*Student Name:* ____ ___________________________

*Student Number:* __________________________________

*Partner's Name and Student #:*

*Demonstrator's Name:* ______________ ____________

**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 _Zahra A.__***

**Table 1. Observations and Discussion**

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

Observation:

- 0.1 M $CuSO_4$: Light blue liquid, no odour
- $NH_3$: colourless, clear liquid, strong odour
- 0.1 M HCl: colourless, clear liquid, strong odour

Procedure:

- When 4 drops of $NH_3$ were added to 20 drops of $CuSO_4$, the solution turned into a deep translucent blue
- 28 drops of HCl were further added to the same solution resulting in transparent solution
- 3 drops of $NH_3$ were added to the solution, changing it to the previous deep translucent blue
- Then 56 drops of HCl were added to the same solution, bringing back the original solution colour

Discussion:

When $CuSO_4$ dissociated in the solution, a single displacement reaction took place between Cu and $NH_{3.}$ This contributed to why the colour changed to translucent dark blue. According to Le Châtelier's Principle, when a system undergoes a particular stress factor, the equilibrium would shift to accommodate that change. The reaction was consuming excess $[Cu (H_2O)_4]^{2+}$ and $NH_3$ ions, resulting in an equilibrium shifting to the right. Product was produced by consuming the reactants to maintain equilibrium.

When HCl was added to the same solution containing the $Cu(NH_3)$ ions, the solution turned back to its original colour; light blue. Because $H^+$ ions were now present in the solution, they resulted in the formation of $[Cu(NH_3)_4]^{2+}$ ions. These ions gave off the light blue colour of the solution. Eventually, the equilibrium would shift to the left to counteract the change in the reaction. The steps were repeated, revealing that the trend was continued and the observations were similar.

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

Observations:

- 0.1 M $Na_2CO_3$: colourless, clear liquid, strong odour
- 0.01 M $AgNO_3$: transparent liquid

Procedure:

- 10 drops of both $Na_2CO_3$ and $AgNO_3$ were mixed in a clean test tube, a brownish liquid was seen. Precipitate was seen at bottom of test tube

Discussion:

A double displacement reaction took place between the reactants, producing two products, one of which is solid. The reaction between the reactants produced a brown solution, as well as a precipitate. This explained the presence of a solid product and the shifting phase of the reaction. Therefore, due to Le Châtelier's Principle; the equilibrium shifted to the right to accommodate the stress that was applied. As a result, a precipitate was produced.

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

Observations:

- 6 M $HNO_3$: colourless. Clear liquid, strong odour

Procedure:

- 4 drops of $HNO_3$ were added to the brown coloured solution. As a result, the solution turned clear immediately

Discussion:

$HNO_3$ was added to the previous solution, which permitted the reversal of the solution to its original transparency. Based on the Le Châtelier's Principle, the increase of $H^+$ ions created a reaction with the solid. Eventually, the equilibrium shifted to the left to accommodate the concentration of $H^+$ ions which caused the reversal of the solid; allowing the solution to revert back to a colourless liquid.

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

Observations:

- 0.1 M HCl: transparent, clear liquid

Procedure:

- 6 drops of HCl were added to the solution. A thick white cloud appeared on the surface of solution
- Precipitate was visual

Discussion:

A thick cloud and precipitate formed as HCl was added. Le Châtelier's Principle stated that in this case, the system was shifting to the right to accommodate the increase of $H^+$ ions in the solution. In addition to that, $Cl^-$ ions were introduced which reacted with $Ag^+$ ions to form a solid AgCl.
The increase of reactant resulted in the formation a precipitate, thus an equilibrium shifting to the right

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

Observations:

- $NH_3$: colourless, clear liquid, very strong dour

Procedure:

- 4 drops of $NH_3$ were added to the solution, slowly turning it back to a clear liquid

Discussion:

The reason the solution became clear again was because $NH_3$ dissolved AgCl, the precipitate that was formed from previous solutions. Dissolving it resulted in the formation of $[Ag(NH_3)_2]^{+}$ ions within an aqueous solution. Therefore, the equilibrium must shift to the left to maintain order.

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

Observations:

- $NH_3$: colourless, clear liquid, very strong odour
- 6 M $HNO_3$: clear, colourless liquid, very strong odour

Procedure:

- When 30 drops of corrosive $HNO_3$ were added to the solution, a thin layer of white cloudiness appeared in the test tube
- Heat was also produced since the test tube was getting warmer
- 7 drops of $NH_3$ were then added, which made the solution change back to a clear, colourless liquid

Discussion:

As $HNO_3$ was added, H+ ions increase. Based on Le Châtelier's Principle, the equilibrium clearly shifted to the right. When $NH_3$ was added, the hydrogen ions reacted with it to form ammonium ions (more product). Eventually, the H+ ions were consumed and the increase in the product caused the equilibrium to shift to the left. This resulted in the clear solution that was initially observed.

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

Observations:

- 0.1 M KI: colourless, clear liquid, no strong odour

Procedure:

- 4 drops of KI were added to the solution. A thick white cloud formed creating a division between the cloudiness and clear area at the bottom of the test tube
- precipitate formed

Discussion:

KI introduced a new ion to the solution; $I^-$ ions. These ions reacted with $Ag^+$ ions which resulted in the reformation of a precipitate (AgI). The reactant concentration increased, causing the equilibrium to shift to the right. As it shift to the right, more product was produced; thus, the explanation behind the white cloudy texture.

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

Observations:

- 0.1 M $Na_2S$: colourless, clear liquid, no strong odour

Procedure:

- 4 drops of $Na_2S$ were added to the solution. The solution turned brownish-grey colour of condensed cloudiness
- Precipitate at bottom of test tube

Discussion:

Based on the observations and Le Châtelier's Principle, the equilibrium shifted to the right to produce more of the product. The addition of $Na_2S$ resulted in the colour change of the solution and the creation of a new precipitate. As $S^{2-}$ ions from $Na_2S$ reacted with $Ag^+$, the result was more product of $Ag_2S$. Therefore, equilibrium shifted to the right to produce more of the product to maintain balance.

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

Observations:

- 0.1 M $CH_3COOH$: clear liquid, no odour
- Universal indicator: orange-red solution, mild odour
- 0.1 M $NaCH_3COOH$: clear liquid, no odour
- 0.1 M NaOH: colourless liquid
- 0.1 M HCl: clear liquid, no odour

Procedure:

**Buffer:**

- 3 drops of universal indicator were added to both wells containing the 10 drops of $CH_3COOH$**.** When mixed, the solution remained liquid and turned slightly red with a mild smell
- Based on the PH chart, the pH of solution was approximately 2.5
- Afterwards, 10 drops of $NaCH_3COOH$ was added to both wells containing the acid, turning the solution into a light orange colour. pH was measured to be 2.5
- First well added with 5 drops of HCl. The solution turned into a light orange colour. A pH of 5
- Second well added with 5 drops of NaOH. The solution turned into a very light orange colour. A pH of 4

**Distilled water:**

- 3 drops on universal indicator into both wells containing 20 drops of distilled water
- 5 drops of HCl into first well - light red -pH of 3
- 5 drops of NaOH into second well – dark blue/ purple – pH of 12

Discussion:

When $NaCH_3OOH$ was added to $CH_3COOH$, the colour's solution turned light orange. The pH of the solution seemed to be similar to the pH of the acid. Therefore, the weak acid and the conjugate base are equal in values and are within a balanced system. When a strong acid (HCl) was added to the solution, hydronium ions increased. The Le Châtelier's Principle tells us that the system would shift in the other direction (equilibrium to the left). Since the values are equal, the pH value would not change so much. Based on the observations, the pH shifted by half. In comparison, adding a base to the solution meant an increase in hydroxide ions. This would cause the system to shift to the right. Meanwhile, the pH does not drastically change. This is due to the fact the buffer resisted the pH change.

When the universal indicator was added to the distilled water, the pH was measured to be 6. However, the pH of water is approximately 7. This meant that the distilled water was acidic. This can be explained due to the fact that distilled water can react with carbon dioxide in the air to produce carbonic acid (weak acid). When HCl was added to the distilled water, the result is a light red substance with a pH of 3. The addition of an acid means an increase in hydrogen ions causing the solution to be more acidic than it already is. There was a significant shift of pH in the solution. The distilled water turned to a blue-purple colour when the base was added. This resulted in a pH of 12. The NaOH is a strong base, it dissociated to form $Na^+$ and $OH^-$. Meanwhile, water acted as a conjugate acid that would give away a hydrogen ion to produce $OH^-$ ions; resulting in the formation of a basic solution. The increase of the hydroxide ions in the solution created an imbalance of concentration and a change in the pH levels.

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

Observations:

- 1 M $CoCl_2$: pink colour, odourless
- 12 M HCl: clear liquid, very strong odour

Procedure:

- 4 drops of HCl were mixed with 5 drops of $CoCl_2$ – resulting in a purple solution
- 5 drops of distilled water were added to the same solution – resulting in a light pick solution

Discussion:

The addition of HCl to the solution changed it to a purple solution. This is because of the $Cl^-$ ions binding to the $Co^{2+}$ ion. This resulted in the common-ion effect. As a result, HCl was ionized, forming $[CoCl_4]^{2-}$ ions. Equilibrium shifted to the right since the increase of chlorine ions created an imbalance. As the chlorine content increased, there were more product of $[CoCl_4]^{2-}$ ions.

The system then shifted left when distilled water was added. Keeping in mind that there was already a proper amount of water at equilibrium. Increasing the amount of water shifted the equilibrium, thus hydrating the $[CoCl_4]^{2-}$ ions. As a result, $[Co(H_2O)_6]^{2+}$ ions were produced allowing the solution to turn back into a light pink solution.

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

Observations:

- CuBr: solid, shiny black colour
- KBr: beige crystals – solid

Procedure:

- 5 drops of distilled water mixed with CuBr – extremely dark brown solution
- Additional 10 drops to solution – lime green
- Additional 25 drops (to reach 2 ml) – extremely light blue
- When 25 drops of distilled water were added o KBr, it dissolved completely

Discussion:

As distilled water was being added to CuBr, the solution's colour became lighter and lighter. First, it turned very dark brown, then green, and eventually extremely light blue. This meant that the system was shifting to the left; causing an increase in the concentration of $[Cu\ (H_2O)_4]^{2+}$ ions. The shifting was taking place due to the imbalance of the concentrations.

KBr is a soluble compound that dissociated when mixed with water. The negative charge on Br- attracted the hydrogen ions ($H^+$) on water, while the positive potassium ions attracted the negatively charged oxygen in the water.

Temperature Effect (continued)

Observations:

- 1 M $CoCl_2$: pink colour, odourless

Procedure:

- the solution of CuBr and distilled water were mixed with a pinch of KBr in a clean test tube
- results in an olive green colour change. This mixture was then placed in a hot bath, which then turns very dark brown. When 20 drops of water were added to the test tube while in the hot bath, the solution turned light blue (continued from previous reaction)
- When plain $CoCl_2$ was put in a hot bath, solution changed to wine colour; however, blue colours were shown on edge of test tube

Discussion:

Adding heat or rising the temperature should move the equilibrium in the opposite direction; therefore, the system shifted left. This resulted in a darker colour as shown in the observation. However, when water was added, the concentration of $H^+$ ions in the solution increased causing the system to shift to the right. As the system shifted to the right to accommodate the increasing amount of H+ ions, the colour of solution turned light blue.

The product was favoured when $CoCl_2$ was placed in the hot bath. As a result, an endothermic reaction took place. Blue colours appeared on the edge of the test tube while it was in the hot bath. Temperature is another kind of stress that could affect the system of an equilibrium reaction. Le Châtelier's principle reacted to the reaction by reducing the heat; it did that by shifting the equilibrium to the opposite side of the heat. Therefore, the equilibrium shifted to the right where $CoCl_4$ exists.

**Calculations:**

**pH of $H_2O$ = 7.0**

**pH of Buffer:**

$CH_3COOH$ (aq) + $H_20$ (l) $\leftrightarrow$ $CH_3COO^-$ (aq) + $H_3O^+$ (aq)
10dr/g 10 dr/g
0.1 M 0.1 M

$[CH_3COOH] = \frac{C1V1}{V2}$

$= \frac{0.1\,M\,(10\frac{dr}{g})}{20\frac{dr}{g}}$

= 0.05 mol/L

$[CH_3COO^-] = [CH_3COOH]$

= 0.05 mol/L

$pH = PK_a + \text{Log}\frac{[A-]}{[HA]}$

$K_a(CH_3COOH) = 1.8 \times 10^{-5}$

$= -\text{Log}\,K_a + \text{Log}\frac{[A-]}{[HA]}$

$= -\text{Log}\,(1.8 \times 10^{-5}) + \text{Log}(\frac{0.05\,mol/L}{0.05\,mol/L})$

**= 4.74**

**pH of acid and water:**

[HCl] = 0.1 M

$V_i$ = 5 dr/g
$V_{water}$ = 20 dr/g
$V_f$ = 25 dr/g

$C_iV_i = C_fV_f$

$C_f = \frac{(o.1\,M)\,(5\frac{dr}{g})}{25\,dr/g}$

$= 0.02\,mol/L = [HCl] = [H^+]$

pH = - Log (0.02)

**= 1.70**

**pH of water and base:**

[NaOH] = 0.1 M

$V_i$ = 5 dr/g
$V_{water}$ = 20 dr/g
$V_f$ = 25 dr/g

$C_iV_i = C_fV_f$

$C_f = \frac{(0.1\ M)\ (5\frac{dr}{g})}{25\ dr/g}$

= 0.02 mol/L = [NaOH] = $[OH^-]$

pOH = - Log ($OH^-$)
= - Log (0.02)
= 1.70

pH = 14 – pOH
= 14 – 1.70
= **12.3**

**pH of buffer and acid:**

| $CH_3COO^-$ (aq) | + HCl (aq) | $\leftrightarrow$ | $CH_3COOH$ (aq) + | $Cl^-$ (aq) |
|---|---|---|---|---|
| 10 dr/g | 5 dr/g | | 10 dr/g | |
| 0.01 M | 0.10 M | | 0.10 M | |

$V_{total}$ = 25 dr/g

$[CH_3COO^-] = \frac{(0.1\ M)(10\frac{dr}{g})}{25\frac{dr}{g}}$
= 0.04 mol/L

$[HCl] = \frac{(0.10\ M)(5\frac{dr}{g})}{25\frac{dr}{g}}$
= 0.02 mol/L

$[CH_3COOH] = [CH_3COO^-]$
= 0.04 mol/L

-------------------------------------------------------------------------------------------------------------------

$CH_3COO^-$ (aq) + HCl (aq) $\leftrightarrow$ $CH_3COOH$ (aq) + $Cl^-$ (aq)

| | $CH_3COO^-$ | HCl | $CH_3COOH$ |
|---|---|---|---|
| I | 0.04 M | 0.02 M | 0.04 M |
| C | - 0.02 M | -0.02 M | + 0.02 M |
| F | 0.02 M | 0 | 0.06 M |

$pH = PK_a + Log \frac{[A-]}{[HA]}$

$= 4.74 + Log \frac{(0.02\ M)}{(0.06\ M)}$

= 4.74 – 0.48

= 4.26

**Change of pH for buffer and acid:**

$\Delta$ pH = $pH_{final} - pH_{initial}$
= 4.26 – 4.74
= **-0.48**

**pH of Buffer and base:**

| | | | | |
|---|---|---|---|---|
| $CH_3COOH$ (aq) + | NaOH (aq) | $\leftrightarrow$ | $CH_3COO^-$ (aq) + | $Na^+$ (aq) |
| 10 dr/g | 5 dr/g | | 10 dr/g | |
| 0.01 M | 0.1 M | | 0.01 M | |

$V_{total}$ = 25 dr/g

$[CH_3COOH] = \frac{(0.1\ M)(10\frac{dr}{g})}{25\frac{dr}{g}}$ = 0.04 mol/L

$[NaOH] = \frac{(o.10\ M)(5\frac{dr}{g})}{25\frac{dr}{g}}$ = 0.02 mol/L

$[CH_3COO^-]$ = $[CH_3COOH]$ = 0.04 mol/L

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$$CH_3COOH\ (aq) + NaOH\ (aq) \leftrightarrow CH_3COO^-\ (aq) + Na^+(aq)$$

| | $CH_3COOH$ | NaOH | $CH_3COO^-$ |
|---|---|---|---|
| I | 0.04 M | 0.02 M | 0.04 M |
| C | - 0.02 M | -0.02 M | + 0.02 M |
| F | 0.02 M | 0 | 0.06 M |

pH = $PK_a$ + Log $\frac{[A-]}{[HA]}$
= 4.74 + Log $\frac{(0.06\ M)}{(o.o2\ M)}$
= 4.74 + 0.48
= **5.22**

**Change of pH for buffer and base:**

$\Delta$ pH = $pH_{final} - pH_{initial}$
= 5.22 – 4.74
= **0.48**

**Conclusion:**

le châteliers principle clearly states that if a system at equilibrium was exposed to a stress, the system would shift as to relieve that stress. A stress that was applied on a system at equilibrium was the modification of the concentration of any of the components. In the experiment, the equilibrium shift took place due to the change in concentration, but then it balances when equilibrium (same amount of concentration on both sides) was reached. The same rule was applied to multiple equilibria; only differs in a way that the reaction was happening sequentially. The buffer system was used to maintain any changes in pH when an acid or base was added. In the buffer system, the pH seemed to be contained in a controlled system; whereas the distilled water faced significant difference in terms of pH levels. A common-ion effect took place due to the fact that there was an increase in ion concentration, causing the equilibrium to shift to accommodate that stress. Another type of stress that could be applied on a system is temperature. The temperature effect on a system can create an endothermic reaction. Due to the absorption of heat by the reaction; equilibrium would shift right to create more of that product.

Experiment 2: Equilibria

**Observation table:**

## Equilibrium Shift

(1) $CuSO_4$ → Light blue colour
20 drops → Liquid
→ no smell

(2) $NH_3$ → clear liquid
4 drops → strong smell

(1)+(2) → translucent dark deep blue
(3)

(4) HCl → strong smell
0.1 → clear liquid
28 drops

(3)+(4) → transparent liquid

repeat Step 2 and 3

when added $NH_3$ → translucent dark color — 3 drops

when added HCl → transparent colour
→ original colour — 6 drops

## Multiple Equilibria

(1) $Na_2CO_3$ → clear liquid, tranparent
10 drops → No odour

(2) $AgNO_3$ → tranparent
10 drops

(1)+(2) → brown
(3) → precipitate at bottom

(4) $HNO_3$ → Strong odour
4 drops → clear liquid

(3)+(4) → clear again
(5) → ~~little amount of percipitate~~
→ test tube warmer

(6) HCl → tranparent - clear liquid
6 drops → Not strong odour

(6)+(5) → percipatate forms
(7) → white colour - cloudy

(8) $NH_3$ → strong odour → clear liquid
4 drops

(8)+(7) → clear liquid (transparent)
(9)

13 drops of $HNO_3$ } step 7 and 9
7 drops of $NH_3$

repeating step 7

(8+7) + 4 → thin white cloud
→ heat (warm test tube)
30 drops

repeat step 9

(8+7)+4+8 → No colour change - clear again
9 → test tube warmer

ten
(10) KI → transparent liquid
4 drops → No strong odour

(9)+(10) → white colour - cloudy
(11) → bottom is clear, doesn't mix with the cloudy area
→ percipitate formed

(12) $Na_2S$ → clear liquid
→ not strong odour

(11)+(12) → brownish/grey colour
4 drops → still a small division at bottom of test tube (they don't mix)
→ Pecipatate at bottom of test tube

Common-ion effect

~~Preparation and Testing of Buffer Solution~~

(1) $CuCl_2$ → pink colour
→ No strong odour

(2) HCl → strong odour
4 drops → clear liquid

(1) + (2) → Purple colour
(3)

(3) + distilled $H_2O$ (5 drops) → pink colour change (a bit lighter)

(4) $CuBr_2$ → solid
→ shiny black colour

$CuBr_2$ + distilled $H_2O$ (5 drops) → extremly dark brown
↳ additional 10 drops → Limish green
↳ additional 25 drops → Light blue to get to 2 mL

KBr → beig crystals

KBr + distilled $H_2O$ → 25 drops - clear liquid

10 drops of $CuBr_2$ solution from well, mix it with a pinch of $KB_r$ in a test tube
↳ results in olive green Colour

~~Common-Ion Effect~~ Preparation and testing of Buffer solution

(1) $CH_3COOH$ → clear, no odour

(2) universal indicater → organge/red solution
slight chemical smell

(1) + (2) → when mixed it creats a small well in the center and smooths over to cover the empty space
(3)
↳ slight arange/redish Colour, slight smell (Liquid glue?)

* based on PH chart; PH of solution is 2.5

(4) $NaCH_3OO$ → clear liquid
No odour

(3) + (4) → made orange colour lighter

* pH of (3) + (4) is 2.5 pH

(5) distilled $H_2O$ + universal indicater = 6 pH

(6) HCl 0.1 mol/L → clear liquid
not strong smell

HCl added to the buffer
↳ PH of ~~3~~ 5 → orangy resembles clear grape fruit
↳ PH of $H_2O$ is ~~6~~ 3 → Light red

2nd buffer added NaOH → a little lighter than before (grape fruit Light colour)
↳ PH is 4
↳ adding NaOH to $H_2O$ turns dark blue/purple
pH = 12

## Temperature Effect

the olive green solution (solution of $CuBr_2$ $Sol^n$ and pinch of KBr)

put in a hot water bath → turned extremly dark brown

↳ when $H_2O$ was added to the test tube in the $H_2O$ bath, the solution turned Light blue.

20 drops

↳ 20 drops added to $CoCl_2$ in a clean test tube then put in a hot bath

Colour turns winy colour, however blue colours are shown on the edge of test tube.