PREPARATION WORKSHEET

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TA: ____________________ Section: ____________________

*Answer the following questions using information obtained from lecture, the textbook, or lab manual. The responses will be collected at the **beginning** of recitation and checked at the start of your lab section for credit.*

1. You have a solution of the weak acid HA and add some of the salt NaA to it.

a. What are the major species in the solution?

HA, A^-

b. What do you need to know to calculate $[H_3O^+]$ concentration and the pH of the solution, and how would you use this information?

K_a, [acid], [base], volume

follow the henderson-hasselbalch equation

c. How does the pH of the HA solution compare with that of the final mixture? Explain.

It is lower because the addition of the base raises the pH

2. A good buffer has: (a) about equal quantities of weak acid and conjugate base present; (b) as well as having a large concentration of each species present. Discuss this statement and why each aspect is important to successful buffering of a solution.

a) $pH = pK_a + \log\left(\frac{[base]}{[acid]}\right)$ base ≈ acid keeps ≈ 1 which allows pH = pKa

b) large concentration of makes any acid/base added have a relatively smaller impact

3. Calculate the pH at 25°C of 100.0 mL of a buffer that is 0.100 M NH_4Cl and 0.100 M NH_3 before and after the addition of 1.0 mL of 6 M HNO_3. The pK_a is given in the experiment. Show your work.

pKa ≈ 9.26

before $pH = 9.26 + \log(1)$
$= 9.26$

acid $\frac{0.1\ mol}{1000\ ml} \cdot 100\ mL = 10\ mmol$

base $\frac{0.1\ mol}{1000\ ml} \cdot 100\ mL = 10\ mmol$

added acid $\frac{6\ mol}{1000\ mL} \cdot 1\ mL = 0.006\ mmol$

added acid reacts with base

10 mmol base − 0.006 mmol = 9.994 mmol

creating acid

10 mmol acid + 0.006 mmol = 10.006

new base $\frac{9.994\ mmol}{101\ ml} = 0.09895$

new acid $\frac{10.006\ mmol}{101\ ml} = 0.09907$

New $pH = 9.26 + \log\left(\frac{0.9895}{0.09907}\right)$
$= 9.259 \approx 9.26$

4. When a person exercises, muscle contractions produce lactic acid. Moderate increases in lactic acid can be handled by the blood buffers without decreasing the pH of blood. However, excessive amounts of lactic acid can overload the blood buffer system resulting in the lowering of the blood pH. A condition called acidosis is diagnosed if the blood pH falls to 7.35 or lower. Assume the primary blood buffer system is the carbonate buffer system shown below. Calculate what happens to the $[H_2CO_3]/[HCO_3^-]$ ratio when the pH decreases from 7.40 to 7.35.

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq) \quad K_{a_1} = 4.3 \times 10^{-7}$$

$pH = pka + \log\left(\frac{[HCO_3^-]}{[H_2CO_3]}\right)$ $pk_a = -\log k_a = 6.367$

$\frac{[HCO_3^-]}{[H_2CO_3]} = 10^{pH-pka} = 10^{7.40-6.367} = 10.8$ $\frac{1}{}$ to flip $\frac{[HCO_3^-]}{[H_2CO_3]}$ to $\frac{[H_2CO_3]}{[HCO_3^-]} = 0.0926$

$= 10^{7.35-6.362} = 9.6$ $\frac{1}{}$ $= 0.1039$

$[H_2CO_3]$ increase while $[HCO_3^-]$ decrease

5. Phosphate ions are abundant in cells, both as the ions themselves and as important substituents on organic molecules. Most importantly, the pK_a for the $H_2PO_4^-$ ion is 7.20, which is very close to the normal pH in the body.

$$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$$

a. What should the ratio $[HPO_4^{2-}]/[H_2PO_4^-]$ be to control the pH at 7.4?

$\frac{[HPO_4^{2-}]}{[H_2PO_4^-]} = 10^{7.4-7.2} = 1.5849$

b. A typical total phosphate concentration in a cell, $[HPO_4^{2-}] + [H_2PO_4^-]$, is 2.0×10^{-2} M. What are the concentrations of HPO_4^{2-} and $H_2PO_4^-$ at pH 7.4?

$[HPO_4^{2+}] + [H_2PO_4^-] = 2 \cdot 10^{-2} M$

$[HPO_4^{2-}] = 2 \cdot 10^{-2} M - [H_2PO_4^-]$

$\frac{2 \cdot 10^{-2} - [HPO_4^{2-}]}{[HPO_4^{2-}]} = 1.5849$

$\frac{2 \cdot 10^{-2}}{[HPO_4^{2-}]} - \frac{[HPO_4^{2-}]}{[HPO_4^{2-}]} = 1.5849$

$\frac{2 \cdot 10^{-2}}{1.5849} = [HPO_4^{2-}] = 0.0126 M$

$2 \cdot 10^{-2} - 0.0126 = [H_2PO_4^-] = 0.0074 M$