AP® CHEMISTRY
2017 SCORING GUIDELINES

Question 3

$$N_2(g) + O_2(g) \rightleftarrows 2\,NO(g)$$

At high temperatures, $N_2(g)$ and $O_2(g)$ can react to produce nitrogen monoxide, $NO(g)$, as represented by the equation above.

(a) Write the expression for the equilibrium constant, K_p, for the forward reaction.

$K_p = \dfrac{(P_{NO})^2}{(P_{N_2})(P_{O_2})}$	1 point is earned for a correct K_p expression.

(b) A student injects $N_2(g)$ and $O_2(g)$ into a previously evacuated, rigid vessel and raises the temperature of the vessel to 2000°C. At this temperature the initial partial pressures of $N_2(g)$ and $O_2(g)$ are 6.01 atm and 1.61 atm, respectively. The system is allowed to reach equilibrium. The partial pressure of $NO(g)$ at equilibrium is 0.122 atm. Calculate the value of K_p.

	$N_2(g)$ +	$O_2(g)$ $\rightleftarrows$	$2\,NO(g)$	
Initial	6.01	1.61	0	1 point is earned for the correct equilibrium partial pressures of reactants and products (may be implicit).
Change	$-x$	$-x$	$+2x$	
Equilibrium	$6.01-x$	$1.61-x$	0.122	
$2x = 0.122$ atm $\Rightarrow$ $x = 0.0610$ atm				
$K_p = \dfrac{(0.122)^2}{(5.95)(1.55)} = 0.00161$				1 point is earned for the correct calculation of K_p.

Nitrogen monoxide, $NO(g)$, can undergo further reactions to produce acids, such as HNO_2, a weak acid with a K_a of 4.0×10^{-4} and a pK_a of 3.40.

(c) A student is asked to make a buffer solution with a pH of 3.40 by using 0.100 *M* $HNO_2(aq)$ and 0.100 *M* $NaOH(aq)$.

(i) Explain why the addition of 0.100 *M* $NaOH(aq)$ to 0.100 *M* $HNO_2(aq)$ can result in the formation of a buffer solution. Include the net ionic equation for the reaction that occurs when the student adds the $NaOH(aq)$ to the $HNO_2(aq)$.

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Question 3 (continued)

NaOH will neutralize some of the HNO_2 to produce NO_2^-. The resulting solution contains a mixture of a weak acid and its conjugate base, which is a buffer solution. $HNO_2 + OH^- \rightarrow NO_2^- + H_2O$	1 point is earned for the recognition that the solution produced is a mixture of a weak acid and its conjugate base. 1 point is earned for the correct net ionic equation.

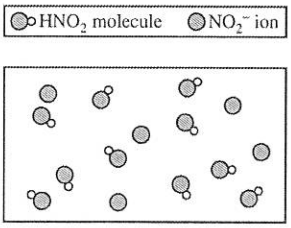
(ii) Determine the volume, in mL, of 0.100 *M* NaOH(*aq*) the student should add to 100. mL of 0.100 *M* HNO_2(*aq*) to make a buffer solution with a pH of 3.40. Justify your answer.

The student should add 50.0 mL of 0.100 *M* NaOH(*aq*). When half of the HNO_2 is converted to the conjugate base, $[HNO_2] = [NO_2^-]$, therefore the buffer has a pH equal to pK_a. OR $pH = pK_a + \log\frac{[NO_2^-]}{[HNO_2]}$, thus $pH = pK_a$ when $[HNO_2] = [NO_2^-]$	1 point is earned for the correct volume. 1 point is earned for clearly indicating a 1 to 1 ratio of HNO_2 and NO_2^- (calculation not required).

(d) A second student makes a buffer by dissolving 0.100 mol of $NaNO_2$(*s*) in 100. mL of 1.00 *M* HNO_2(*aq*). Which is more resistant to changes in pH when a strong acid or a strong base is added, the buffer made by the second student or the buffer made by the first student in part (c) ? Justify your answer.

The buffer made by the second student is more resistant to changes in pH because it contains a higher concentration of HNO_2 and NO_2^- to react with added H^+ or OH^- ions.	1 point is earned for the correct choice and a valid justification.

(e) A new buffer is made using HNO_2(*aq*) as one of the ingredients. A particulate representation of a small representative portion of the buffer solution is shown below. (Cations and water molecules are not shown.) Is the pH of the buffer represented in the diagram greater than, less than, or equal to 3.40? Justify your answer.



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Question 3 (continued)

The pH of the solution is less than 3.40. If $[HNO_2] = [NO_2^-]$, pH = pK_a, and the pH of the solution would be 3.40. Since $[HNO_2] > [NO_2^-]$, as represented in the diagram, the solution has a pH less than 3.40. OR $pH = pK_a + \log\frac{[NO_2^-]}{[HNO_2]} \Rightarrow pH = 3.40 + \log\frac{5}{10} \Rightarrow pH = 3.10$	1 point is earned for the correct choice. 1 point is earned for a valid justification.

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