

CHAPTER EXERCISE #3 (Group 4)

Section: CAS – 02 – 302A

Names of the Members:

- Ansay, Micah Abigail D.
- Camarillo, Kate T.
- Rañin, Jean Amor L.
- Tapel, Azalea G.

CHAPTER EXERCISE #3 BUFFER PREPARATION

Understand the problem carefully. Show your solution and derivation and mark your final answer in **red**. Chapter exercises must be done and submitted by group. Do not forget to write your section, group number, and all the names of each member. Submit your file in Word document format by downloading this file. You have **one (1) week** to work on this.

1. Calculate the mass of NaH_2PO_4 (FW= 119.98) and of Na_2HPO_4 (FW= 141.96) needed to prepare 500 mL of 1.0M $\text{NaH}_2\text{PO}_4/\text{Na}_2\text{HPO}_4$ buffer at pH = 8.0.

Given:

FW of NaH_2PO_4 = 119.98 g/mol

FW of Na_2HPO_4 = 141.96 g/mol

Molarity = 0.1 M

Volume = 500 mL $\times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.500 \text{ L}$

pH = 8.0



$\text{Moles}_{\text{total}} = MV = \text{mole NaH}_2\text{PO}_4 + \text{mole Na}_2\text{HPO}_4$

$\text{Moles}_{\text{total}} = (0.1 \text{ mol}) (0.500 \text{ L}) = \underline{0.05 \text{ moles}} = \text{mole NaH}_2\text{PO}_4 + \text{mole Na}_2\text{HPO}_4$

$\text{Moles Na}_2\text{HPO}_4 = 0.05 \text{ moles} - \text{mole NaH}_2\text{PO}_4$

$= 0.05 \text{ mole} - \text{mole NaH}_2\text{PO}_4 = 6.17 \text{ mole NaH}_2\text{PO}_4$

$= 0.05 \text{ mole} = 6.17 \text{ mole NaH}_2\text{PO}_4 + \text{mole NaH}_2\text{PO}_4$

$= 0.05 \text{ mole} = 7.17 \text{ mole NaH}_2\text{PO}_4$

$$\text{Mole NaH}_2\text{PO}_4 = \frac{0.50 \text{ mole}}{7.17} = \underline{6.97 \times 10^{-3} \text{ or } 0.00697}$$

$$\begin{aligned}\text{Mole NaHPO}_4 &= 0.05 \text{ moles} - \text{mole NaH}_2\text{PO}_4 \\ &= 0.05 \text{ moles} - 0.00697 \\ &= \underline{0.043 \text{ moles}}\end{aligned}$$

Using the Henderson-Hasselbalch equation:

$$\text{pH} = \text{pKa} + \log \frac{(\text{mole Na}_2\text{HPO}_4)}{(\text{mole NaH}_2\text{PO}_4)}$$

$$\text{pH} = \text{pKa} + \log \frac{(0.05 \text{ moles} - \text{mole Na}_2\text{HPO}_4)}{(\text{mole NaH}_2\text{PO}_4)}$$

$$8.0 = 7.21 + \log \frac{(0.05 \text{ moles} - \text{mole Na}_2\text{HPO}_4)}{(\text{mole NaH}_2\text{PO}_4)}$$

$$8.0 - 7.21 = \log \frac{(0.05 \text{ moles} - \text{mole Na}_2\text{HPO}_4)}{(\text{mole NaH}_2\text{PO}_4)}$$

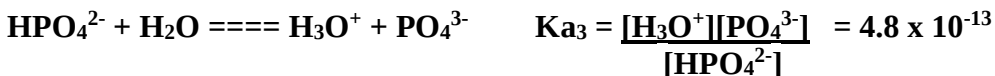
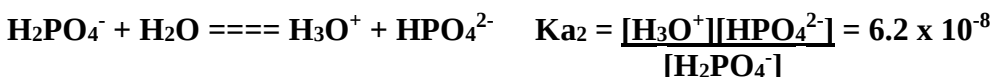
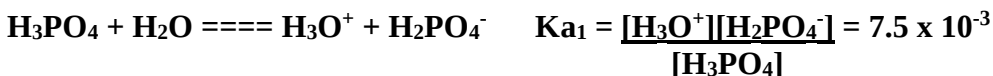
$$\log \frac{(0.05 \text{ moles} - \text{mole Na}_2\text{HPO}_4)}{(\text{mole NaH}_2\text{PO}_4)} = 0.79$$

$$\begin{aligned}\log \frac{(0.05 \text{ moles} - \text{mole Na}_2\text{HPO}_4)}{(\text{mole NaH}_2\text{PO}_4)} &= \text{antilog } (0.79) \\ &= \underline{6.17}\end{aligned}$$

$$\begin{aligned}\text{Mass Na}_2\text{HPO}_4 &= \text{mole Na}_2\text{HPO}_4 \times \text{FW of Na}_2\text{HPO}_4 \\ &= 0.043 \text{ moles} \times 141.96 \text{ g/mol} \\ &= \underline{6.10 \text{ g}}\end{aligned}$$

$$\begin{aligned}\text{Mass NaH}_2\text{PO}_4 &= \text{mole NaH}_2\text{PO}_4 \times \text{FW of NaH}_2\text{PO}_4 \\ &= 0.00697 \text{ moles} \times 119.98 \text{ g/mol} \\ &= \underline{0.84 \text{ g}}\end{aligned}$$

Calculate the change in pH and the buffer capacity of the buffer in #1 if 0.25 moles of H_3O^+ was added to the buffer.



$$\begin{aligned} \text{pH } K_{a1} &= -(\log 7.5 \times 10^{-3}) + \log \frac{[0.431]}{[0.069]} \\ &= 2.21 + 0.80 \\ &= \underline{3.01} \end{aligned}$$

$$\begin{aligned} \text{Buffer Capacity} &= \frac{0.25 \text{ moles}}{(3.01)(0.500 \text{ L})} \\ &= \frac{.25 \text{ mol}}{1.505 \text{ L}} \\ &= \underline{0.166} \end{aligned}$$

$$\begin{aligned} \text{pH } K_{a2} &= -(\log 6.2 \times 10^{-8}) + \log \frac{[0.431]}{[0.069]} \\ &= 7.21 + 0.80 \\ &= \underline{8.01} \end{aligned}$$

$$\begin{aligned} \text{Buffer Capacity} &= \frac{0.25 \text{ moles}}{(8.01)(0.500 \text{ L})} \\ &= \frac{.25 \text{ mol}}{4.005 \text{ L}} \\ &= \underline{0.0624} \end{aligned}$$

$$\begin{aligned} \text{pH } K_{a3} &= -(\log 4.8 \times 10^{-13}) + \log \frac{[0.431]}{[0.069]} \\ &= 12.32 + 0.80 \\ &= \underline{13.12} \end{aligned}$$

$$\begin{aligned} \text{Buffer Capacity} &= \frac{0.25 \text{ moles}}{(13.12)(0.500 \text{ L})} \\ &= \frac{.25 \text{ mol}}{6.56 \text{ L}} \\ &= \underline{0.0381} \end{aligned}$$