Chemistry 105 B  Midterm Exam 2:

YOU MUST SHOW YOUR COMPLETE WORKING ON EACH OF PROBLEMS 5-8 TO RECEIVE FULL CREDIT. STATE AND JUSTIFY ANY APPROXIMATIONS YOU MAKE.

Relevant Data:

\[ K_p = \frac{K}{(RT)^{\Delta n}} \]
where \( \Delta n \) is the sum of the coefficients of gaseous products minus the sum of the coefficients of gaseous reactants in the balanced equation.

Gas Constant:
\[ R = 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1} \]

Autoionization of water:
\[ K_w = [H_3O^+] [OH^-] = 1.0 \times 10^{-14} \text{ @ 25 °C} \]
\[ pH = -\log_{10} [H_3O^+] \]
\[ pOH = -\log_{10} [OH^-] \]

The Henderson-Hasselbalch Equation:
\[ pH = pK_a + \log_{10}(\frac{[A^-]}{[HA]}) \]

Solution of the quadratic equation \( ax^2 + bx + c = 0 \) is given by:
\[ x = \frac{-b \pm \sqrt{(b^2 - 4ac)}}{2a} \]
Multiple Choice and short answer problems (you do not need to show working for problems 1 - 4):

1. Find the value of the equilibrium constant ($K$) at 500 K for

$$N_2 (g) + 3H_2 (g) \leftrightarrow 2 \text{NH}_3 (g)$$

The value for $K_p$ at 500 K is $1.5 \times 10^{-5}$ atm$^{-2}$.

(a) $7.5 \times 10^{-2}$ mol$^{-2}$ L$^2$
(b) $1.3 \times 10^{-2}$ mol$^{-2}$ L$^2$
(c) $8.9 \times 10^9$ mol$^{-2}$ L$^2$
(d) $2.5 \times 10^{-2}$ mol$^{-2}$ L$^2$
(e) $6.2 \times 10^{-4}$ mol$^{-2}$ L$^2$

(5 points)

2. Consider the following equilibrium:

$$4\text{NH}_3 (g) + 5 \text{O}_2 (g) \leftrightarrow 4 \text{NO} (g) + 6 \text{H}_2\text{O} (g)$$

The reaction is exothermic from left to right. Which of the following statements is completely correct in describing the effect of two different changes made to the system at equilibrium:

(a) if heat were added to the system, the equilibrium will shift to the right AND more ammonia would be produced if oxygen were added to the reaction mixture
(b) if heat were added to the system, the equilibrium will shift to the right AND more water would be produced if oxygen were added to the reaction mixture
(c) if heat were added to the system, the equilibrium will shift to the left AND the concentration of NO would increase if the overall pressure was decreased
(d) if heat were added to the system, the equilibrium will shift to the left AND the concentration of NO would decrease if the overall pressure was decreased

(5 points)
3. Ephedrine is a central nervous system stimulant. This compound is a weak base as follows:

$$\text{C}_{10}\text{H}_{15}\text{ON (aq) } + \text{ H}_2\text{O (l) } \leftrightarrow \text{C}_{10}\text{H}_{15}\text{ONH}^+ (aq) + \text{ OH}^- (aq)$$

$K_b$ for ephedrine is $1.4 \times 10^{-4}$. What $[\text{OH}^-]$ would you expect for a 0.10 M aqueous solution of ephedrine:

(a) $3.7 \times 10^{-3}$ M
(b) $8.5 \times 10^{-7}$ M
(c) $1.4 \times 10^{-4}$ M
(d) $7.1 \times 10^{-11}$ M

(5 points)

4. Complete the following table by calculating the missing entries and indicating whether the room temperature solution is acidic or basic.

<table>
<thead>
<tr>
<th>$[\text{H}^+]$</th>
<th>$[\text{OH}^-]$</th>
<th>pH</th>
<th>Acidic or basic?</th>
</tr>
</thead>
<tbody>
<tr>
<td>$2.5 \times 10^{-4}$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>$6.9 \times 10^{-8}$</td>
<td>3.2</td>
<td></td>
</tr>
</tbody>
</table>

(10 points)
Long Answers. You must show all your working.

5. Rainwater is acidic because several acidic oxides dissolve in water in the atmosphere. One such process involves carbon dioxide, which when dissolved exists in equilibrium with carbonic acid according to the following:

\[ \text{CO}_2 \ (aq) + \text{H}_2\text{O} \ (l) \leftrightarrow \text{H}_2\text{CO}_3 \ (aq) \] \hspace{1cm} (1)

\[ \text{H}_2\text{CO}_3 \ (aq) + \text{H}_2\text{O} \ (l) \leftrightarrow \text{HCO}_3^- \ (aq) + \text{H}_3\text{O}^+ \ (aq) \] \hspace{1cm} (2)

(a) Write the equilibrium constant expression for reactions 1 and 2.

(b) We wish to calculate the pH of rainwater that contains $3.7 \times 10^{-3}$ mol L$^{-1}$ dissolved CO$_2$. Normally equilibrium (1) is assumed to lie far to the right hand side, thus the available initial $[\text{H}_2\text{CO}_3] = 3.7 \times 10^{-3}$ mol L$^{-1}$. The value for the carbonic acid dissociation constant is $K_a = 4.3 \times 10^{-7}$. From this information, find the pH of the rainwater.

(c) Give two important gases, other than CO$_2$, produced by mankind that lead to increased acidity in rainwater.

(20 points)
6. I\textsubscript{2} dissolved in aqueous I\textsuperscript{-} solution exists in equilibrium with the tri-iodide ion, I\textsubscript{3}\textsuperscript{-} according to:

\[ \text{I}^-(\text{aq}) + \text{I}_2(\text{aq}) \rightleftharpoons \text{I}_3^-(\text{aq}) \]

If the equilibrium constant at room temperature is 768 mol\textsuperscript{-1} L, what are the equilibrium concentrations of I\textsubscript{3}\textsuperscript{-} and I\textsubscript{2} if 0.10 mol of I\textsubscript{2} is added to a 1.0 mol L\textsuperscript{-1} solution of I\textsuperscript{-} (aq)?

(20 points)
7. Blood is a buffer based on the $\text{H}_2\text{CO}_3 / \text{HCO}_3^-$ system.

$$\text{H}_2\text{CO}_3 (aq) + \text{H}_2\text{O} (l) \leftrightarrow \text{HCO}_3^- (aq) + \text{H}_3\text{O}^+ (aq)$$

The acid dissociation constant is given by $K_a = 4.3 \times 10^{-7}$.

In normal blood, the pH is 7.41. What is the ratio of CO$_2$ (usually written H$_2$CO$_3$) to HCO$_3^-$ in blood?
8. Explain why nitric acid is a strong acid based on its molecular structure. (Hint: consider both the ability of HNO₃ to donate H⁺ and the stability of the NO₃⁻ anion).

(15 points)