Student Name: Anjali Thomas

Student Number: 8581091

Partner's Name and Student #: Julia Hancott

Demonstrator's Name: Joel

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

Lab Week (even/odd): odd

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

Laboratory Report Form

Experiment 2.

Equilibria

Checklist:

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

Table 1. Observations and Discussion

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

Initial : CuSO4 – light blue, transparent solution
NH3 – clear liquid with strong smell
HCl – clear liquid with strong acidic smell

a) In the experiment, the CuSO4 was in an aqueous form thus, the ions in CuSO4 disassociate. In the equation above, the Cu ion forms an ion-water complex of [Cu(H2O)4]2+. This solution has been observed to be a transparent, light blue solution.

b) The reaction between [Cu(H2O)4]2+ and NH3 created an immediate reaction of a transparent, dark blue solution. In this reaction, a single displacement reaction with [Cu(H2O)4]2+ and NH3 occur. The Cu ion reacts with the NH3 ion. This produces the products listed above in the equation.

c) As HCl was added, the colour of the solution became lighter until it was a shade of light blue. The HCl present in the aqueous solution initially dissociates to from H+ and Cl-. The H+ reacted with NH3 in the reactants side to produce NH4. This ultimately caused the NH3 concentration to decrease which then caused an equilibrium shift towards the reactants. Therefore, an increased amount of [Cu(H2O)4]2+ formed which created the light blue colour seen initially in just [Cu(H2O)4]2+ alone.

1. By repeating steps 2 and 3, it demonstrated the equilibrium shift favouring the products with the addition of NH3. This is seen by the solution turning a dark blue once again. The addition of NH3 causes an influx in reactants. The equilibrium then shifts to the products side to obey Le Chatelier's Principle of opposing the change in a reaction with an opposite reaction. In step 3, the equilibrium shift occurs once to the reactants side and this is indicated by the light blue solution colour.

2AgNO3 (aq) + Na2CO3 (aq) → Ag2CO3 (s) + 2NaNO3 (aq)
←

Initials: Na2CO3 – clear , no odour
AgNO3 – clear liquid, no odour

d) The colours that were observed was a clear liquid. Na2CO3 disassociates in water to produce Na^1+ and CO3^2-. These ions caused the solution to appear clear.

e) As AgNO3 is added to the Na2NO3, the solution turned orange tinted at first then beige and opaque. A double displacement reaction is occurring and the change in colour indicates the presence of products. The opaque colour was the Ag ion and CO3 ion reacting to form a precipitate.

$$2\,H^{+}\,(aq) + CO_3^{2-}\,(aq) \rightleftharpoons H_2CO_3\,(aq) \rightleftharpoons H_2O\,(l) + CO_2\,(g)$$

f) The addition of HNO3 initially created a precipitate and then transitioned into a clear solution with fog being visible throughout the solution. HNO3 is a strong acid therefore, it dissociates completely when reacted with water. Therefore, since HNO3 is in an aqueous solution, it dissociates. The H+ ion reacted with CO3 2- ion from the products of the previous reaction. A reaction occurred to balance the influx of reactants which resulted in the productions of the products. Since this is a 3 part equilibria, all three reactions must have equal rate of reactions. According to the observations that we made, there was no noticeable gas that had formed therefore, the reaction consisted of bicarbonate. This is reversible since all reactions demonstrate La Chatlier's Principle.

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

g) The HCl ions disassociate when put into a water solution to form an aqueous solution. The HCl addition turned the previous solution white and opaque after 10 drops. The chloride ion from HCl and the silver from the AgCO3 (from product of original reaction) formed a precipitate that appeared white and opaque. (It was not observed in the raw data sheet that a precipitate had formed but, it was observed that the solution was white and opaque.)

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

h) The addition of NH3 to the current solution caused all molecules in the solution to become soluble as the Ag ion (in the previous reaction) from the precipitate reacted with NH3 to form the soluble [Ag(NH3)2]+.
This caused the observations of a formation of a white substance on the bottom (precipitate) then disappeared as the test tube was gently tapped.

$$H^{+}\,(aq) + NH_3\,(aq) \rightleftharpoons NH_4^{+}\,(aq)$$

Repeating steps 7 & 9 : The same changes was not observed.

2. As HNO3 is added to the previous solution, the test tube became very warm and it appeared as if gas was forming. The reaction that occurred was exothermic and it was caused by the disassociation of HNO3. As HNO3 dissociated, it caused the release of hydrogen as well as the reaction with the ammonia from the product ([Ag(NH)3]2+) in the previous reaction. This created NH4. The decreasing concentration of NH3 causes the Ag+ to react with the Cl- that is present. As a result, the reaction Ag+Cl=AgCl takes place and a precipitate forms. Ammonium was then added to the previous solution. It was observed more gas had appeared to form, the test tube was cold to the touch and the reaction produced a lot of odour (may have been caused by use of ammonia). The reaction increased the amount of NH3 that was available. Therefore, the aqueous solution for [Ag(NH3)2]+ formed and the precipitate disappeared.

$$Ag^{+}\ (aq)\ +\ I^{-}\ (aq) \rightleftharpoons AgI\ (s)$$

i) The addition of KI to the solution caused a white precipitate to form with clear liquid underneath. The Kl dissociated to form K+ and I-. The Ag from [Ag(NH)2]+ reacted with I to produce the AgI precipitate.

$$2Ag^{+}\ (aq)\ +\ S^{2-}\ (aq) \rightleftharpoons Ag_2S\ (s)$$

j) The addition of Na2S to the previous solution created a brown precipitate that formed and clear liquid that formed underneath. This occurred from the reaction of the Ag ion from the previous product of the reaction (AgI (s)) with the dissociated S^2-. This formed the brown precipitate of Ag2S.

$$CH_3COOH\ (aq)\ +\ H_2O\ (l) \rightleftharpoons H_3O^{+}\ (aq)\ +\ CH_3COO^{-}\ (aq)$$

k&l) The universal indicator containing the acetic acid was red. The red colour indicates an acidic solution is present. The pH colour was red/orange which meant the pH was 2.5. Both pH's correspond to the solution being acidic.

m) The universal indicator containing acetic acid and NaCH3COO was red. This meant that the acidic solution was still present.

n) The pH paper turned orange which indicated that the pH was approximately a 3. This corresponded to the universal indicator because both exhibited acidic solutions present.

o) The universal applicator containing H2O was red which meant that an acidic solution was present. The pH paper turned yellow which meant that it had a pH of 6.

p&q) The addition of HCl to the distilled water created a solution that was more acidic. The universal indicator of the H2O and HCl was red which meant an acidic solution was present. The pH paper corresponded with a yellow/orange therefore, it had a pH of approximately 3-6. The pH level was significantly lower than the original pH of the water.

The addition of HCl to the buffer caused the pH to remain constant. The buffer and HCl solution appeared light red with the addition of universal applicator. This implied that an acidic solution was present. The pH paper corresponded with an orange colour which implied that the pH level stayed constant.

r) The buffer and NaOH appeared orange with the presence of the universal indicator. This implied that the solution was still acidic. The pH of the solution remained constant even though a base was added. The H2O and NaOH created a purple solution with the presence of the universal applicator. This indicated a basic solution was present. The pH of the solution increased.

s) The pH paper of the buffer and NaOH was a green/blue colour which indicated a pH of 11-12. This may have been brought on by an error because the universal applicator and pH does not correspond. The pH of the H2O and NaOH was blue colour which indicated that the pH was approximately a 13-14. This corresponded with the universal applicator.

3) is done in calculations area.

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

t) Initial observation of CoCl2 at room temperature was red/violet colour and odourless.

u) After the 3 drops of HCl to CoCl2, the solution turned blue and fumed. Initially, the CoCl2 reacted with water because it was in aqueous form. This caused CoCl2 to dissociate into Co^2- and Cl^1- forming [Co(H2O)6]^2+. The HCl is also a strong acid in aqueous form therefore, it dissociates into H+ and Cl-. When both reactants are added, a single displacement reaction occurs forming the above products. The blue colour is produced by the formation of the products.

v) The addition of H2O causes equilibrium to shift towards the reactants because of the influx in the product H2O. This occurs in order to achieve equal rate of reaction for reactants and products by opposing the imbalance. The addition of water created a light red solution which indicated the presence of the reactants.

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

w) The appearance of CuBr2 is a glittery obsidian.

x&y) The solid and H_2O became a solution (dissolved $CuBr_2$). This solution appeared to be a mint colour after H2o was added. The solution stayed a mint colour throughout.

z) The addition of water caused the solution to remain a mint colour. The equilibria involved is the reaction stated above because it exhibits the Cu ionizing and reacting with H_2O forming $[Cu(H_2O)_4]$.

aa) The initial appearance of KBr was observed to be white, snow-like crystals. When dissolved in H_2O, it became a clear solution. This implied that the dissociation of KBr formed K^+ and Br^- which caused the colour of the solution to be clear.

bb) As $CuBr_2$ was added to the KBr solution, the solution changed to a green colour. This exhibited the common ion effect because the green colour was previously seen in the $CuBr_2$ which implied that the solubility of the $CuBr_2$ decreased in the aqueous environment.

Hot water bath

(cc) As the previous solution was heated, the solution turned a dark, green colour with a murky appearance. There was also observed to be a precipitate that had formed. To explain this, the reaction may have shifted its equilibrium towards its products which was indicated in the precipitate forming.

(dd) The $CoCl_2$ solution did not appear to change as it was heated.

Calculations:

Ka of CH_3COOH = 1.8 X10^-5

Initial :pH of H_2O: Assumed to be 7 because it's a neutral solution

Initial :pH of buffer: $C_1V_1=C_2V_2$

C_2=(10 dr/g)(0.1M)/(20dr/g)

C_2=0.05M

pH= -log 1.8 x 10^-5 + log (0.05/0.1)

= 4.74

3a) **pH of H_2O +HCl:** H2o + HCl → $H_3O^+ + Cl^-$

$C_1V_I=C_2V_2$

C_2= (0.1M)(5dr/g)/(25dr/g)

$C_2 = 0.02M$

pH = -log(0.02) = 1.7

pH change from H_2O to H_2O with addition of acid

7-1.7= 5.3 drop in pH

pH of H_2O+ NaOH: $H_3O^+ + NaOH \rightarrow H_2O + Na^+$

$C_1V_1=C_2V_2$

$C_2=(5dr/g)(0.1M)/(25dr/g)$

$C_2=0.02M$

pOH= - log (0.02) = 1.7 pH= 14.00 – 1.7 =12.3

pH change from H_2O to H_2O and NaOH

7-12.3= 5.3 increase in pH

b) **pH of buffer + HCl**: $CH_3COO^-(aq) + HCL \rightarrow CH_3COOH + Cl^-$

	CH_3COO^-	HCL	CH_3COOH
I	0.04M	0.02M	0.04M
C	-0.02M	-0.02M	+0.02M
E	0.02M	0	0.06M

$C_1V_1=C_2V_2$

$C_2= (10dr/g)(0.1M)/(25dr/g)$

$[CH_3COO^-]=0.04M$

$C_1V_1=C_2V_2$

$C_2 = (0.1M)(5dr/g)/(25dr/g)$

[HCl] =0.02M

pH= 4.74 + log (0.02/0.06) = 4.26

pH change from buffer to buffer and HCl

4.74-4.26 = 0.48 drop in pH

pH of buffer + NaOH: $CH_3COO^-(aq) + NaOH \rightarrow CH_3COOH + Na^+$

	CH_3COO^-	NaOH	CH_3COOH
I	0.04 M	0.02M	0.04M
C	-0.02M	-0.02M	+0.02M
E	0.02M	0	+0.06M

$C_1V_1=C_2V_2$

$C_2=(0.1M)(10dr/g)/(25dr/g)$

$[CH_3COO^-]=0.04 M$

$C_1V_1=C_2V_2$

$C_2=(5dr/g)(0.1M)/(25dr/g)$

[NaOH]=0.02M

pH=4.74 + log(0.02/0.06) = 4.26

pH change from buffer to buffer and NaOH

4.74-4.26 = 0.48 drop in pH

Additional Discussion (if desired....otherwise, discussion can be combined with the observations in the table):

The noticeable source of error was seen in the pH paper when measuring the buffer solution with the addition of the base. The pH was supposed to remain somewhat constant which was demonstrated in the calculations. This may have been brought on by other solutions that travelled/was absorbed down the length of the pH paper. This could have resulted in contaminations and altered the pH level of the solution being measured. For instance, the solution of water and base could have contaminated the results for the buffer and base creating an illusion that the buffer and base were basic.

Conclusion

Equilibrium Shift
La Chatelier's Principle was effectively demonstrated in the reactions. This was achieved by observing the reaction minimizing the change in equilibrium by either favouring the reactants or products. The addition of HCl favoured the reactants and the addition of NH_3 favoured the products.

Multiple Equilibria
The use of numerous reactants provided the demonstration of equilibriums acting for sub parts of a reaction that were all linked together (usually with the use of the products from the previous reaction reacting with the addition of a chemical). Equilibria can be achieved through numerous reactions.

Buffer system
The changes of pH in buffer system and the base and acid were extremely small (-0.48). This demonstrates the characteristics of buffer systems as they are in constant equilibrium and does not change with the addition of other substances. The changes in pH of the water and the pH with the addition of base (+5.3) and acid (-5.3) was extremely large because the strong acid and base dissociated completely therefore favouring the products.

Common Ion Effect

The common ion was Br-. The solution returned to a similar colour of the original $CuBr_2$ which implied that the solubility of it decreased causing the solution to change into a green colour. This resulted in an equilibrium shift to the products $[Cu(Br)_4](aq)$.

Temperature Effect

The solution exhibited equilibrium shift as underwent temperature change. It favoured its products which was indicated in the precipitate form that was observed.

(5) If steps 2 & 3 repeated, sol'n turns ~~dark blue~~ then sol'n turns light blue

eq. shift +

(1) Before

$CuSO_4$ - light, blue sol'n, transparent

NH_3 - ~~base~~ strong smell, clear

(3) HCl: strong, acidic; smell, clear, liquid

(2) After ≃ 2 drops

- dark blue, ~~transparent~~, immediate rxn

27/08/16

(4) w HCl

w every drop, color becomes lighter, shades of blue until total light blue

(3) (6) $CuSO_4 + NH_3$ sol'n

dark blue, transparent, immediate rxn

(7) w HCl

→ white, opaque

~~w HCl ≃ 10 drops~~

w HNO_3

- percipatate formed

- clear, foggy, liquid

Angela T

Sept. 27 2016

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Multiple Equilibria

(1) Before

Na_2CO_3

→ no odour

→ clear

$AgNO_3$

- clear liquid
- no odour

(4) w̄ HCl ≈ 10 drops

→ white opaque

(7) when steps 7-9 repeated

→ w̄ HNO_3

- test tube heated up
- gas formd (formed) @ top

→ w̄ NH_3

- alot of gas formed
- test tube cooled
- alot of odour

(9) w̄ Na_2S

- percipitate turns brown
- perciptate floats @ top

→ rest of liquid clear

(2) After

→ orange tinted @ first then beige & opaque

(3) w̄ HNO_3

→ percipatate formed

→ clear, foggy liquid

(6) w̄ NH_3

→ white substance formed on bottom then disapeared when gently tapped

(8) w̄ KI

→ perciptate formed @ top

- white in colour
- rest of liquid clear

27/09/16

Sept. 27th 2016

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	Ph colour ~~=~~ red	
CH_3COOH	red/orange	red
CH_3COOH + ~~H_2O~~ $NaCH_3COO$	orgage	red
H_2O	yellow	red
Buffer + HCl	orange	red light
Buffer + NaOH	green/blue	orange
H_2O + HCl	yellow/orange	red
H_2O + NaOH	blue	purple

Buffer System

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Sept 27th 2016
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27/09/16

~~Equation~~ common ion effect

$CuBr_2$ + KBr sol'n

Before

CuBr - light green liquid KBr - clear liquid

After ~~KBr~~ - darker green/aqua

$CoCl_2$ + KBr in hot bath
- dark, murky green
  - almost opaque
  - black hue
  - liquid splashed on testtube ∴ black/green dots

$CoCl_2$ Before
- light red

$CoCl_2$ in hot water bath
- darker red

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

→ HCl strong scent, clear

→ $CoCl_2$ red / violet color, odorless

→ after ≃ 3 drops HCl the solution turned blue + fumed

→ after ≃ 10 drops H_2O, solution turned to light red

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

→ $CuBr_2$: glittery, obsidian

↳ w̄ H_2O: turned mint solution

→ KBr : white, snow-like crystals, clear when added to $CuBr_2$ (solution green)

→ Hot water bath

↳ w̄ solution: ~~clear, may have to much water~~ dark, murky green, precipitate seen

↳ w̄ ~~$CoCl_2$~~: No change

22/09/16