Student Name: Jenna Palombaro

Student Number: 8686718

Partner's Name and Student #: Sam Touma 8620325

Demonstrator's Name: Michael Forberg

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): Thursday

Lab Week (even/odd): Odd (7)

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

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 JP

Observations and Discussion

Part 1: Equilibrium Shift

The initial copper sulfate is observed to be light blue with no odour. This can be expected as a result of the Cu^{2+} ions because they are blue whereas the SO^{2+} ions have no colour.

When concentrated ammonia is added to copper sulfate, an intense violet blue colour projects throughout the solution upon mild mixing. The colour change is attributable to LeChatelier's Principle. The concentration of the ammonia is increased and therefore there is an equilibrium shift to the right in order to minimize the effect of changes imposed on the system. The equilibrium constant becomes smaller as the excess NH_3 is used up to produce larger amounts of $[Cu(NH_3)_4)^{2+}]$. This ion has an even darker colour of blue than the original Cu^{2+} ion making it responsible for the change in colour to violet blue.

Upon the addition of the first four drops of hydrochloric acid, the solution lightens in colour slightly but obtains a cloudy appearance. More hydrochloric acid is added and the solution returns to its original transparent light blue state. This occurs because water molecules replace the NH_3 ion as the Cl^- ion displaces it. Therefore, the more HCl added, the less NH_3 present in the solution. This decrease in reactant causes a left shift in the equilibrium as $(Cu(NH_3)_4)$ breaks apart and releases $NH_{3,}$ therefore decreasing the concentration of the ion. This decrease results in the recurring characteristics of the solution.

It can be expected that by adding more NH_3 and HCl accordingly, the same changes will occur. This is because it is just a repetitive action of LeChatelier's Principle. By adding more $NH_{3,}$ the equilibrium is shifted back towards the right and concentration of ammonia increases. By adding more HCl, the concentration of ammonia is decreased therefore it is shifted to the left. This reaction will continue to reoccur but higher levels of solution may be needed each time.

Part 2: Multiple Equilibria

The 0.1mol/L Na_2CO_3 and $AgNO_3$ are clear colourless liquids and neither have an odour. When 0.5mL of $AgNO_3$ is added to the Na_2CO_3 the result is the formation of a cloudy, light brown solution with a precipitate forming. No temperature change is observed.

Next, 6mol/L HNO_3, a clear liquid, is added to the solution. After 5 drops is added to the test tube, the solution begins to revert to its clear state. An evident increase in temperature is observed. This is a result of the dissociation of HNO_3 in water. The H^+ ions bond with whichever CO_3^{2+} molecules are of availability and form H_2CO_3. This formation is accountable for the return of the clear and colourless state as shown in the equation: $2\ H^+\ (aq)\ +\ CO_3^{2-}\ (aq) \rightarrow H_2CO_3\ (aq) \rightarrow H_2O\ (l)\ +\ CO_2\ (g)$.

The addition of 4 drops of 1.0mol/L HCl results in a cloudy, translucent white solution after gentle mixing where there is no temperature or odour change detected. When adding the HCl, it dissociates into its H^+ and Cl^- ions. These ions make new bonds as shown in the equation of $Ag^+ (aq) + Cl^- (aq) \rightarrow AgCl (s)$. The Cl^- ion bonds with Ag^+ to form the precipitate of $AgCl_{(s)}$.

10 drops of concentrated NH_3 are then added to the solution. It results in a clear transparent liquid, appearing much like the original state of the solution. The added NH_3 is now available to bond with Ag^+ ions and it replaces the Cl^- ion in all precipitates previously formed. These newly introduced bonds are responsible for any odours detected.

When steps 7 and 9 are repeated, the same changes occur. The solution becomes transparent and colourless once again.

Continuing on, 7 drops of clear liquid 0.1mol/L KI are added to the solution. A foggy liquid with a yellow tint is produced. No noticeable changes in temperature or odour are observed. KI is broken down into its ions of K^+ and I^- where the I^- bonds with the Ag^+, replacing the NH_3. AgI is increased in the solution. This dissociation and new bond cause the characteristics of the reaction.

The last step of Multiple Equilibria includes mixing in clear colourless 0.1mol/L Na_2S to the solution. 7 drops are added before seeing a change. A dark grey colour begins to form on the top of the test tube but after gentle mixing, it spreads to the entire sample. The liquid appears to be opaque and there is a possibility of a precipitate being formed. No temperature or odour change is detected. Na_2S dissociates in the solution into its ions. The concentration of S^{2-} ions increases and they bond with all available Ag^+ ions. This results in Ag_2S.

Part 3: Buffers

To begin, using a plastic transfer pipet, 10 drops of clear 0.1mol/L CH_3COOH with a vinegar like odour are distributed into two wells in a well plate. To this, 3 drops of universal indicator - transparent red tone with a weak odour - are added. The universal indicator has a pH of roughly 3. When tested, the solution has a pH of 4 as the paper turns light orange for both of the wells.

Next, 10 drops of clear, odourless 0.1mol/L $NaCH_3COO$ are added to each well. Since the sodium acetate has a strong base parent and a weak acid parent, it is slightly basic in water. The acetic acid provides sodium acetate with H^+ ions to bond with so the basic properties of the solution can be absorbed as seen in the equation: $CH_3COOH (aq) + H_2O (l) \rightarrow H_3O^+ (aq) + CH_3COO^- (aq)$.
Due to this, the pH of the wells has not changed and the solution still becomes light orange and transparent. Therefore, the solution after the addition of the salt is a buffer system.

The pH for both wells still remains at 4 with the colour of light orange on the pH paper. They have not changed because of the absorption of basicity done by the buffer.

Next, in two new wells, 20 drops of distilled water are transferred into each and then 3 drops of the universal indicator are added. This results in a transparent red liquid with a pH of 6 as when the pH paper is used, the colour shown is an olive green.

5 drops of clear, liquid 0.1mol/L HCl is added to one well containing the buffer and to one well containing the water. No visible change in either of the wells is observed. The pH of the buffer remains at 4 (light orange pH paper) because the buffer system resists change by absorbing the acidity of the HCl added. The weak base takes on the H^+ ions and absorbs the acid. On the other hand, the pH of the water solution changed to 2 (dark orange) as there was no buffer present to resist change.

Lastly, an addition of 5 drops of clear liquid 0.1mol/L NaOH was made to the other wells of the buffer and water solutions. The buffer solution lightened a small amount but the pH remained at 4 (light orange pH paper). This happens because the weak acid takes on the basicity as a result of the buffer resisting change in pH. On the contrary, the water turns transparent and purple with a pH change to 13 (dark blue pH paper). The base had such a strong effect on the solution because there was nothing used to absorb the basicity.

Part 4: Common-Ion Effect

This section of the experiment begins with 5 drops of magenta, odourless 1mol/L $CoCl_2$ being combined with concentrated, clear, strong 12 M HCl with a strong odour. After 4 drops of HCl is mixed in, the solution turns to an opaque blue colour. Upon gentle mixing, the colour darkens to a darker pink colour. The excess chloride added in the HCl shifts the equilibrium to favour the blue chloride complex. By adjusting amounts of Cl^- present in the solution, the equilibrium can be rearranged with both complexes present in the concentration. This would produce a purple solution.

The addition of five drops of distilled water causes the solution to lighten significantly. This is because the hydrate ($Co(H_2O)_6^{+2}$) is formed and equilibrium is being re-established. The solution is beginning to work towards the complete equilibria of the purple solution. Therefore, the water causes a left shift.

Next, a small amount of metallic, blue-grey solid $CuBr_2$ is added to an empty well and then is mixed with five drops of distilled water. This results in a black opaque liquid. When ten more drops of water are added, the solution becomes a light blue liquid with all the $CuBr_2$ dissolved. Lastly, twenty-five more drops are added which lightens the colour slightly, makes the solution transparent and obviously increases the volume.

These changes in colour occur as the hydrate $Cu(H_2o)_4^{-2}$ is formed as seen in the equation: $CuBr_2 + 2H_2O \rightarrow Cu(H_2O)_2Br_2$. When H_2O is added, the equation: $4Br^-(aq) + [Cu(H_2O)_4]^{2+}(aq) \approx [CuBr_4]^{2-}(aq) + 4\ H_2O(l)$ - shifts to the left as it becomes more and more aqua blue.

A small amount of solid white KBr crystals is then added into an empty well. Twenty drops of water are added to entirely dissolve the crystals and results in a clear liquid. The crystals are dissociated into their K^+ and Br^- ions. No reaction occurs because the solution is neutral and it can be assumed the pH of the solution will be 7. It can also be assumed that both ions in the salt are colourless because even if there was an equilibrium shift, it cannot be detected by the naked eye due to the lack of colour change.

Lastly, a sample of clear solid KBr is added to 0.5mL of the solution set aside in step 25. The KBr turns brown and sinks to the bottom of the test tube. After gentle mixing, the liquid turns green. There is no temperature or odour change.

Part 5: Temperature Effect

The solution previously prepared is submerged into a hot water bath. The solution immediately changes colour to dark brown with no odour change and an obvious increase in temperature.

$CoCl_2$, a magenta liquid, is then placed in the boiling water bath. Nothing significant takes place but small blue droplets form on the side of the test tube. It can be assumed that if the solution is submerged for a longer period of time, the entire sample will turn blue. According to LeChatelier's Principle, the equilibrium is shifting to the right as heat is added to the reaction. The equilibrium is endothermic because it becomes more blue as it is heated.

Within this lab, there are many possible sources of error. Some measurements given in the experiment were given in "pinches" or "drops" using a pipet. It is possible that an incorrect amount was distributed due to excess drops or different drop sizes. Whether too little or too much was added, it is possible that this could skew any results.

Calculations:

Change in pH of buffer with the addition of an acid:

1. Find concentration

$C_F = C_I \times (V_I/V_F)$
$[CH_3COOH]_F = 0.1$ mol/L x ((10dr/g)/(20dr/g))
= 0.05 mol/L

2. Find initial pH
$pH_{water} = 7.0$

pH = pKa + log[A^-] / [HA]
pH = pKa + log(0.05/0.05)
pH = pKa
pH = -logKa
= $-\log 1.76 \times 10^{-5}$
= 4.75

3. Find initial concentration

[HCl]= (0.10 mol/L)(5 drops/25 drops)
= 0.02 mol/L

[CH_3COO^-] and [CH_3COOH] due to 1:1 ratio
= (0.10 mol/L)(5 drops/25 drops)
= 0.04mol/L

4. Make an ICE Chart

	CH_3COO^-	HCl	CH_3COOH
Initial	0.04 mol/L	0.02 mol/L	0.04 mol/L
Change	-0.02 mol/L	-0.02 mol/L	+0.02 mol/L
Equilibrium	0.02 mol/L	0 mol/L	0.06 mol/L

5. Calculate pH

pH= pKa+ log [A^-]/[HA]
= 4.74 + log [0.02 mol/L]/[0.06 mol/L]
= 4.74- 0.48
= 4.26

Change in pH of buffer with the addition of a base:

1. Find Concentrations (initial)

[NaOH] = 0.1 mol/L (5 drops/25 drops)
= 0.02 mol/L

[CH_3COOH] and CH_3COONa (1:1 ratio) = 0.05 mol/L (20 drops/25 drops)
= 0.04 mol/L

2. Make an ICE chart

	NaOH	CH_3COOH	CH_3COONa
Initial	0.02 mol/L	0.04 mol/L	0.04 mol/L
Change	-0.02 mol/L	-0.02 mol/L	+0.02 mol/L
Equilibrium	0 mol/L	0.02 mol/L	0.06 mol/L

3. Calculate pH

pH=Pka+log(0.06mol/L)/(0.02mol/L)
= 4.74 +0.48

Change in pH of water with the addition of an acid:

H2O = 20 drops = 1 mL
[HCl] = 0.1 mol/L
HCl = 5 drops = 0.25 mL
$[HCl]_F = [HCl]_I \times V_i/V_F$
$[HCl]_F$ = 0.1 mol/L × (0.5 dr/g / 25 dr/g)
$[HCl]_F$ = 0.2 mol/L pH = -log[H+]
= -log(0.02)
= 1.7

Change in pH of water with the addition of a base:

H2O = 20 drops = 1 mL
NaOH = 5 drops = 0.25 mL
[NaOH] = 0.1 mol/L
CF = C1× (V1/VF)
= (0.1)(0.25) / 1.25
= 0.02 mol/L
[OH-] = 0.02 mol/L
pOH = -log[OH-]
pOH = -log(0.02)
pOH = 1.7
pH = 14 – pOH = (14 – 1.7) = 12.3

Conclusion:

The experiment exemplified the way that LeChatelier's Principle can be used to describe reactions. Equilibrium will always work in a way that minimizes any effects or changes on a system at equilibrium. It will do this by giving or taking reactants or products. Reactions already at equilibrium can be reversed simply with the addition of the right reagents and buffer systems will work to absorb any acidity or basicity added to a solution to maintain equilibria.

④ go back to clear.

Table 1. Observations and Discussion

$[Cu(H_2O)_4]^{2+}$ (aq) + 4 NH_3 (aq) ≈ $[Cu(NH_3)_4]^{2+}$ (aq) + 4 H_2O (l) ① $CuSO^4$ -clear. -no odour. - light blue ② NH_3 conc -odour strong mix -turned intense blue. -bubbling/fogging -odour weakened -no temp change. ③ mix -cloudy -first 4 drops light blue -more added -returned to clear original state/light blue. HCl -clear -odour. -no temp change
2$AgNO_3$ (aq) + Na_2CO_3 (aq) ≈ Ag_2CO_3 (s) + 2 $NaNO_3$ (aq) ① Na_2CO_3 -clear -no odour ② $AgNO_3$ -colourless -odourless -transparent. mixed -cloudy -no temp -brown. -possible formation of precipitate ③ HNO_3 -odour - clear mixed -reverting to clear state - ~~little bit effect~~ ~~white~~. ④ mix w/ HCl
2 H^+ (aq) + CO_3^{2-} (aq) ≈ H_2CO_3 (aq) ≈ H_2O (l) + CO_2 (g) 5 drops reverted to clear state. -heat produced
Ag^+ (aq) + Cl^- (aq) ≈ AgCl (s) ⑧ add 1mol/L HCl. → 4 drops -became white -cloudy -cannot see throug -no temp or odour change

Equ shift

Mult Equilibria.

$Ag^+ (aq) + 2NH_3 (aq) \approx [Ag(NH_3)_2]^+ (aq)$ add concentrated NH_3 → 10drops -reverted to clear state again -heat produced
$H^+ (aq) + NH_3 (aq) \approx NH_4^+ (aq)$ (10) goes back to same state as before (7) smoke. (9) clear HNO_3 ~~-heat produce at~~ -heat producted
$Ag^+ (aq) + I^- (aq) \approx AgI (s)$ (11) KI → 7 drops clear on its own mix -foggy -yellow tint -no temp change -no odour
$2Ag^+ (aq) + S^{2-} (aq) \approx Ag_2S (s)$ Na_2S clear mix → 7 drops -dark grey colour change → formed on top first then spread -no temp change -foggy - possible precipitate.

(14) no visible change in either of the wells
pH buffer → 4 (lighter orange)
pH water → 2 (dark orange)

(20) ... (clear)
buffer → lightened a small amt pH → light 4 orange.
water → turned purple / clear. pH → dark blue 13

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

-clear
-vinegar odour.
+
Universal ~~indicat~~ indicator.
= transparent red colour. (light).

→ weak odour.
-scent of CH_3COOH.

pH paper.

Well ①	Well ②
- light orange -acidic	- light orange. -acidic

pH of 4 for both.

(17) $NaCH_3COO$ → clear / no odour
pH → pH has not changed
-same odour & colour (tad lighter)
(18) water + U.I. → clear / red
pH → 6 (olive green)

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

(21) $CoCl_2$
magenta
-no scent

(22) mix w/ HCl → 4 drops.
- first few drops changed to blue (not transparent)
-mixed to dark pink colour
↳ pink is transparent
-no temp change
(23) distilled water
-lightened colour.

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

Solid $CuBr_2$.
dark blue / grey
shiny / sparkly
powderish.

X / dark liquid / black.
Y / light blue liquid
-all metal dissolved
Z / 25 drops added
-increase in volume
-slightly lighter colour.

Solid KBr.
white crystals
+
distilled water
clear liquid
20 drops water

(27) sltn + KBr. KBr.
- brown ~~precipitate~~ at bottom.
- ~~water~~ liquid turned green after mixing
-no temp change / no odour change

Test tube in boiling water

- no odour change
- obvious + temp
- liquid to dark brown

Place $CoCl_2$ (1mol/L) in clean test tube. Put in boiling water bath

- blue dots on side of test tube
- no major change in main liquid