

\[ \frac{1}{[A]_t} = kt + \frac{1}{[A]_o} \]

\[ [A] = -kt + [A]_o \]

Quadratic equation: \[ x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} \]

\[ \pi = MRT \]

Arrhenius equation: \[ k = A e^{E_a/RT} \]

or \[ \ln \frac{k_2}{k_1} = \frac{E_a}{R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right) \]

Half-life formulae

\[ t_{1/2} = \frac{1}{k[A]_o} \]

\[ t_{1/2} = \frac{0.693}{k} \]

\[ t_{1/2} = \frac{[A]_o}{2k} \]