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PLEASE NOTE:If ANY of the above information is UNCLEAR or not provided, your grade will NOT be recorded!!

Lab Day (T/W/Th/F):

Lab Week (even/odd):

Lab time (10:00, 2:30, 6:30):

Laboratory Report Form

Experiment 2.

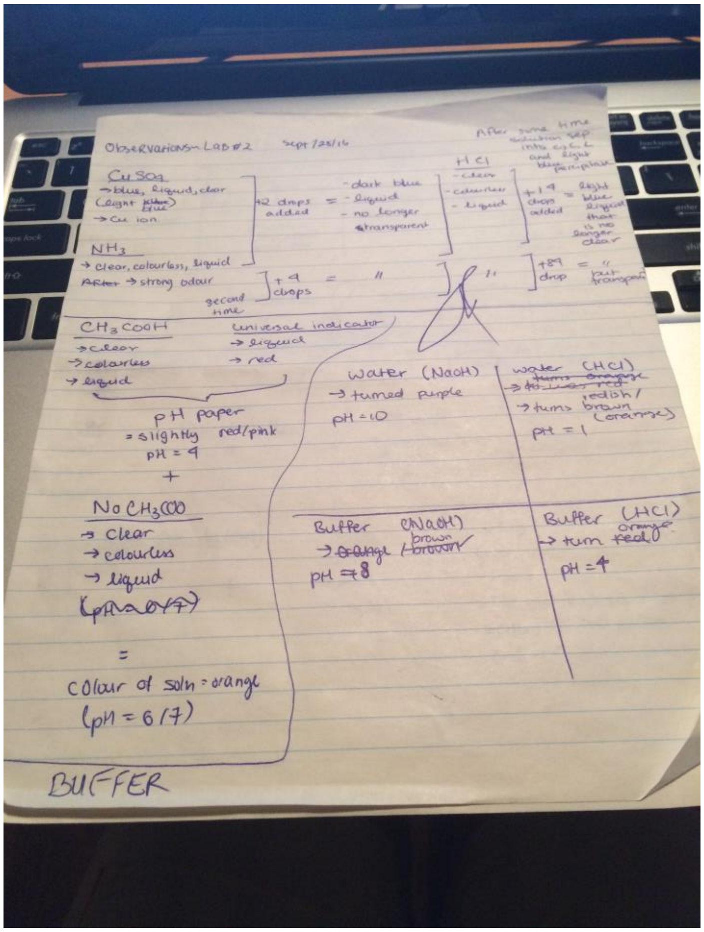
Equilibria

Checklist:

- **Raw Data Sheet written in pen, signed by TA and attached**
- **Completed Report Form attached**

Student's Initials _JS_

Raw Data


Observations - Lab #2 Sept /28/16

$CuSO_4$
→ blue, liquid, clear (light blue)
→ Cu ion

+2 drops added = - dark blue - liquid - no longer transparent

After some time solution sep into c,c,L and light blue precipitate

HCl
- clear
- colourless
- liquid

+14 drops added = light blue liquid that is no longer clear

NH_3
→ clear, colourless, liquid
→ strong odour

+9 drops = " second time

+89 drop = " but transparent

CH_3COOH
→ clear
→ colourless
→ liquid

universal indicator
→ liquid
→ red

pH paper
= slightly red/pink
pH = 4

+

$NaCH_3COO$
→ clear
→ colourless
→ liquid

=

colour of soln = orange
(pH = 6/7)

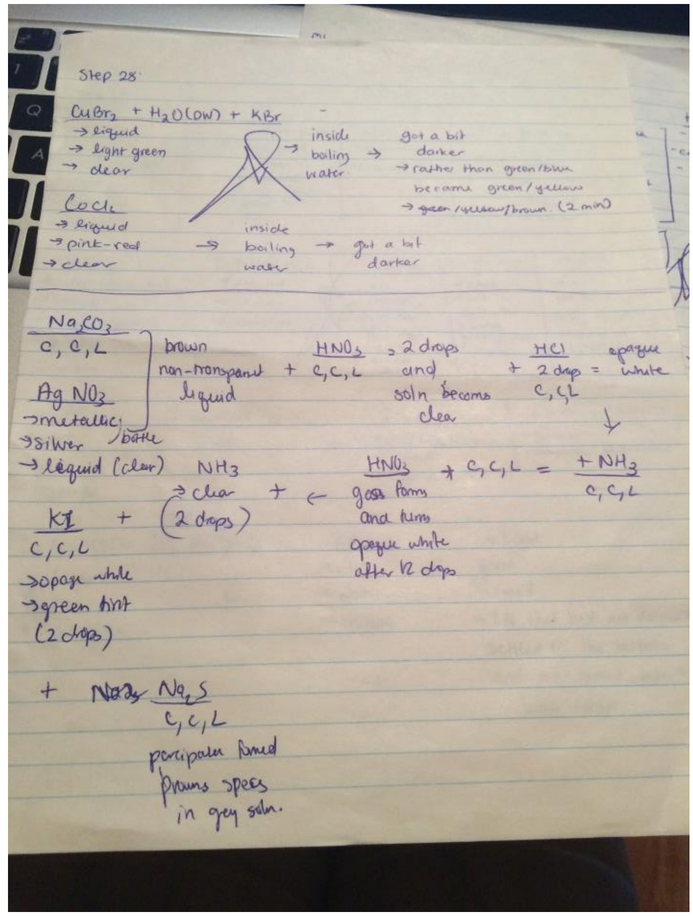
BUFFER

water (NaOH)
→ turned purple
pH = 10

water (HCl)
→ turns reddish/brown (orange)
pH = 1

Buffer (NaOH)
→ brown
pH = 8

Buffer (HCl)
→ turn orange
pH = 4


Step 28

$CuBr_2 + H_2O(DW) + KBr$
→ liquid
→ light green
→ clear

→ inside boiling water → got a bit darker
→ rather than green/blue became green/yellow
→ green/yellow/brown (2 min)

$CoCl_2$
→ liquid
→ pink-red
→ clear

→ inside boiling water → got a bit darker

Na_2CO_3
c, c, L

brown non-transparent liquid

+ HNO_3 c, c, L = 2 drops and soln becomes clear

+ HCl 2 drop c, c, L = opaque white

$AgNO_3$
→ metallic/bottle
→ silver
→ liquid (clear)

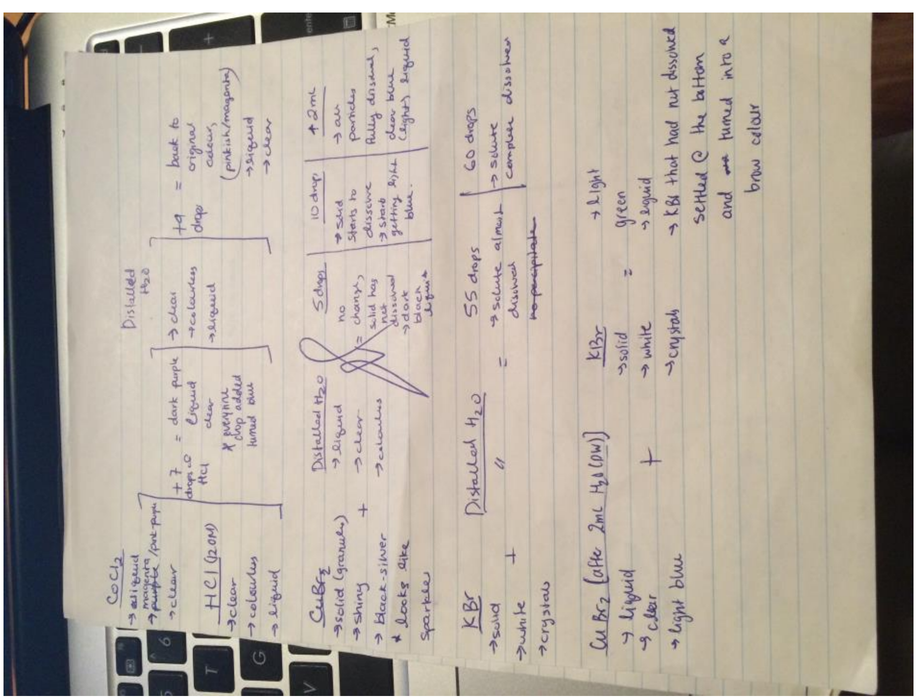
NH_3
→ clear
(2 drops)

HNO_3 + c, c, L = gas forms and turns opaque white after 12 drops

+ NH_3 c, c, L

KI + c, c, L
→ opaque white
→ green tint
(2 drops)

+ Na_2S
c, c, L
precipitate formed
browns specs in grey soln.


$CoCl_2$
→ liquid
→ magenta purple / pink-purple
→ clear

+7 drops of HCl = dark purple liquid clear
✗ purple one drop added turned blue

HCl (12M)
→ clear
→ colourless
→ liquid

Distilled H_2O
→ clear
→ colourless
→ liquid

+9 drops = back to original colour (pinkish/magenta)
→ liquid
→ clear

$CuBr_2$
→ solid (granules)
→ shiny
→ black-silver
✗ looks like sparkles

+ Distilled H_2O
→ liquid
→ clear
→ colourless

= 5 drops: no change, solid has not dissolved → dark black liquid

10 drops → solid starts to dissolve → starts getting light blue

+2mL → all particles fully dissolved, clear blue (light) liquid

KBr
→ solid
→ white
→ crystals

+ Distilled H_2O "

= 55 drops → solute almost dissolved, no precipitate

60 drops → solute completely dissolves

$CuBr_2$ (after 2mL H_2O (DW))
→ liquid
→ clear
→ light blue

+ KBr
→ solid
→ white
→ crystals

= → light green
→ liquid
→ KBr that had not dissolved settled @ the bottom and turned into a brown colour

Table 1. Observations and Discussion

$$[Cu(H_2O)_4]^{2+} (aq) + 4\ NH_3 (aq) \approx [Cu(NH_3)_4]^{2+} (aq) + 4\ H_2O (l)$$

Reactants:
$CuSO_4$ = liquid, light blue, transparent/clear (Cu ion in charge of blue colour)
NH_3 = liquid, clear, transparent, colourless
HCl = liquid, clear, transparent, colourless

a. The colour of the $CuSO_4$ solution is light blue and the ion responsible for this is the metal ion Copper.
b. The NH_3 took 2 drops to change the solution of the $CuSO_4$ into a darker, richer blue. It remains as a liquid that is transparent. Here the NH_3 is a weak base. When added to the $CuSO_4$ solution, the water is replaced with the NH_3 ion. The $[Cu(NH_3)_4]^{2+}$ solution is dark blue in colour not because of the Cu ion, but the complex its self.
c. The HCl took 14 drops to change the solution back to the original colour of $CuSO_4$.

$$Cu(NH_3)_4 (aq) + HCl (aq) = NH_4 (aq) + CuCl_{2 (S)}$$
$$H^+ (aq) + NH_3 (aq) \approx NH_4^+ (aq)$$

However, it was noticed that the solution was murky, almost opaque and after sometime, the solution separated into blue precipitate and clear liquid.

1. The changes that were observed in step 4 of the reaction were expected since the reaction was put through the same motions of shifting from the left side to the right. The only difference that was noted when doing this same steps, the second time were that the number of drops increased from 2 to 4 for NH_3 and from 14 to 84 drops in HCl.

Analysis:

It should be noted that the main principle at work here to explain the constant shift between reactants and products would be LeChatelier's principle. In order for the reaction above to maintain an equilibrium, the reaction must shift when new reactants are added to the equation. Thus, for the addition of the NH_3 solution, the reaction shifted to the right to accommodate the amount of NH_3 reacted with the Cu ion to create the rich blue solution which was $Cu(NH_3)_4$ (aq). This changed when HCl was added since, here the acid reacted with the base NH_3, turning it into ammonium. In order to compensate for the loss of NH_3 molecules on the left side, and to relieve the stress caused by the removal of NH_3, the reaction shifts left.

The following 7 reactions are for part 2, multiple equilibrium. Discussion is embedded within the observations.

$$2AgNO_3 (aq) + Na_2CO_3 (aq) \approx Ag_2CO_3 (s) + 2\ NaNO_3 (aq)$$

Reactants:
$AgNO_3$ = clear, colourless, liquid
Na_2CO_3 = clear, colourless, liquid

d) Na_2CO_3 is a clear colourless liquid, in which both ions are responsible for non-existent colour.

e) This reaction is a double displacement reaction in which the Ag+ ion swaps places with Na+. This reaction makes a brown precipitate (Ag_2CO_3 (s)) and liquid with the colour of a brownish grey $NaNO_3$ (aq). It is an opaque solution. According to LeChatelier's principle, a dynamic equilibrium has been reached as reaction happens back and forth. There are no spectator ions.

$$2\ H^{+}\ (aq) + CO_3^{2-}\ (aq) \approx H_2CO_3\ (aq) \approx H_2O\ (l) + CO2\ (g)$$

Reactants:
HNO_3 = clear, colourless, liquid

f) Here 2 drops of HNO_3 are added, however NO_3^- is a spectator ion which currently does not affect the outcome of the products. The H+ ion replaces the Ag+ ion which, like the NO_3^- it is a spectator ion. It should be noted that the solution had become clear. This shows that the H_2CO_3 further dissociated into water and carbon dioxide gas. The carbonic acid, is unstable at room temperature and pressure thus decomposing into water and carbon dioxide. The addition of the HNO_3 causes the equilibrium to shift to the right in order to accommodate for the need of CO_3^{2} ions, so that it can react with the HNO_3.

$$Ag^{+}\ (aq) + Cl^{-}\ (aq) \approx AgCl\ (s)$$

Reactants:
HCl = liquid, clear, transparent, colourless

g) The spectator ions in the previous reaction are now used here to form a precipitate, AgCl. With 2 drops of HCl, we see an opaque white solution appear, and thus we are aware that the precipitate has formed. With reference to LeChatelier's principle, we see that the reaction shifts right since the once spectator ion, is the ion that is being combined to produce the precipitate in the solution. It should be noted that the H+ ions are spectators at the moment.

$$Ag^{+}\ (aq) + 2NH_3\ (aq) \approx [Ag(NH_3)_2]^{+}\ (aq)$$

Reactants: NH_3 = liquid, colourless, clear h) Here the solution is once again turned into a clear solution. The NH_3 is the strong ion that replaces the original anion (Cl) with itself. This makes the precipitate dissociate back into ions where Ag+ is now part of the $Ag(NH_3)_2$ complex. This complex gives the non-existent colour in the solution. With reference to LeChatelier's principle, the reaction shifts left since equilibrium needs to be satisfied. This can only be done when there is an abundant amount of Ag+ ions present so that the $Ag(NH_3)_2$ complex can be formed. Thus, the reaction shifts left so that the precipitate dissociates into ions.
$H^+ (aq) + NH_3 (aq) \approx NH_4^+ (aq)$ Reactants: HNO_3= liquid, colorless, clear NH_3 = liquid, colorless, clear - With the addition of the HNO3, the original $Ag(NH_3)_2$ complex is broken apart so that the NH3 could form NH_4, thus allowing Ag ions to become precipitate with the Cl- spectator ions. - With the addition of NH_3, the AgCl compound once again is dissociated and the $Ag(NH_3)_2$ complex is reformed again, thus making the solution clear. - The reaction shifts left when the NH_3 so added causing the formation of NH_4, and $Ag(NH_3)_2$ complex to dissociate.
$Ag^+ (aq) + I^- (aq) \approx AgI (s)$ Reactants: KI = solid, white, crystal/powder i) With the addition of KI, the solution returns to its opaque white, with a slight greenish tint. The formation of a precipitate is from the I- anion binding with the silver producing AgI. Relating back to Le Chatelier's Principle, as the number of reactants increases, the system shifts forward to compensate and thus more products are produced.
$2Ag^+ (aq) + S^{2-} (aq) \approx Ag_2S (s)$ Reactants: Na_2S = colourless, transparent

j) With the addition of Na_2S, the solution turns gray in color. Precipitate in the color of brown specs is formed from the production of Ag_2S. According to Le Chatalier's principle, the equilibrium shifts left causing the AgI to dissolve and later form into Ag_2S.

$$CH_3COOH\ (aq)\ +\ H_2O\ (l)\ \approx\ H_3O^+\ (aq)\ +\ CH_3COO^-\ (aq)$$

k) CH3COOH, a clear colorless liquid turns red when the universal indicator is added. This color represents a pH of 0 (very acidic).
l) The pH paper indicates a pH of 4.
m) The pH changes to 7 with the addition of the salt, thus making the solution a clear orange solution.
n) Both the universal indicator and the pH read the solution is neutral.
o) The universal indicator in distilled water is red, this corresponds with a pH of 0.
p) pH decreases to 1 when HCl is added to the water, and the color of the solution was red. When added to the buffer solution, pH decreases to 4, making the color orange.
q) The pH and the color coincide for the water. However, pH is slightly lower than expected for the more neutral color given off by the buffer.
r) pH of the water increases 10 when the base is added to the water, giving it a purple color. When added to the buffer, the pH increases to 8, as well, giving it a brown color.
s) The pH of the buffer solution generally coincided with the color of the solution. The pH of the water coincided with the color of the water.

Analysis:
Here two systems are being compared with against an acid and a base. We are able to see the different outcomes from both the water and the buffer system when acid and base are added. However, the reasoning behind why the reactions occur the way they do is because of equilibrium which can be explained through LeChatelier's principle. We know that in order for these reactions to maintain balance they need to shift left or right to accommodate for certain products/reactants. We see this in the water wells, as the pH jumps from neutral to both acidic and basic pH levels. However, this is not the same for the buffers which had minimal changes to the pH. This proves that the buffer was able to maintain equilibrium.
The reason why the buffer's pH doesn't change is because the acid adds H+ ions to bind to CH3COO- so that it produces CH3COOH. When a base is added CH3COOH loses an H+ ion thus producing H2O which also doesn't affect the pH. Overall, there the weak acid and bases and their conjugates do not affect the pH to the same extent as these acids/bases in water.

$$4Cl^-\ (aq)\ +\ [Co(H_2O)_6]^{2+}\ (aq)\ \approx\ [CoCl_4]^{2-}\ (aq)\ +\ 6\ H_2O\ (l)$$

Reactants:
CoCl2 = liquid, clear, pink
HCl = liquid, clear, colorless
CuBr2= solid, shiny crystal, metallic
t) At room temperature the CoCl2 solution is a liquid that's magenta/pink in color and transparent.
u) The solution changes color and becomes a transparent dark purple. The chemistry that explains this is that adding HCL, increasing the amount of CL- ions and thus results in a

equilibrium shift right to balance the products with the excess reactants. The increase in CoCl4 makes the solution a strong violet color.
v) The addition of water changes the equilibrium because it adds more products, which causes a shift in the reverse direction. The shift, resulting in an increase of reactants, returns the solution to its original pink color.
w) CuBr2 is a shiny, crystal solid, black.
x) With 5 drops of H_2O added, the solid is barely dissolved.
y) When an additional 10 drops are added, the solid starts to dissolve to a light blue solution
z) After making the solution a volume of 2ml, the entire sold dissolves showcasing a blue solution.
aa) The dissolved crystals, which were white colored, turned brown when in the solution.
bb) When the CuBr2 was added to the solution, the result was a light green liquid with KBr left at the bottom and tuned into a dark brown color.

Analysis:
With reference to the Le Chatelier's principle, it is to be noted that the reaction shifted right during the addition of Cl- ions, but this changes in the opposite direction when H2O is added to the solution. It should be taken into consideration that the common ion effect is in play as there are many of the same ions added into the equation thus causing the equilibrium to shift due to change in anion concentration.

dd) The original solution of CoCl2 began as a pinkish/red. When the solution was heated, the solution turned into a darker color, like a magenta.

Analysis:
For the second solution, the dark magenta color occurs when the solution is exposed to high heat. However, it is to be noted that the change in color did not become blue which is known to be the end result of this heated reaction. Thus an error had occurred since the reaction did not fully shift to the left due to the fact that the amount of time given to the reaction in the boiling water was enough.

$$4\ Br^-\ (aq) + [Cu(H_2O)_4]^{2+}\ (aq) \approx [CuBr_4]^{2-}\ (aq) + 4\ H_2O\ (l)$$

Reactants:
Solution of KBr, H_2O, and

cc) The original solution of the three compounds, CuBr2, H2O and KBr produced a light green, liquid that is clear. When heated, the color of the solution changes from, the light green to a yellow green, which slowly proceed to be a dark brown color.

Analysis:
The first solution was heated the color changed from a green to a brown. Cu(H2O)4 ion was responsible for the original light green color. With the addition of heat to the reaction, the equilibrium to shift towards the CuBr4 ion's causes it to turn brown in color. The reaction is endothermic, as added energy favors the reactants, making the reaction shift to right.

Calculations:

Find the pH of the water solutions and different buffer systems.

1. pH of water
2. pH of buffer
3. pH of water + HCl
4. PH of water + NaOH
5. pH of buffer + HCl
6. pH of buffer + NaOH

1. pH of water = 7.0

2. pH of buffer (CH_3COOH + $NaCH_3COO$)

$CH_3COOH\ (aq) + H_2O\ (l) \approx H_3O^+\ (aq) + CH_3COO^-\ (aq)$

C = 0.1 mol/L
V = 10 drops
V_T = 20 drops

$$C = 0.1\frac{mol}{L} \times \frac{10\ drops}{20\ drops}$$
$$= 0.05\ mol/L$$

C = 0.1 mol/L
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$$C = 0.1\frac{mol}{L} \times \frac{10\ drops}{20\ drops}$$
$$= 0.05\ mol/L$$

Henderson – Hassel Balch Eqn

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$
$$pH = -\log K_a + \log\frac{[A^-]}{[HA]}$$
$$= -\log(1.8\times10^{-5}) + \log\frac{[0.05]}{[0.05]}$$
$$= -\log(1.8\times10^{-5})$$
$$= 4.74 \text{ (buffer pH)}$$

3. pH of water + HCl

V_{HCl} = 5 drops
C_{HCl} = 0.1 mol/L

V_{H2o} = 20 drops

$$V_f = V_{H2O} + V_{HCl}$$
$$V_f = 20\ drops + 5\ drops$$

$$V_iC_i = V_fC_f$$
$$(5\ drops)\left(\frac{0.1mol}{L}\right) = (25\ drops)\ (X)$$
$$\frac{(5\ drops)}{(25\ drops)}\left(\frac{0.1mol}{L}\right) = (X)$$
$$(X) = 0.02\ mol/L$$

**X is the concentration of H+ ions

$$pH = -\log(H^+)$$
$$pH = -\log(0.02mol/L)$$
$$pH = 1.70$$

4. pH of water + NaOH

$$V_f = V_{H2O} + V_{NaOH}$$
$$V_f = 20\ drops + 5\ drops$$
$$V_iC_i = V_fC_f$$
$$(5\ drops)\left(\frac{0.1mol}{L}\right) = (25\ drops)\ (X)$$
$$\frac{(5\ drops)}{(25\ drops)}\left(\frac{0.1mol}{L}\right) = (X)$$
$$(X) = 0.02\ mol/L$$

**X is the concentration of OH- ions

$$pOH = -\log(OH^-)$$
$$pOH = -\log(0.02mol/L)$$
$$pOH = 1.70$$

$$pH = -\log(H^+)$$
$$pH = 14 - pOH$$
$$pH = 12.3$$

5. pH of buffer + HCl

CH_3COO^- (aq) + HCl ≈ CH_3COOH + Cl^-

CH_3COO^- (aq)	HCl	CH_3COOH
C = 0.1 mol/L	C = 0.1 mol/L	C = 0.1 mol/L
V = 10 drops	V = 5 drops	V = 10 drops
V_T = 25 drops	V_T = 25 drops	V_T = 25 drops

$CH_3COO^- = \left(0.1\frac{mol}{L}\right)\left(\frac{10\ drops}{25\ drops}\right)$

$CH_3COO^- = 0.04mol/L$

$HCl = \left(0.1\frac{mol}{L}\right)\left(\frac{5\ drops}{25\ drops}\right)$

$HCl = 0.02mol/L$

$CH_3COOH = \left(0.1\frac{mol}{L}\right)\left(\frac{10\ drops}{25\ drops}\right)$

$CH_3COOH = 0.04mol/L$

CH_3COO^- (aq) + HCl ≈ CH_3COOH + Cl

I	0.04 mol/L	0.02mol/L	0.04mol/L
C	-0.02mol/L	-0.02mol/L	+0.02mol/L
E	0.02mol/L	0 mol/L	0.06mol/L

Henderson – Hassel Balch Eqn

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

$$pH = 4.74 + \log\frac{[0.02]}{[0.06]}$$

pH = 4.74 – 0.48
= 4.26

6. pH of buffer + NaOH

CH_3COOH (aq) + NaOH ≈ $NaCH_3COO$ + H_2O

CH_3COOH	NaOH	$NaCH_3COO$
C = 0.1 mol/L V = 10 drops V_T = 25 drops	C = 0.1 mol/L V = 5 drops V_T = 25 drops	C = 0.1 mol/L V = 5 drops V_T = 25 drops

$CH_3COOH = \left(0.1\frac{mol}{L}\right)\left(\frac{10\ drops}{25\ drops}\right)$

$CH_3COOH = 0.04mol/L$

$NaOH = \left(0.1\frac{mol}{L}\right)\left(\frac{5\ drops}{25\ drops}\right)$

$NaOH = 0.02mol/L$

$NaCH_3COO = \left(0.1\frac{mol}{L}\right)\left(\frac{10\ drops}{25\ drops}\right)$

$NaCH_3COO = 0.04mol/L$

CH_3COOH (aq) + NaOH ≈ $NaCH_3COOH$ + H_2O

I	0.04 mol/L	0.02mol/L	0.04mol/L
C	-0.02mol/L	-0.02mol/L	+0.02mol/L
E	0.02mol/L	0 mol/L	0.06mol/L

Henderson – Hassel Balch Eqn

$$pH = pK_a + \log\frac{[HA]}{[A-]}$$
$$pH = 4.74 + \log\frac{[0.06]}{[0.02]}$$
$$pH = 5.22$$

Difference in pH

Buffer
pH = 5.22 – 4.26
= 0.96

Water
pH = 12.3 -1.7
= 11.6

Thus, it should be noted that the difference in pH was minimal in the buffer system where as in the water wells, the difference of pH was a lot. This shows that the buffer system was able to maintain equilibrium, as mentioned in part 4 of observations.

Conclusion:

To summarize, the equilibrium of a reaction can be effected by the addition of various chemicals, temperature, and by the reactants/products in a reaction. We were able to explore this through pH levels of buffer solution, colour change of endo and exothermic reactions, as well as through LeChatelier's principle. This principle shows that reactions will adjust certain circumstances in order to maintain the equilibrium thus causing reactions to shift left, right or to remain the same.