

Chapter 17 – Practice Problems with Buffers - ANSWERS

$$1. \quad (a) \quad K_a = \frac{[\text{H}^+][\text{OCl}^-]}{[\text{HOCl}]} = \frac{(x)(x)}{(0.025 \text{ M})} = 3.0 \times 10^{-8}$$

$$x^2 = 7.5 \times 10^{-10}$$

$$x = [\text{H}^+] = 2.7 \times 10^{-5} \text{ M}$$

$$(b) \quad \text{pH} = -\log(2.7 \times 10^{-5}) = 4.57$$

$$(c) \quad \% \text{ ionization of HOCl} = \frac{[\text{H}^+]}{[\text{HOCl}]} \times 100\% = \frac{2.7 \times 10^{-5}}{0.025} \times 100\% = 0.11\%$$

2. (a) The presence of additional OCl^- ions causes the equilibrium position to shift toward the **reactants**.

(b) The value of $[\text{H}^+]$ in Solution #2 should be **less than** the value of $[\text{H}^+]$ in Solution #1.

(c) The pH of Solution #2 should be **greater than** the pH of Solution #1.

(d) The % ionization of HOCl in Solution #2 should be **less than** the % ionization of HOCl in Solution #1.

$$3. \quad (a) \quad K_a = \frac{[\text{H}^+][\text{OCl}^-]}{[\text{HOCl}]} = \frac{[\text{H}^+](0.015 \text{ M})}{(0.025 \text{ M})} = 3.0 \times 10^{-8}$$

$$[\text{H}^+] = 5.0 \times 10^{-8} \text{ M}$$

$$(b) \quad \text{pH} = -\log(5.0 \times 10^{-8}) = 7.30$$

Alternatively, you could have used the Henderson-Hasselbalch equation as follows.

$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right) = -\log(3.0 \times 10^{-8}) + \log\left(\frac{0.015}{0.025}\right) = 7.52 + \log(0.60) = 7.52 - 0.22 = 7.30$$

$$(c) \quad \% \text{ ionization of HOCl} = \frac{[\text{H}^+]}{[\text{HOCl}]} \times 100\% = \frac{5.0 \times 10^{-8}}{0.025} \times 100\% = 0.0002\%$$

4. Solution B. A mixture of $\text{HNO}_2(aq)$ and $\text{NaNO}_2(aq)$ is a better buffer because it consists of a weak acid (HNO_2) and its conjugate base (NO_2^-). A buffer solution that contains a weak acid and its conjugate base is able to resist changes in pH because it can neutralize small quantities of strong acid or strong base while still maintaining a ratio of $\frac{[\text{A}^-]}{[\text{HA}]}$ that is approximately the same value.

On the other hand, Buffer A contains $\text{HNO}_3(aq)$, which is a strong acid. $\text{NaNO}_3(aq)$ is a neutral salt. This combination would not make a good buffer because it does not consist of a weak acid – conjugate base pair.

5. (a) The value of pK_a for HNO_2 is equal to $-\log(4.5 \times 10^{-4}) = 3.35$

(b)

Solution	Prediction	Justification
a mixture of 1.0 M HNO_2 and 1.0 M NaNO_2	$\text{pH} = pK_a$	$[\text{HNO}_2] = [\text{NO}_2^-]$
a mixture of 0.75 M HNO_2 and 0.55 M NaNO_2	$\text{pH} < pK_a$	$[\text{HNO}_2] > [\text{NO}_2^-]$
a mixture of 0.83 M HNO_2 and 1.1 M NaNO_2	$\text{pH} > pK_a$	$[\text{HNO}_2] < [\text{NO}_2^-]$

6. (a)
$$K_a = \frac{[\text{H}^+][\text{NO}_2^-]}{[\text{HNO}_2]} = \frac{[\text{H}^+](1.0 \text{ M})}{(1.0 \text{ M})} = 4.5 \times 10^{-4}$$

$$[\text{H}^+] = 4.5 \times 10^{-4} \text{ M}$$

$$\text{pH} = -\log(4.5 \times 10^{-4}) = 3.35$$

Alternatively, you could have used the Henderson-Hasselbalch equation as follows.

$$\text{pH} = pK_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right) = -\log(4.5 \times 10^{-4}) + \log\left(\frac{1.0}{1.0}\right) = 3.35$$

(b)
$$K_a = \frac{[\text{H}^+][\text{NO}_2^-]}{[\text{HNO}_2]} = \frac{[\text{H}^+](0.55 \text{ M})}{(0.75 \text{ M})} = 4.5 \times 10^{-4}$$

$$[\text{H}^+] = (4.5 \times 10^{-4}) \left(\frac{0.75}{0.55}\right) = 6.1 \times 10^{-4} \text{ M}$$

$$\text{pH} = -\log(6.1 \times 10^{-4}) = 3.21$$

Alternatively, you could have used the Henderson-Hasselbalch equation as follows.

$$\text{pH} = pK_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right) = -\log(4.5 \times 10^{-4}) + \log\left(\frac{0.55}{0.75}\right) = 3.35 - 0.13 = 3.22$$

$$6. \quad (c) \quad K_a = \frac{[\text{H}^+][\text{NO}_2^-]}{[\text{HNO}_2]} = \frac{[\text{H}^+](1.1 \text{ M})}{(0.83 \text{ M})} = 4.5 \times 10^{-4}$$

$$[\text{H}^+] = (4.5 \times 10^{-4}) \left(\frac{0.83}{1.1} \right) = 3.4 \times 10^{-4} \text{ M}$$

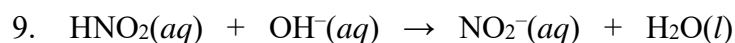
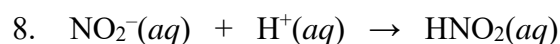
$$\text{pH} = -\log(3.4 \times 10^{-4}) = 3.47$$

Alternatively, you could have used the Henderson-Hasselbalch equation as follows.

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{A}^-]}{[\text{HA}]} \right) = -\log(4.5 \times 10^{-4}) + \log \left(\frac{1.1}{0.83} \right) = 3.35 + 0.12 = 3.47$$

7. The value of $\text{p}K_a$ for HNO_2 is equal to $-\log(4.5 \times 10^{-4}) = 3.35$

The pH range of an equimolar mixture of HNO_2 and NaNO_2 is between pH **2.35** and pH **4.35**.



10. Buffer X is more resistant to changes in pH when a strong acid or a strong base is added to it because it contains a higher concentration of HNO_2 and NO_2^- to react with added H^+ or OH^- ions.

11. The volume of 1.0 M NaOH needed is equal to 50.0 mL.

$$\text{Moles of HNO}_2 = 0.100 \text{ L} \times \frac{1.0 \text{ mol HNO}_2}{\text{L}} = 0.10 \text{ mol HNO}_2$$

Moles of NaOH needed to react with HALF of the $\text{HNO}_2 = 0.05 \text{ mol NaOH}$

When half of the HNO_2 is converted to the conjugate base, $[\text{HNO}_2] = [\text{NO}_2^-]$, and the pH of the solution is equal to the $\text{p}K_a$ of the weak acid, which is 3.35.

Another way to demonstrate this is

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{NO}_2^-]}{[\text{HNO}_2]} \right) \text{ and } \text{pH} = \text{p}K_a \text{ when } [\text{HNO}_2] = [\text{NO}_2^-]$$

12. (a) $\text{pH} = \text{p}K_a = -\log(1.5 \times 10^{-5}) = 4.82$

(b) $\text{pH} = \text{p}K_a = -\log(6.2 \times 10^{-8}) = 7.21$

(c) $\text{pH} = \text{p}K_a = -\log(5.6 \times 10^{-10}) = 9.25$

13.

Buffer Solution	Particulate Diagram of the Buffer Solution	Prediction	Justification
#1		$\text{pH} < 3.17$	$[\text{HF}] > [\text{F}^-]$
#2		$\text{pH} = 3.17$	$[\text{HF}] = [\text{F}^-]$
#3		$\text{pH} > 3.17$	$[\text{HF}] < [\text{F}^-]$