8.10 What Are Buffers? – How Do They Resist Drastic pH Changes

Addition of Strong Acid – H_3O^+

1M CH_3CO_2H / 1M $CH_3CO_2^-$

$\underbrace{CH_3CO_2H}_{\text{Weak Acid}}$ $\underbrace{CH_3CO_2^-}_{\text{Conj. Base}}$

$H_3O^+ \longrightarrow$

SA + WB = 100%

$H_3O^+ + \underbrace{CH_3CO_2^-}_{\text{Buffer base}} = H_2O(l) + \underbrace{CH_3CO_2H}_{\text{Buffer acid}}$

Overall Changes:

$[CH_3CO_2^-]$: ↓ Reacts with the added H_3O^+.

$[CH_3CO_2H]$: ↑ product of the reaction that removed the H_3O^+.

$[H_3O^+]$: ↑ not by much, increase due to $[CH_3CO_2H]$↑

pH: ↓ not by much, $[CH_3CO_2H]$↑ cause slight decrease in pH.

8.10 What Are Buffers? – How Do They Resist Drastic pH Changes

Addition of Strong Base – OH^-

1M CH_3CO_2H / 1M $CH_3CO_2^-$

Weak acid   Conj. base

OH^-

WA + SB = 100%

$\underbrace{CH_3CO_2H}_{\text{Buffer acid}} + OH^- = H_2O(l) + \underbrace{CH_3CO_2^-}_{\text{Buffer base}}$

Overall Changes:

$[CH_3CO_2H]$	:	$\downarrow$	Reacts with the added OH^-.
$[CH_3CO_2^-]$	:	$\uparrow$	Product of the Reaction that Removed the OH^-.
$[OH^-]$	:	$\uparrow$	Not by much. Increase due to $[CH_3CO_2^-]\uparrow$.
pH	:	$\uparrow$	Not by much. $[CH_3CO_2^-]\uparrow$ causes a slight increase in pH

8.10 What Are Buffers? – How Do They Resist Drastic pH Changes

A buffer solution made from HF and KF has a pH = 2.84.
Addition of OH- will cause –

1. Increase significantly *
2. Increase slightly
3. Decrease significantly*
4. Decrease slightly
5. Increase
6. Decrease

* Used only when Buffer Breaks down ... more on this Later ... not relevant here.

a)  pH ? 2 Adding OH^- will cause the solution to become a little more basic.

b)  pOH ? 4 If pH increases slightly then pOH will do reverse: pH + pOH = 14

c)  [HF] ? 6 $\underset{\text{Buffer acid}}{\underline{HF(aq)}} + OH^- \rightleftharpoons H_2O(l) + \underset{\text{Buffer base}}{\underline{F^-}}$

d)  $\frac{[F^-]}{[HF]}$? 5 See explanation for c): $[HF]\downarrow$, $[F^-]\uparrow$

8.10 What Are Buffers? – Making an Optimal Buffer Solution – pH and pKa

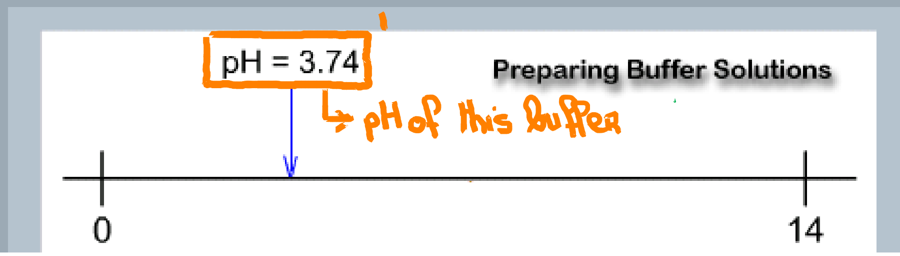
See class web site to see whether this holds true for other systems.
Note : [Buffer acid] = [Buffer base]

Acid: $[HCO_2H]$ = 0.10 M

Conj base HCO_2^-: $[NaHCO_2]$ = 0.10 M

Buffer system	K_a	pK_a
● $HCO_2H/NaHCO_2$	$K_a = 1.8 \times 10^{-4}$	$\boxed{pK_a = 3.74}$
○ $H_2CO_3/NaHCO_3$	$K_a = 4.2 \times 10^{-7}$	$pK_a = 6.38$
○ $HOCl/NaOCl$	$K_a = 3.5 \times 10^{-8}$	$pK_a = 7.46$
○ H_3BO_3/NaH_2BO_3	$K_a = 7.3 \times 10^{-10}$	$pK_a = 9.14$
○ NH_4Cl/NH_3	$K_a = 5.6 \times 10^{-10}$	$pK_a = 9.25$
○ $NaHCO_3/Na_2CO_3$	$K_a = 4.8 \times 10^{-11}$	$pK_a = 10.32$

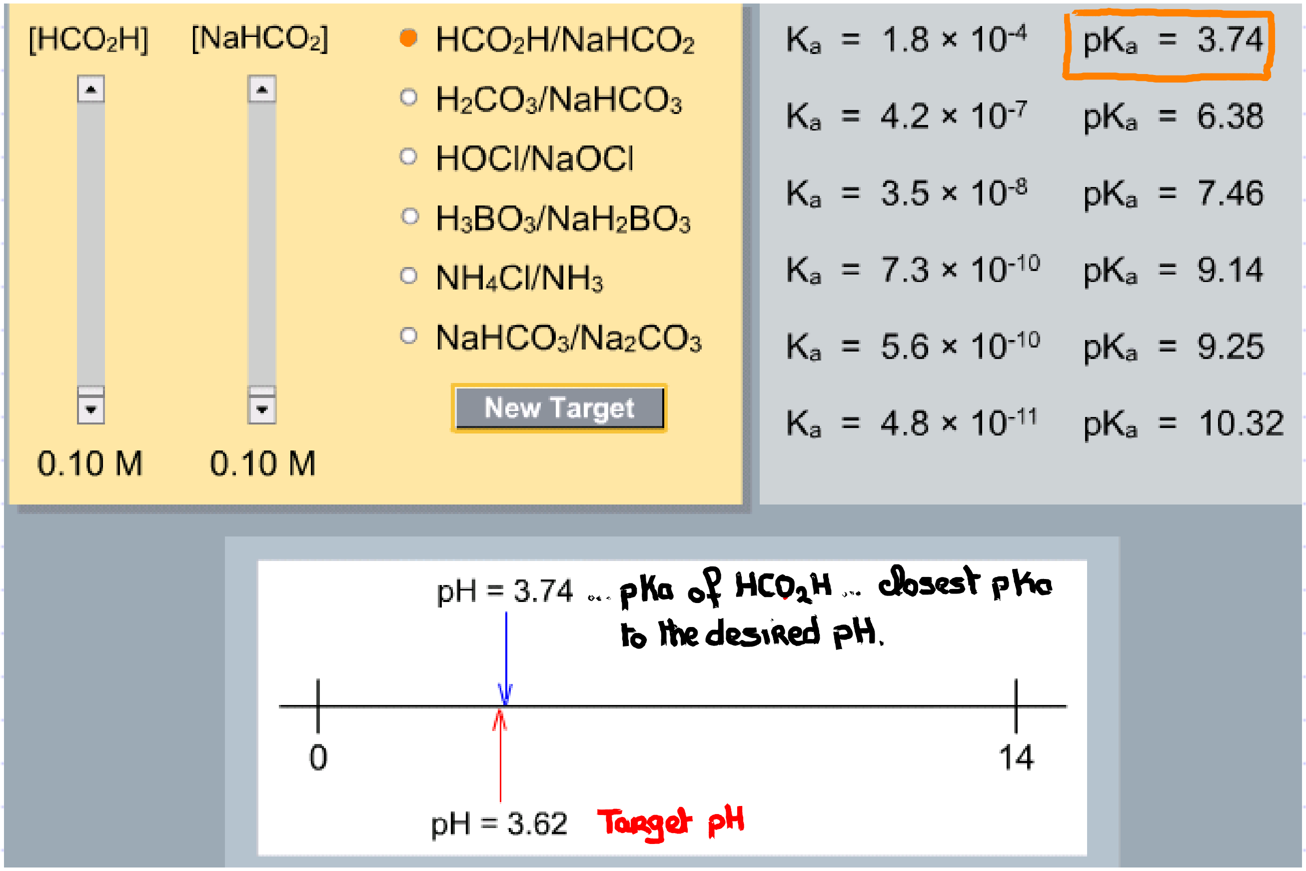
New Target



When [Buffer acid] = [Buffer base], then the pH of the buffer is equal to the pKa of the Buffer acid.

When chosing a buffer system one usually selects one whose pKa of the weak acid is closest to the desired pH.

9.10 What Are Buffers? – Making an Optimal Buffer Solution
Adjusting the pH of a Buffer



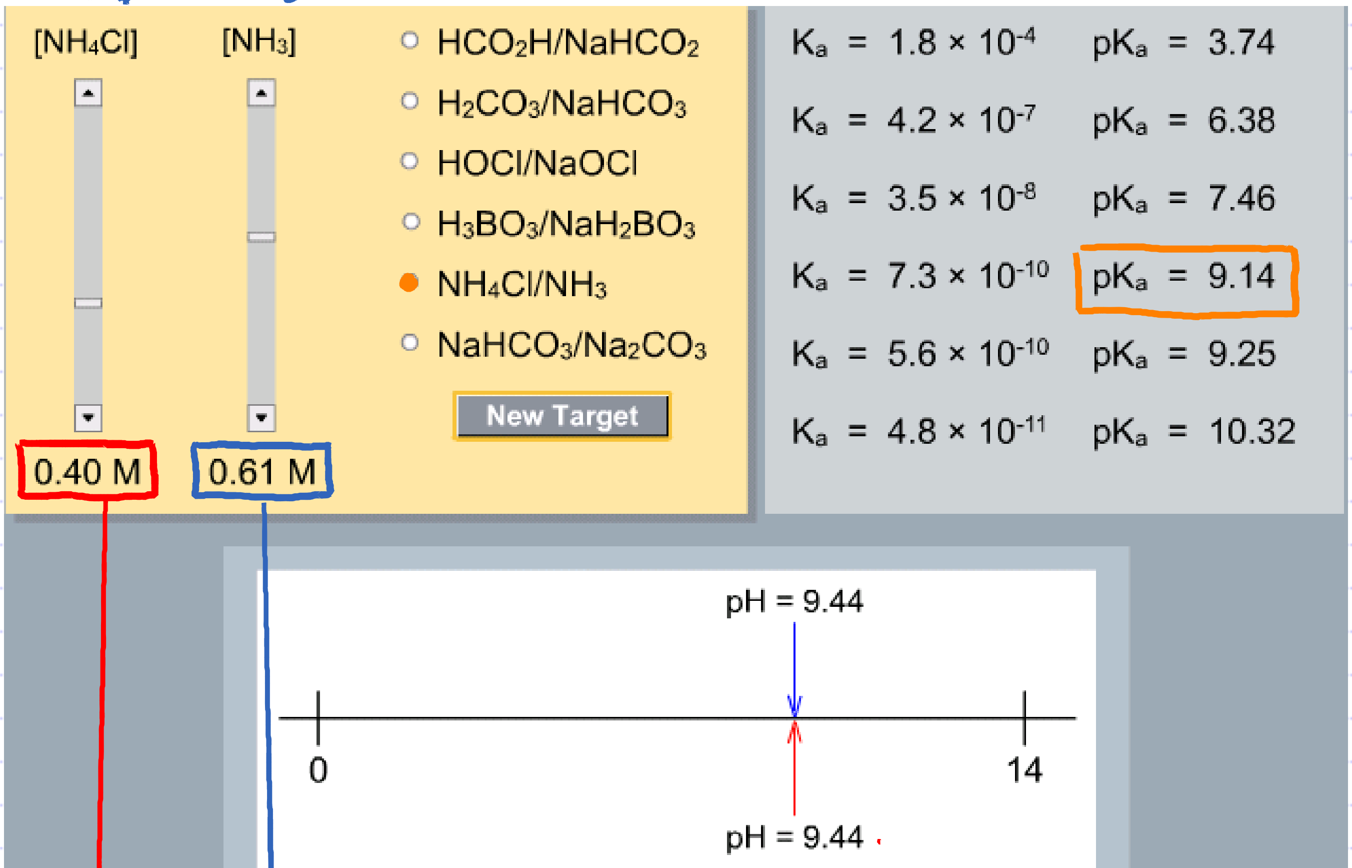
Since the desired pH is more acidic than the starting pH, either increase the [Buffer acid] or decrease the decrease [Buffer base]

8.10 What Are Buffers? – Making an Optimal Buffer Solution

Buffer Capacity

Buffer acid : NH_4^+
Buffer base : NH_3

NH_4^+ NH_3



Maximum amount of H_3O^+ that can be removed
Maximum amount of OH^- that can be removed

Buffer Capacity

8.10 What Are Buffers? – Identifying Buffer Solutions

How many of the following aqueous solutions are buffers? 2

a) $0.24\ M\ HI + 0.18\ M\ NaI$	X: HI is a strong acid.
d) $0.10\ M\ CH_3COOH + 0.18\ M\ CH_3COOK$	✓: CH_3COOH (WA) / CH_3COO^- its conjugate base.
c) $0.27\ M\ NH_4Br + 0.31\ M\ NH_3$	✓: NH_4^+ (WA) / NH_3 its conjugate base.
b) $0.34\ M\ NH_4NO_3 + 0.39\ M\ NaNO_3$	X: NH_4^+ (WA) but NO_3^- is not its conjugate base.

8.10 What Are Buffers? – How Do They Resist Drastic pH Changes

A 1L solution contains 0.25 mol of NaCN and 0.15 mol of HCN.

1. Increase significantly
2. Increase slightly
3. Decrease significantly
4. Decrease slightly
5. Increase
6. Decrease

a) Addition of 0.1 mol of HCl will case the [HCN] to – 5

$H_3O^+ + CN^- = H_2O(l) + HCN(aq)$

b) Addition of 0.1 mol of HCl will case the pOH to – 2

pH, ↓ slightly. pH + pOH = 14 thus pOH ↑ slightly.

c) Addition of 0.1 mol of NaOH will case the [HCN] to – 6

$OH^- + HCN(aq) = H_2O(l) + CN^-$

d) Addition of 0.2 mol of NaOH will case the pH to – 1

$OH^- + HCN$: Buffer capacity exceeded.

8.11 How do We Calculate the pH of a Buffer?

Purely for information purposes only ... the final formula is all you need.

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+ + A^- \qquad K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$[H_3O^+] = K_a\left(\frac{[HA]}{[A^-]}\right)$$

$$\log_{10}[H_3O^+] = \log_{10}K_a + \log_{10}[HA] - \log_{10}[A^-]$$

$$-\log_{10}[H_3O^+] = -\log_{10}K_a - \log_{10}[HA] + \log_{10}[A^-]$$

$$pH = pK_a + \log_{10}[A^-] - \log_{10}[HA]$$

$$pH = pK_a + \log_{10}\left(\frac{[A^-]}{[HA]}\right)$$

HA .. Weak acid ... Buffer Acid.
A^- .. Conj base ... Buffer Base.

$$\boxed{\text{Buffer pH:} \quad pH = pK_a + \log_{10}\left(\frac{[\text{Buffer Base}]}{[\text{Buffer Acid}]}\right)}$$

Henderson-Hasselbach Equation.

8.11 How do We Calculate the pH of a Buffer?

MAIN QUESTION

Question

A solution contains the following components

0.208 M HCO_2H

0.376 M $NaHCO_2$

What is the pH of the solution?

Ka HCO_2H = 1.8×10^{-4}

Answer

Enter a response, then Submit.

4

Buffer Acid : $[HCO_2H] = 0.208\,M$

Buffer Base : $[HCO_2^-] = 0.376\,M$

$$pH = pK_a + \log_{10}\left(\frac{[\text{Buffer Base}]}{[\text{Buffer Acid}]}\right)$$

$$pH = -\log_{10}(1.8\times10^{-4}) + \log_{10}\left(\frac{0.376}{0.208}\right)$$

$$\begin{aligned} pH &= 3.745 + \log_{10}(1.807) \\ &= 3.745 + 0.257 \\ &= 4.002 \end{aligned}$$

8.11 Buffers – A Summary

a) Buffer, Buffer acid + Buffer base	Weak acid + its conjugate base
b) [Buffer acid] = [Buffer base]	$pH = pKa$
c) Buffer Capacity	: [ ] of the Buffer acid or the Buffer base.
d) How a Buffer works	OH^- + Buffer acid = $H_2O(l)$ + Buffer base H_3O^+ + Buffer base = $H_2O(l)$ + Buffer acid
e) Buffer pH	$pH = pKa + \log_{10}\left(\frac{[\text{Buffer base}]}{[\text{Buffer acid}]}\right)$