

Suggested problems: (Chapter 19)

- 1-3, 5, 6, 11-20, 21-28, 33-36, 39-42, 44, 48,
- 50-54, 58, 62, 63, 65-78, 82, 86-89

What is a buffer sol'n?

A solution that contains an acid and its conjugate (both weak)

What does a buffer sol'n do?

Resists changes in pH when small quantities of a strong acid or strong base are added.

Why does a buffer sol'n resist changes in pH?

Because of the equilibrium established by the acid/base conjugate pair.

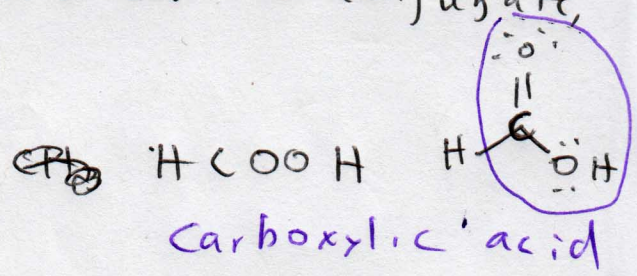
How much strong acid or base can the buffer resist?

Depends on quantity of the acid/base conjugate pair (buffer capacity)

What is an ideal buffer sol'n composition and the resulting pH?

Equal moles of the buffer acid and its conjugate, pH = pKa of weak acid

Formic acid $K_a = 1.8 \times 10^{-4}$



Given 0.50M HCOOH and

0.50M HCOONa, calculate pH

$$K_a = \frac{[H^+][HCOO^-]}{[HCOOH]}$$

	<chem>HCOOH</chem>	<chem>HCOO^-</chem>	<chem>H^+</chem>
I	0.5	0.5	0
C	-x	+x	+x
E	0.5-x	0.5+x	+x

The given concentrations of the reagents do not really exist in sol'n because the two reagents will try to establish an equilibrium.

However, in a weak acid/base conjugate system, the changes in concentration are often small enough that they can be ignored,

If $x \ll 0.5 \text{ M}$, then $0.5 - x \approx 0.5 = 0.5 + x$

$$K_a = \frac{[H^+][.5]}{[.5]} \Rightarrow K_a = [H^+]$$

$$pK_a = pH = \boxed{3.75}$$

When the changes in concentration ~~can't~~ due to establishing equilibrium can be ignored, the Henderson-Hasselbalch equation can be used.

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

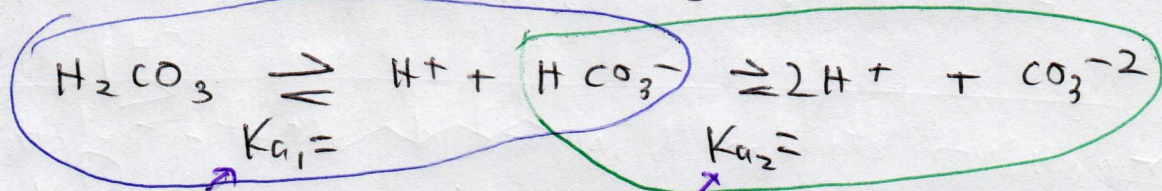
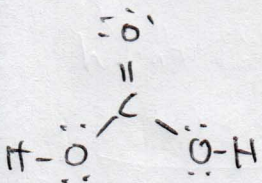
p. 835

pH 10.00

Na_2CO_3

how many grams?

1.5 L of 0.20 M NaHCO_3 to make the buffer?



for carbonic acid

for hydrogen carbonate

Since HCO_3^- and CO_3^{2-} are involved in this problem, K_{a2} is used, not K_{a1} . $K_{a2} = 4.7 \times 10^{-11}$

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

$$pK_{a2} = -\log_{10} K_a = 10.33$$

$$10.00 = 10.33 + \log \left(\frac{[A^-]}{[HA]} \right)$$

$$\log \left(\frac{[A^-]}{[HA]} \right) = -0.33$$

$$\frac{[A^-]}{[HA]} = 10^{-0.33} = 0.47$$

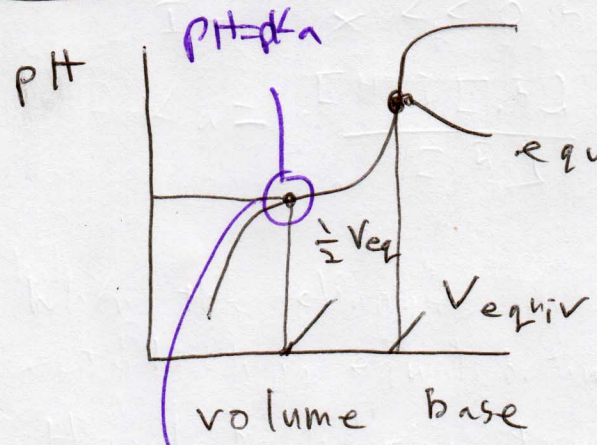
~~[A-]~~ [A-] = 0.47 x [HA]

mol A- = 0.47 x mol HA

= 0.47 x 1.5 L NaHCO3 x 0.2 M NaHCO3

= .14 mol CO3-2 x MM Na2CO3

→ grams needed



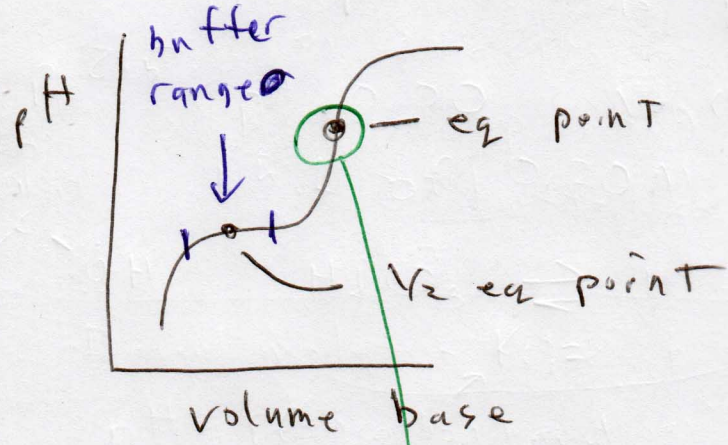
weak acid / strong base

equivalence point

(how is this different from endpoint?)

Assumed to be equal in buffer systems

$\frac{1}{2}$ equivalence point $\Rightarrow [HA] \approx [A^-]$



pH of buffer range = $pK_a \pm 1$

Indicator - The pH at which an indicator normally changes color is equal to pK_a of the indicator.

$\frac{1}{2}$ Equivalence point of indicator should ideally occur at equivalence point of the titration.