Key ✓

PRACTICE

1. For each of the solutions below, first identify if it is a buffer or not and, if it is a buffer, identify the acid and base component of the buffer system.

$H_2O + HClO_2 \rightleftharpoons H^+ + ClO_2^-$

a) equal volumes of equimolar solutions of $HClO_2$ and $KClO_2$ are mixed.

Buffer: YES NO weak Acid + weak Base

If yes, acid is $HClO_2$ and base is ClO_2^-.

b) 100 mL of 0.2 M NH_3 and 125 mL of 0.1 M NH_4Cl are mixed.

$NH_3 + NH_4Cl \rightleftharpoons NH_3 + H^+$

$\underset{Acid}{NH_4^+}(aq) \rightleftharpoons \underset{CB}{NH_3} + H^+$

Buffer: YES NO

If yes, acid is NH_4^+ and base is NH_3.

c) 50 mL of 0.1 M HCl and 25 mL of 0.2 M NaCl are mixed.

$HCl(aq) + NaCl(aq) \rightarrow HCl + NaCl$ All spectors NO Base

Buffer: YES NO

If yes, acid is ______ and base is ______.

d) equal volumes of 0.2 M $NaHSO_4$ and 0.25 M Na_2SO_4 are mixed.

$\cancel{Na}HSO_4 + \cancel{Na_2}SO_4 \rightarrow$

$HSO_4^-(aq) \rightleftharpoons H^+ + SO_4^{-2}$

Buffer: YES NO

If yes, acid is HSO_4^- and base is SO_4^{-2}.

e) 100 mL of 0.25 M HF and 75 mL of 0.40 M $Li_2C_2O_4$ are mixed. HF Strong Acid

Buffer: YES NO

If yes, acid is ______ and base is ______.

f) equal volumes of 0.1 M CH_3NH_2 and 0.2 M CH_3NH_3Br are mixed.

Buffer: YES NO

If yes, acid is NH_3 and base is NH_2^+.

$CH_3NH_2 + CH_3NH_3Br \rightleftharpoons$

$\underset{Acid}{CH_3NH_3} \rightleftharpoons \underset{CB}{CH_3NH_2^+} + H^+$

Key 2/

2. Find the pH of each of the following buffer solutions:

a) Find the pH of a 300 mL solution created by adding 4 moles of hypochlorous acid (HClO) to 8 moles of sodium hypochlorite (NaClO).

Reference Table

The acid is HClO and there are 4 mol or mmol $K_a = 2.9 \times 10^{-8}$

The base is ClO^- and there are 8 mol or mmol

$4\ HClO_{(aq)} + 8\ \cancel{Na}ClO_{(aq)} \rightleftharpoons 4H^+ + 8ClO^-$

Acid CB 300 mL

$[H^+] = K_a \frac{[Acid]}{[Base]}$ $pH = -\log[H^+]$

$= 2.9 \times 10^{-8} \left(\frac{4}{8}\right)$ $\boxed{pH = 7.8}$

$[H^+] = 1.5 \times 10^{-8}\ M$

b) Find the pH of a solution created by mixing 100 mL of 0.2 M lactic acid ($HC_3H_5O_3$) and 2.5 g of sodium lactate($NaC_3H_5O_3$).

The acid is $HC_3H_5O_3$ and there are 20 mol or mmol $K_a = 1.4 \times 10^{-4}$

The base is ______ and there are 22 mol or mmol

Instead of 1 mole

$HC_3H_5O_3 + \cancel{NaC_3H_5O_3} \rightleftharpoons H^+ + C_3H_5O_3^-$

Acid 2.5g CB

.2 M 2.5g

100 mL

$\left(\frac{2.5g\ NaC_3H_5O_3}{}\right)\left(\frac{1000\,mmol\ NaC_3H_5O_3}{112.03\,g\ NaC_3H_5O_3}\right)$

$[H^+] = K_a \frac{[Acid]}{[Base]}$

$HC_3H_5O_3$

$(.20\ M)(100\ mL)$ $= (1.4 \times 10^{-4})\left(\frac{20}{22}\right)$ $= 22\ mmol\ NaC_3H_5O_3$

20 mmol $[H^+] = 1.3 \times 10^{-4}\ M$

$pH = -\log[H^+]$

$\boxed{pH = 3.9}$

c) Find the pH of a solution created by mixing 1.7 g of ammonium chloride (NH_4Cl) with 25 mL of 1.0 M ammonia (NH_3).

The acid is NH_4^+ and there are 32 mol or mmol

The base is NH_3 and there are 25 mol or mmol $K_b = 1.8 \times 10^{-5}$

$NH_4\cancel{Cl}_{(aq)} + NH_3 \rightleftharpoons NH_4^+$ $NH_3\ (1.0M)(25\,mL) = 25\ mmol$

$NH_4^+ \rightleftharpoons NH_3 + H^+$ $NH_4Cl\left(\frac{1.7g}{}\right)\left(\frac{1000\,mmol}{53.492\,g}\right) = 32\ mmol$

Acid CB

1.7g 1.0M

25 mL $[H^+] = K_a \frac{Acid}{Base}$

$K_a K_b = 1 \times 10^{-14}$ $= 5.6 \times 10^{-10}\left(\frac{32}{25}\right)$

$K_a = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}}$

$K_a = 5.6 \times 10^{-10}$ $[H^+] = 7.2 \times 10^{-10}$ $pH = -\log[H^+]$

$\boxed{pH = 9.14}$

Key 3/

3. A buffer solution is created by mixing 200 mL of 0.2 M formic acid (HCO_2H) with 30 mmol of potassium formate ($KHCO_2$).

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

The acid is HCO_2H and there are 40 mol or mmol HCO_2H (.2M)(200mL) = 40mmol

The base is HCO_2^- and there are 30 mol or mmol $[H^+] = K_a \frac{[Acid]}{[Base]}$

a) Determine the pH of this solution.

$HCO_2H + KHCO_2 \rightarrow$ $HCO_2H \rightleftharpoons H^+ + HCO_2^-$

Acid .2M 200mL; CB 30 mmol

$= (1.8 \times 10^{-4})\left(\frac{40}{30}\right)$

$[H^+] = 2.4 \times 10^{-4}$

$pH = -\log[H^+]$

pH = 3.6

b) Write the net ionic equation for the reaction that occurs when hydrochloric acid is added to the solution. $HCO_2^- + H^+_{(aq)} \rightleftharpoons HCO_2H$

c) After the addition of 30 mL of 0.1 M HCl, 3 mmol of invader has been added. (30mL)(.10M) = 3 mmole

*d) Find the pH of the solution after the HCl was added.

$[H^+] = K_a \frac{[Acid + mole]}{[Base - mole]} = 1.8 \times 10^{-4} \frac{(40+3)}{(30-3)}$

$[H^+] = 2.9 \times 10^{-4}$

$pH = -\log[H^+]$

$\boxed{pH = 3.5}$

4. A buffer solution is created by mixing 200 mL of 0.1 M benzoic acid (C_6H_5COOH) with 4.5 g of sodium benzoate (C_6H_5COONa).

$K_a = 6.3 \times 10^{-5}$ $C_6H_5COOH \rightleftharpoons C_6H_5COO^- + H^+$ 200mL .1M; 4.5g

The acid is C_6H_5COOH and there are 20 mol or mmol

The base is $C_6H_5COO^-$ and there are 31 mol or mmol

a) Determine the pH of this solution.

$[C_6H_5COOH] = (200mL)(.1M) = 20$ mmol

$[C_6H_5COONa] \left(\frac{4.5g}{1}\right)\left(\frac{1000 \text{ mmol}}{144.1 g}\right) = 31$

$[H^+] = K_a \frac{Acid}{Base} = 6.3 \times 10^{-5}\left(\frac{20}{31}\right)$

$[H^+] = 4.1 \times 10^{-5}$

$pH = -\log[H^+]$

$\boxed{pH = 4.4}$

b) Write the net ionic equation for the reaction that occurs when sodium hydroxide is added to the solution. $C_6H_5COOH + OH^- \rightleftharpoons C_6H_5COO^- + H_2O$

c) After 0.01 moles of sodium hydroxide is added to the buffer 10 mmol of invader has been added

$(.01 \text{ moles})\left(\frac{1000 \text{ mmoles}}{1 \text{ mole}}\right) = 10$

*d) Find the pH of the solution after the sodium hydroxide was added.

$[H^+] = K_a \frac{Acid - mmol}{Base + mmol}$

$= 6.3 \times 10^{-5} \frac{(20-10)}{(31+10)}$

$[H^+] = 1.5 \times 10^{-5}$

$pH = -\log[H^+]$

$\boxed{pH = 4.8}$

Key 4

5. A buffer solution is created by mixing 100 mL of 0.1 M sodium bicarbonate ($NaHCO_3$) with 5.8 g of sodium carbonate (Na_2CO_3). $NaHCO_3 + Na_2CO_3$ $HCO_3^- \rightleftharpoons H^+ + CO_2^{-2}$

$K_a = 4.7\times10^{-11}$ The acid is HCO_3^- and there are 10 mol or mmol $[NaHCO_3]$ $(.1M)(100ml) = 10$ mmol

The base is CO_2^{-2} and there are 55 mol or mmol Na_2CO_3 $\left(\frac{5.8g}{1}\right)\left(\frac{1000\,mmol}{105.81g}\right) = 55$

a) Determine the pH of this solution.

$[H^+] = K_a \frac{[Acid]}{[Base]} = 4.7\times10^{-11}\left(\frac{10}{55}\right)$ $pH = -\log[H^+]$ $\boxed{pH = 11.1}$

$[H^+] = 8.5\times10^{-12}$

b) Write the net ionic equation for the reaction that occurs when hydrobromic acid is added to the solution. $H^+ + CO_2^{-2} \rightleftharpoons HCO_3^-$ HBr

c) After 8 mL of 0.2 M HBr is added to the buffer 1.6 mmol of invader has been added.

$[HBr] = (.2M)\,8ml = 1.6\,mmol$

*d) Find the pH of the solution

$[H^+] = K_a \frac{[Acid + mmol]}{[Base - mmol]}$ $pH = -\log[H^+]$ $\boxed{pH = 11}$

$= 4.7\times10^{-11}\frac{(10+1.6)}{(55-1.6)}$

$[H^+] = 1.0\times10^{-11}$

6. Consider a buffer solution created by 500 mL of 0.1 M hydrocyanic acid (HCN) and 2.0 g of sodium cyanide (NaCN). $HCN \rightleftharpoons H^+ + CN^-$ $[HCN]$ $(.1M)(500ml) = 50$ mmol

a) Determine the pH of the solution. 500 ml .1M Acid $K_a = 6.2\times10^{-10}$ 2.0g NaCN CB $[NaCN]$ $\left(\frac{2.0g}{1}\right)\left(\frac{1000\,mmol}{49.01g}\right) = 41$ mmol

$[H^+] = K_a \frac{(Acid)}{[Base]}$

$= 6.2\times10^{-10}\left(\frac{50}{41}\right)$ $pH = -\log[H^+]$

$[H^+] = 7.6\times10^{-10}$ $\boxed{pH = 9.1}$

b) What volume of 0.5 M HCl could be added to this solution before the buffering capacity ended?

Acid = 50 mmol HCl will combine w/ Base which has 41 mmol

Base = 41 mmol ∴ 41 mmol is MAX HCl that can be Added

$[HCl]$ $(.5M)(Xml) = 41\,mmol$ $\boxed{X = 82\,ml}$

c) What mass of sodium hydroxide could be added to this solution before the buffering capacity ended? ? g NaOH Buffer can Handle 50 mmol (Acid) of Base Being Added

$\left(\frac{50\,mmol}{1}\right)\left(\frac{39.998g\ NaOH}{1000\,mmole}\right) = 2.0\ g\ NaOH$

Key 5

7. Consider a buffer solution created by 500 mL of 0.2 M hydrocyanic acid and 4 g of sodium cyanide.

HCN NaCN

$HCN \rightleftharpoons H^+ + CN^-$

.2M 500mL $K_a = 6.2\times10^{-10}$ Acid; 4g CB

$[HCN] = (.2M)(500mL) = 100$ mmol

$[NaCN]\ \frac{4g}{}\left(\frac{1000 mmol}{49.01g}\right) = 82$ mmol

a) Determine the pH of the solution.

$[H^+] = K_a \frac{Acid}{Base} = (6.2\times10^{-10})\left(\frac{100}{82}\right)$

$[H^+] = 7.6\times10^{-10}$

$pH = -\log[H^+]$

$\boxed{pH = 9.1}$

b) Compare the pH of the original buffer solution in #6 and this buffer solution. Explain.

The pH values are the same because the ratio of acid to base is the same and the same conjugate pair is used.

8. Which buffer solution, #6 or #7, has a greater capacity for handling added HCl? Why?

#7 has a greater capacity for invading HCl because it has twice the conjugate base to react with it.

9. A buffer solution is created by mixing 2.8 g of sodium acetate ($NaC_2H_3O_2$) with 200 mL of 0.1 M acetic acid ($HC_2H_3O_2$).

$HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$

200mL .1M $K_a = 1.8\times10^{-5}$ Acid; 2.8g $NaC_2H_3O_2$ CB

a) Determine the pH of this solution.

$[H^+] = K_a \frac{[Acid]}{[Base]} = 1.8\times10^{-5}\left(\frac{20}{34}\right)$

$[H^+] = 1.1\times10^{-5}$

$pH = -\log[H^+]$

$\boxed{pH = 5.0}$

$[HC_2H_3O_2]\ (.1M)(200mL) = 20$ mmol

$[NaC_2H_3O_2]\ \left(\frac{2.8g}{1}\right)\left(\frac{1000 mmol}{82.034g}\right) = 34$ mmol

b) What effect will adding 50 mL of water to the buffer solution have on the pH? Explain.

Adding water will not affect the pH of the buffer because the ratio of acid to base components of the buffer remain the same.

10. Calculate the specified ratio of concentrations for each of the following to achieve the desired pH:

a) hydrofluoric acid (HF) to fluoride ion (F^-) in a buffer solution that has a pH of 3.2.

$K_a = 7.2\times10^{-4}$

$pH = -\log[H^+]$

$3.2 = -\log[H^+]$

$[H^+] = 6.3\times10^{-4}$

$[H^+] = K_a \frac{Acid}{Base}$

$6.3\times10^{-4} = 7.2\times10^{-4}\left[\frac{HF}{F^-}\right]$

$\left[\frac{HF}{F^-}\right] = 0.88$

b) propanoate ion ($C_3H_5O_2^-$) to propanoic acid ($HC_3H_5O_2$) in a buffer solution that has a pH of 5.2.

$K_a = 1.3\times10^{-5}$

$pH = -\log[H^+]$

$[H^+] = 6.3\times10^{-6}$

$[H^+] = K_a \frac{[Acid]}{[Base]}$

$6.3\times10^{-6} = 1.3\times10^{-5}\ \frac{[HC_3H_5O_2]}{[C_3H_5O_2^-]}$

$\frac{[HC_3H_5O_2]}{[C_3H_5O_2^-]} = 2.1$

c) hydrocyanic acid (HCN) to cyanide ion (CN-) in a buffer solution that has a pH of 9.1.

$K_a = 6.2\times10^{-10}$

$pH = -\log[H^+]$

$[H^+] = 7.9\times10^{-10}$

$[H^+] = K_a \frac{[Acid]}{[Base]}$

$7.9\times10^{-10} = 6.2\times10^{-10}\ \frac{[HCN]}{[CN^-]}$

$\frac{[HCN]}{[CN^-]} = 1.3$

$pH = -\log[H^+]$
put in Ka

Key 4

pH		
3.8	Lactic acid $HC_3H_5O_3$	Sodium lactate $NaC_3H_5O_3$
4.7	Acetic acid $HC_2H_3O_2$	Sodium acetate $NaC_2H_3O_2$
	Sodium bicarbonate $NaHCO_3$	Sodium carbonate Na_2CO_3
pOH = 4.7, pH = 9.3	Ammonia - Base NH_3	Ammonium chloride NH_4Cl

11. Using the table above, choose the acid component and base component you would use to create a buffer solution of the given pH. (No quantities are necessary.) Then calculate the pKa for the acid component of your choice using the previously provided reference table.

$pKa = -\log Ka$

a) pH = 3.8 $Ka = 1.4 \times 10^{-4}$

Acid lactic Acid Base $NaC_3H_5O_3$ pK_a 3.85 $pKa = -\log(1.4\times10^{-4})$

b) pH = 9.3

Acid Ammonium Chloride Base Ammonia pK_a 9.25

$Kb = 1.8\times10^{-5}$
$Ka = \frac{Kw}{Kb}$
$Ka = 5.6\times10^{-10}$
pKa =

c) pH = 10.4

Acid Sodium Bicarbonate Base Sodium Carbonate pK_a 10.33

HCO_3^-
$Ka = 4.7\times10^{-11}$
pH = 10.3

I picked it Because its exponate was close to pH 10.4

$pKa = -\log Ka$
=

* These (#12-14) are Finding PH of Buffer system Problems

Questions 12-14 refer to the following table of available chemicals.:

Ka	PH		
1.8×10^{-4}	0.1 M formic acid (HCO_2H) 3.7	sodium formate, $NaHCO_2$ (MW = 68.01g/mol)
1.3×10^{-5}	0.2 M propanoic acid ($HC_3H_5O_2$) 3.8	sodium propanoate, $NaC_3H_5O_2$ (MW = 96.06 g/mol)
4.4×10^{-7}	0.5 M carbonic acid (H_2CO_3) 6.36	sodium bicarbonate, $NaHCO_3$ (MW = 84.01 g/mol)
Kb 1.8×10^{-5}, Ka: 5.56×10^{-10}	0.4 M ammonia (NH_3) 9.26	ammonium chloride, NH_4Cl (MW = 53.49 g/mol)

12. Describe how to make 100 mL of a solution buffered at a pH of 4.3. Your description should include the identity of the acid and base component of your buffer as well as quantities of each chemical. So How much Acid + Base do I need? Conjugate in grams!

Choice of conjugate pair: Propanoic Acid (A) / Sodium propanoate (CB)

pH = 4.3

$pH = -\log[H^+]$

$[H^+] = 10^{-4.3}$

$[H^+] = 5.0 \times 10^{-5}$

$Ka = 1.3 \times 10^{-5}$

$HC_3H_5O_2 = .2M$

100 mL = .1 L

$[H^+] = Ka \frac{[Acid]}{[Base]}$

$[Base] = Ka \frac{[Acid]}{[H^+]}$

$NaC_3H_5O_2 = (1.3 \times 10^{-5}) \frac{.2 M}{5.0 \times 10^{-5}}$

$NaC_3H_5O_2 = .052 M$

$\left(\frac{.052 \text{ moles } NaC_3H_5O_2}{1L}\right)\left(\frac{.1 L}{1}\right)\left(\frac{96.06\ NaC_3H_5O_2}{1 \text{ mole } NaC_3H_5O_2}\right) = .50 g\ NaC_3H_5O_2$

$\boxed{\therefore \text{ Add } .50 g\ NaC_2H_5O_2 \text{ to } 100 ml \text{ of } .2M\ HC_3H_5O_2}$

13. Describe how to make 200 mL of a solution buffered at a pH of 8.7. Your description should include the identity of the acid and base component of your buffer as well as quantities of each chemical.

Choice of conjugate pair: Ammonia (B) / Ammonium Chloride (CA)

PH = 8.7

$pH = -\log[H^+]$

$[H^+] = 10^{-8.7}$

$[H^+] = 2.0 \times 10^{-9}$

$Kb = 1.8 \times 10^{-5}$

$Ka = \frac{Kw}{Kb}$

$Ka = 5.6 \times 10^{-10}$

200ml = .2 L

$[H^+] = Ka \frac{[Acid]}{[Base]} \rightarrow .4 M$

$[Acid] = \frac{[H^+][Base]}{Ka}$

$= \frac{(2.0 \times 10^{-9})(.4 M)}{5.6 \times 10^{-10}}$

$[NH_4Cl] = 1.4 M$

$\left(\frac{1.4 \text{ moles } NH_4Cl}{1L}\right)\left(\frac{.2L}{1}\right)\left(\frac{53.49 g\ NH_4Cl}{1 \text{ mole } NH_4Cl}\right) = 15 g\ NH_4Cl$

$\boxed{\therefore \text{ Add } 15 g\ NH_4Cl \text{ to } 200 ml \text{ of } 0.4 M\ NH_3 \text{ + Mix}}$

AP Chem - Unit 10 - WKst: Skill Builders - Buffers

14. How to make 50 ml soln Buffered at pH = 6.9

Carbonic Acid / Sodium Bicarbonate

Acid — CB

$K_a = 4.4 \times 10^{-7}$

Given:

$pH = 6.9$

$pH = -\log[H^+]$

$\therefore [H^+] = 1.3 \times 10^{-7}$

$50\,ml = .05\,L$

$[H_2CO_3] = .5\,M$

Soln.

$[H^+] = K_a \frac{[Acid]}{[Base]}$

$[Base] = \frac{K_a[Acid]}{[H^+]}$

$= \frac{(4.4 \times 10^{-7})(.5\,M)}{1.3 \times 10^{-7}}$

$[BASE] = 1.75\,M$

$\left(\frac{1.75 \text{ moles } NaHCO_3}{1L}\right)\left(\frac{.05\,L}{1}\right)\left(\frac{84.01\,g\ NaHCO_3}{1 \text{ mole } NaHCO_3}\right) = 7.3\,g\ NaHCO_3$

$\therefore$ Add 7.3 g $NaHCO_3$ to 50 ml of 0.4 M H_2CO_3 + mix

15) You have 100 ml of 0.2 M $HC_2H_3O_2$

Add NaOH (5 g Available) to make a Buffered soln pH = 4.9

** This is an strong Base Being Added "Invading" **

Given:

$pH = 4.9 \quad pH = -\log[H^+]$

$\therefore [H^+] = 1.3 \times 10^{-5}$

$K_a = 1.8 \times 10^{-5}$ Acetic Acid

Acetic Acid $(0.2\,M)(100\,ml) = 20\,mmol$

Soln.

$[H^+] = K_a \frac{[Acid]}{[BASE]}$ * Add BASE

$[H^+] = K_a \frac{[Acid - X]}{[Base + X]}$ X = mmol NaOH

$1.3 \times 10^{-5} = 1.8 \times 10^{-5} \frac{(20 - X)}{\emptyset + X}$ ∅ Base to start w/

$.72X = 20 - X$

$1.72X = 20$

$X = 11.6$ mmol NaOH

$\left(\frac{11.6 \text{ mmol NaOH}}{1}\right)\left(\frac{40.0\,g\ NaOH}{1000\,mmol}\right) = .46\,g$ NaOH

$\therefore$ Add .46 g NaOH to 100 ml of 0.2 M Acetic Acid

16) Given:

How to make 200 ml Soln Buffered pH = 3.6

Invader problem!!

200 mL of 0.1 M $HC_9H_7O_4$

$Ka = 3.4 \times 10^{-4}$

10 g $KOH_{(s)}$ strong Base!

Soln:

① $pH = -\log[H^+]$

$[H^+] = 10^{-3.6}$

$[H^+] = 2.5 \times 10^{-4}$

$[HC_9H_7O_4] = 200\text{ml}(0.1) = 20\text{ mmol}$

② $[H^+] = Ka \frac{[Acid]}{[Base]}$ Add Base KOH

$2.5 \times 10^{-4} = 3.4 \times 10^{-4} \frac{(20-X)}{0+X}$

$.74X = 20 - X$

$1.74X = 20$

$X = 11.5$ mmol KOH

③ $\left(\frac{11.5 \text{ mmol KOH}}{1}\right)\left(\frac{56.11\text{ g KOH}}{1000\text{ mmol KOH}}\right) = 0.65\text{ g KOH}$

∴ Add 0.65 g KOH to 200 ml of 0.1 M acetylsalicylic Acid

17) Given:

Invader Problem!!

100 ml of 0.5 M CH_3NH_2 weak Base

100 ml 0.5 M HCl strong Acid!

make soln Buffered @ pH = 10.2

Soln:

CH_3NH_2 $Kb = 4.4 \times 10^{-4}$

$Ka = \frac{Kw}{Kb} = \frac{1 \times 10^{-14}}{4.4 \times 10^{-4}}$

$Ka = 2.3 \times 10^{-11}$

$pH = -\log[H^+] = 10.2$

$[H^+] = 6.3 \times 10^{-11}$

$(100\text{ml})(.5\text{M}) = 50$ mmol CH_3NH_2 Base!!

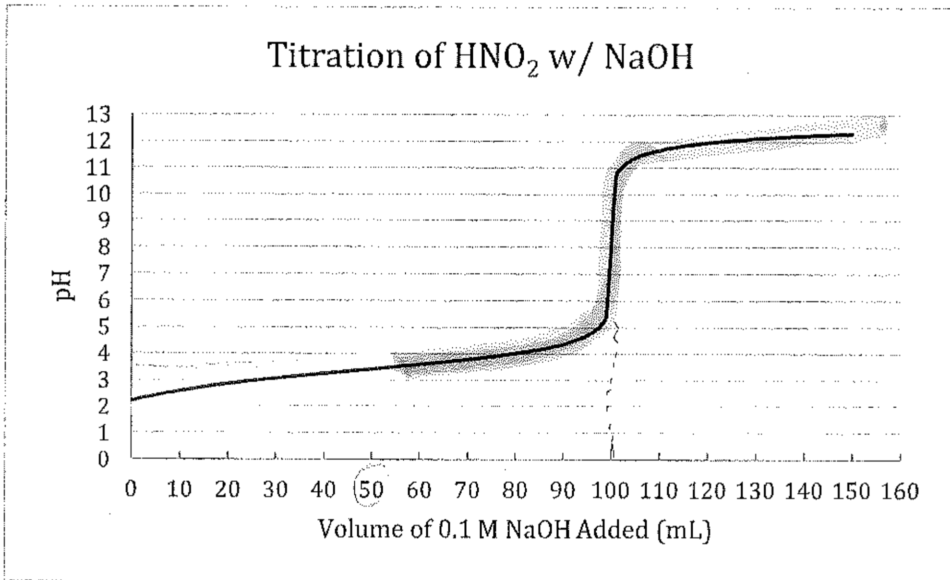
$[H^+] = Ka \frac{[Acid]}{[Base]}$ Add HCl $X =$ mmol HCl

$6.3 \times 10^{-11} = 2.3 \times 10^{-11} \frac{0+X}{50-X}$

$2.74 = \frac{X}{50-X}$

$X = 37$ mmol HCl

18. Consider the titration curve below that is created from adding 0.1 M NaOH to 100 mL of 0.1 M HNO_2.



a) Show the chemical reaction that occurs to result in a buffer solution when 20 mL of NaOH has been added.

$HNO_2 + NaOH \longrightarrow HOH + NaNO_2$

.1M .1M

100mL

10mmol

b) Using the curve, determine the K_a value of HNO_2. Show the buffer calculation that relates to this determination. Ka = ?

looking for the ½ equivalence pt → 50mL NaOH Added (Took 100mL/2)

pH ≈ 3.5

(50mL)(.1M) = 5 mmol Invader (NaOH)

$[H^+] = K_a \frac{[Acid]}{[Base]}$ Base Added

$[H^+] = 10^{-3.5}$

$[H^+] = 3.2 \times 10^{-4}$

$K_a = [H^+] \frac{[Base]}{[Acid]} = (3.2 \times 10^{-4}) \left(\frac{0+5}{10-5}\right)$

c) Using one color highlighter or pen, trace the values on the curve for which $[HNO_2] > [NO_2^-]$.

From Volumes .1 to 50mL

$K_a = 3.2 \times 10^{-4}$

d) Using another color highlighter or pen, trace the values on the curve for which $[NO_2^-] > [HNO_2]$.

From Volumes 50mL to 150mL

$$3.2\times10^{-4}=K_a\left[\frac{10^{-5}}{0+5}\right]$$ Acid Base

$K_a=3.2\times10^{-4}$

-or- $pH=pK_a$

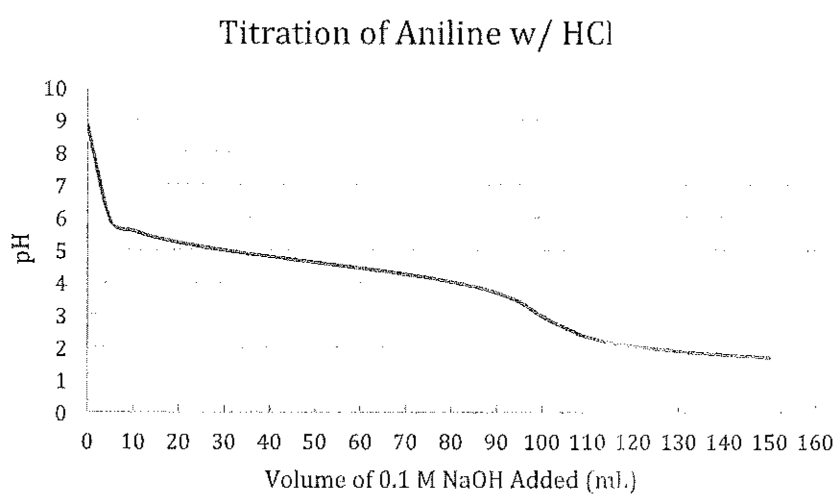
$K_a=10^{-pH}$

c) Using one color highlighter or pen, trace the values on the curve for which $[HNO_2]>[NO_2^-]$.
From volumes = 0.1 to 50 mL

d) Using another color highlighter or pen, trace the values on the curve for which $[NO_2^-]>[HNO_2]$.

From volumes 50 mL to 150 mL

19. Consider the titration curve that is created from adding 0.1 M HCl to 100 mL of 0.1 M aniline ($C_6H_5NH_2$).



a) Circle the region on the graph that represents a buffer solution present.

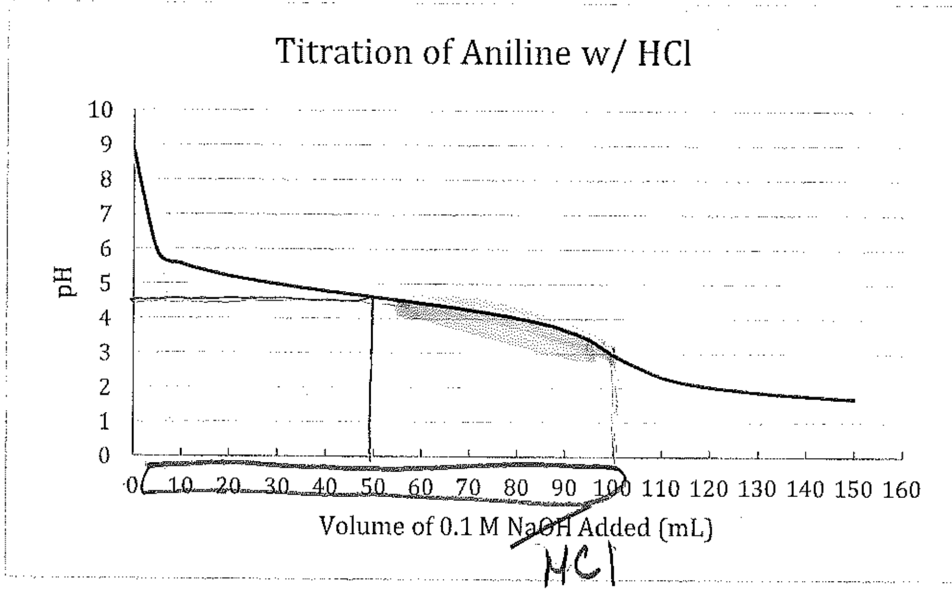
From 0.1 mL to the equivalence point at 100 mL of HCl

b) In the buffer region, using one color highlighter or pen, trace the values on the curve for which $[C_6H_5NH_2]>[C_6H_5NH_3^+]$.

From 0.1 mL to the half equivalence point at 50 mL of HCl

19. Consider the titration curve that is created from adding 0.1 M HCl to 100 mL of 0.1 M aniline ($C_6H_5NH_2$).

Strong Acid



a) Circle the region on the graph that represents a buffer solution present.

b) In the buffer region, using one color highlighter or pen, trace the values on the curve for which $[C_6H_5NH_2]>[C_6H_5NH_3^+]$. .1 ml to 100ml

c) In the buffer region, using one color highlighter or pen, trace the values on the curve for which $[C_6H_5NH_2]<[C_6H_5NH_3^+]$. 50 to 100ml

d) Estimate the value of K_b for aniline using the graph.

$K_a = 10^{-pH}$

$K_a = 10^{-4.5}$

$K_a = 2.0 \times 10^{-5}$

$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{2.0 \times 10^{-5}}$

$K_b = 5.0 \times 10^{-10}$

20. If the pH of a buffer solution of ethylamine ($CH_3CH_2NH_2$) and ethylammonium nitrate ($CH_3CH_2NH_3NO_3$) is 10.25, which component of the buffer, acid or base is present in higher quantity. Show your reasoning.

ethylamine $K_b = 5.6 \times 10^{-4}$

$K_a = \frac{K_w}{K_b} = \frac{1 \times 10^{-14}}{5.6 \times 10^{-4}}$

$K_a = 1.8 \times 10^{-11}$

$pH = -\log[H^+]$

$[H^+] = 10^{-10.25}$

$[H^+] = 5.6 \times 10^{-11}$

$[H^+] = K_a \frac{[Acid]}{[BASE]}$

$\frac{[Acid]}{[BASE]} = \frac{[H^+]}{K_a}$

$\frac{[Acid]}{[BASE]} = \frac{5.6 \times 10^{-11}}{1.8 \times 10^{-11}}$

$\frac{Acid}{Base} = 3.1$

∴ Acid is present in higher quantity

21. If the pH of a buffer solution of acetic acid ($HC_2H_3O_2$) and sodium acetate ($NaC_2H_3O_2$) is 5.5, which component of the buffer, acid or base is present in higher quantity. Show your reasoning.

$pH = 5.5$

$pH = -\log[H^+]$

$[H^+] = 10^{-5.5}$

$[H^+] = 3.2 \times 10^{-6}$

$K_a = 1.8 \times 10^{-5}$

$[H^+] = K_a \frac{Acid}{Base}$

$\frac{Acid}{BASE} = \frac{[H^+]}{K_a} = \frac{3.2 \times 10^{-6}}{1.8 \times 10^{-5}}$

$\frac{Acid}{BASE} = .18$

∴ Base is present in higher quantity

REFLECTION QUESTIONS

1. **When calculating the pH of a buffer system, what quantities do you need to have?**
   You need the K_a value for the acidic component of the buffer along with moles or molarity values for the acidic and basic components of the system.

2. **Why is it important to be working with moles or millimoles when dealing with an invader in a buffer?**
   These problems are actually chemical reactions happening and quantities of change of the acid and base part of the buffer must be calculated in moles.

3. **When dealing with a buffer invader problem, what other chemistry problems are these similar to?**
   These are similar to limiting reactant problems.

4. **Why does the pH of a buffer system not change significantly when invaded by a small quantity of strong acid or base?**
   Because the invading strong acid or base is converted by one component of the buffer into the other component of the buffer. Thus, there is not a large change in free H^+ or OH^- ions which are responsible for pH changes.

5. **Why is it not necessary to use molarity values when calculating $[H^+]$ for a buffer but molarity values must be used when dealing with a simple weak acid equilibrium expression?**
   Buffer calculations have a ratio of acid to base and the volume values used in molarity calculations cancel out in the calculation. In an equilibrium expression, these volumes do not cancel out and molarity must be used.

6. **Why does the pH = pKa at the ½ equivalence point in the titration?**
   At the ½ equivalence point, there are equivalent moles of acid and base component in the buffer. Therefore the ratio of acid to base is 1 and $[H^+] = K_a$. Therefore, $pH = pK_a$.

7. **Look back at problems 2(b) and 2(c). What did you have to do differently in the problems?**
   The K_a of the acidic component of the buffer was given directly for 2(b), but only the K_b for the basic component was given in 2(c). Therefore, a K_a value had to be calculated in 2(c) in order to calculate the pH.

8. **When creating a buffer, what is the most important factor in the pH of the system?**
   The value of the Ka is the most important factor to setting the pH. You want the pK_a to be as close as possible to the desired pH.

9. **When creating a buffer, what is the most important factor in the capacity of the system for handling invading acid or base?**

   The quantity of acid and base components in the buffer determines the capacity of the system for handling invaders.