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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 (circle): Tues aft Tues night Wed Thurs aft Thurs night Fri

Lab Week (circle): 1 2

Laboratory Report Form

Experiment 4.

Equilibria

Checklist:

○ **Raw Data Sheet written in pen, signed by TA and attached**

○ **Completed Report Form attached**

Student's Initials B.A.

Table 1. Observations and Discussion

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

$CuSO_4 \rightarrow Cu^{2+} + SO_4^{2-}$
$Cu^{2=} + H_2O \rightarrow Cu(H_2O)_4]^{2+}$

a) The colour that was observed was a light transparent blue colour. The copper sulphate in the presence of water turns to tetraaquacopper(II). Initially the Copper sulphate ions break down to for Cu and SO_4 ions. Then the copper ions then combine with the water molecules to form $[Cu(H_2O)_4]^{2+}$. This solution has the pale blue colour seen by the eye.

b) When NH_3 was added to the Solution, the solution had transitioned into a dark purple colour. Initially, the addition of the ammonia had formed a precipitate in the solution. This precipitate was copper(II) hydroxide. Since excess ammonia was added, it had transitioned to a dark purple solution without a precipitate. This was because the solution had formed tetraamminodiaquacopper(II) which as a deep purple colour. The precipitate had left the solution because the ammonia had joined the tetraaminecopper(II) which changed the shape of the compound which turned it aqueous and dark purple in colour.

c) $HCl \rightarrow H^+ + Cl^-$
$H^+ + NH_3 \rightarrow NH_4^+$

Adding HCl turned the solution clear. The Solution had shifted the the equilibrium back to the reactants eliminating the deep purple colour of the tetraamminodiaquacopper(II).

4. When ammonia was added again to the solution, it had shifted the back towards the products forming the product of tetraaminecopper(II). This brought back the colour purple seen in the solution previously. When the HCl was added again, it shifted the equilibrium again back towards the reactants. This was seen as the solution had turned clear again.

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

d) The original colour of the sodium carbonate was clear.

e) The solution had formed a precipitate when silver nitrate was added to the solution. This is because the reaction formed silver carbonate and sodium nitrate. In the reaction the Silver from the silver nitrate joins the carbonate from the sodium carbonate and formed the insoluble silver carbonate molecule. From the solubility rules, all carbonates are insoluble except the ones with sodium, potassium and calcium.

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

f) $HNO_3 \rightarrow H^+ + NO_3^-$
$Ag_2CO_3 \rightarrow Ag^{2+} + CO_3^{2-}$

The addition of the nitric acid had turned the appearance of the solution back to a clear colour. This is because the nitric acid was able to react with the silver carbonate to from bicarbonate. This addition shifted the equilibrium to favour the reactants.

Ag^{+} (aq) + Cl^{-} (aq) ≈ AgCl (s)

HCl -> H^{+} + Cl^{-}

g) When the addition of HCl occurred, the solution turned to a cloudy white solution. According to solubility rules, chlorine will precipitate with the joining of a silver ion. In this particular reaction, the silver from the silver nitrate had joined the chlorine from the hydrochloric acid which lead to the precipitation of AgCl. The remaining ions formed nitric acid.

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

h) Adding the concentrated NH_3 turned the cloudy solution to a clear solution. Since the precipitate that was present was the silver chloride. So the addition of the ammonia reacted with the silver chloride molecules in the solution. The two molecules reacted to form silver (II) nitrate and hydrochloric acid. Both of the new molecules are soluble.

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

$HNO_3 \rightarrow H^{+} + NO_3^{-}$

2) When nitric acid was added, the solution became whitish and cloudy. The contents became warm meaning that the reaction was exothermic. From the presence of hydrogen ions in the solution, the hydrogen ions react with the ammonia ions to form ammonium. By adding the nitric acid, it removes the ammonium that had originally dissolved the precipitate, hence the precipitate is formed again. When the ammonia was added, the solution becomes colourless and transparent. This is because the addition of the ammonia had removed the precipitate.

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

$KI \rightarrow K^{-} + I^{+}$

i) By adding the potassium iodide, the solution turns a cloudy off white. In the reaction stated above, the potassium iodide had split during decomposition and the iodide ion had joined a silver ion that was present in the solution. When the molecule of silver iodide is formed, it appears as a precipitate. This molecule is insoluble.

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

$Na_2S \rightarrow Na^{+} + S^{2+}$

j) By adding the sodium sulphate to the solution, the solution turned to a dark brown/grey colour which was opaque. Adding the sodium sulphate to the solution allowed silver ions to join with the sulphur ions from the sodium sulphate. These silver ions would have been from the initial solvent of silver nitrate.

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

$CH_3COOH \rightarrow H^{+} + CH_3Coo^{-}$

k) The the colour universal indicator was red. This colour would relate to an acidic solution. On the pH scale, the colour red is stating an acidic solution.

l) The colour of the pH paper was a mild orange colour which on a pH scale would represent a pH of 3. The two pH's correspond meaning indeed the solution was acidic.

m) By adding the salt $NaCH_3COO$, the solution states that the solution is still acid. The solution was still a reddish colour. The products of the reaction would be H_3O^{+} and acetate. The addition of the salt sodium acetate is used to increase the concentration of the acetate ion and ultimately makes a buffer system.

n) When pH paper was used to measure the pH, the paper turned yellow indicating a pH of 5. Since the solution was still reddish and the paper hinted at an acidic solution, they do in fact match up.

o) When the indicator was added to water, the solution was a clear red colour. This would mean the solution was acidic. When tested with the pH paper, the paper was yellow and had a value of about 6 on the pH scale.

p) When the acid was added to both solutions, the solution with water turned to a darker shade of red and the buffer solution stayed the same. (No visible change.) The pH of the water solution when the acid was added should make the solution more acidic then before and when the acid is added to the buffer solution, the solution should stay the same because a buffer solution is a constant pH.

q) The pH of the water solution was 3 and the pH of the buffer solution was 4. These both match up with their respective colours for the indicator.

r) The addition of the base had turned the colour of the water solution a purple colour. The addition of the base in the buffer solution had no visible affect on the colour of the solution. This is what to be expected for each of the tests.

s) The pH of the buffer was 4 and the pH of the water solution was about 11.5. These results match the results indicated in point r).

3. For part 3. Refer to ***Calculations***

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

t) At room temperature, the cobalt(II) chloride is a red coloured solution.

u) By adding conc. HCl, the solution changes from red to a dark blue colour. In the solution before the the HCl was added, the cobalt(II) chloride went through a displacement with water to form $[Co(H_20)_6]^{2+}$. This compound then reacted with the chloride from the conc. HCl and produced $[CoCl_4]^{2-}$ and water. The blue colour that appears is the cobalt(II) complex ion of $[CoCl_4]^{2-}$. In this reaction chlorine ions and $[Co(H_20)_6]^{2+}$ are in equilibrium with $[CoCl_4]^2$ and water molecules.

v) By adding the water, the solution turns to a light pink colour. Since in this equilibrium reaction two cobalt(II) complex ions are present, one pink and one deep blue, it would show that by adding the water shifted the equilibrium towards the pink complex ion. By looking at the reaction, by adding the water it should theoretically shift the equilibrium towards the reactants. In this case, the pink complex ion was $Co(H_2O)_6]^{2+}$, which is one one of the reactants. So it can be concluded that the addition of water shifts the equilibrium to favour the reactants.

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

w) The copper(II) bromide solid had a metallic dark grey/black colour. It had resembled gun powder.

x) The addition of water was able to begin dissolving the solid into a very faint blue colour solution. As it was dissolving, it appeared as though the solid was rusting until it became aqueous. This process occurred because by adding water, the water was able to break apart the copper(II) and the bromine ions. From the reaction stated above, it appears that the bromine ions were present in the new solution as well as $[Cu(H_2O)_4]^{2+}$.

y) When the extra water was added, the solid was almost completely dissolved. The liquid the solid is present in had a faint blue tinge to it.

z) When adding water until the volume reached two mL, the solid completely dissolved and the remaining solution had a very faint blue tinge to it (negligible). The water was used to break down the bonds of the copper(II) bromide so that the equilibrium of 4 Br - (aq) + [Cu(H2O)4]2+ (aq) [CuBr4]2- (aq) + 4 H2O (l).

aa) The colour of the dissolved potassium bromide is colourless and clear. This would mean that complex ions were clear in the solution.

bb) $KBr \rightarrow K^{+} + Br^{-}$

When the salt was added to the copper(II) bromide solution, there was no physical change in the of the initial solution. The salt had dissolved in the solution. This is because all the products in the solution present are all aqueous

cc) When the contents from step 27) are heated, the solution very faintly becomes blue. This would mean the reaction was favouring the reactants because the product that has the bluish characteristic is the $[Cu(H_2O)_4]^{2}$

dd) When the cobalt(II) chloride was heated up, the colour transitioned into a dark red colour. By heating this equilibria, the equilibrium shifted towards the products. Since the colour went to a darker shade of red then the initial red seen in observation t) , the formation of $[CoCl_4]^2$ was more dominant then before.

Calculations:

1) pH of the buffer solution
2) pH of the water solution
3) pH of the buffer solution when the acid was added
4) pH of the water solution when the acid was added
5) pH of the buffer solution when the base was added
6) pH of the water solution when the base was added

$Ka=1.8x10^{-5}$ (dissociation constant)
$pH=pKa+log([A^+]/[HA])$ pKa = 4.74
C1V1=C2V2

1) C1V1=C2V2

$[CH_3COOH]=[0.1][0.5/1]$
=0.05 mol/L

$pH= -log1.8x10^{-5} + log[0.05/0.05]$ CH_3COOH
=4.7 pH for CH_3COOH

2) Since distilled water was used, it can be assumed that the pH of the water was initially 7.

3) $CH_3COO^- + HCl \rightarrow CH_3COOH + Cl^+$

$[CH_3COO^-] = [0.05](0.001)/0.00125$
=0.04 mol/L CH_3COO^-

$[HCl]=[0.1](0.00025)/(0.00125)$
=0.02 mol/L HCL

$[CH_3COOH] =[0.05](0.001)/(0.00125)$
=0.04 mol/L CH_3COOH

$CH_3COO^- + HCl \rightarrow CH_3COOH + Cl^+$

I,	0.04	0.02	0.04	…
C.	-0.02	-0.02	+0.02	…
E.	0.02	…	0.06	…

pH = pH=pKa+log([A^+]/[HA])
= 4.74+ log([0.02]/[0/06])
= 4.26 pH of the buffer solution.

4) $H_2O + HCl \rightarrow H_3O^+ + Cl^-$

C1V1=C2V2
C = [0.1](0.00025)/0.00125
=0.02

pH=-log[H^+]
=-log[0.02]
=1.70 pH of the water solution

5) $CH_3COO^- + NaOH \rightarrow CH_3COOH + Na^+$

[CH_3COO^-] = [0.05](0.001)/0.00125
=0.04 mol/L CH_3COO^-

[NaOH]=[0.1](0.00025)/(0.00125)
=0.02 mol/L NaOH

[CH_3COOH] =[0.05](0.001)/(0.00125)
=0.04 mol/L CH_3COOH

$CH_3COO^- + NaOH \rightarrow CH_3COOH + Na^+$

I,	0.04	0.02	0.04	…
C.	-0.02	-0.02	+0.02	…
E.	0.02	…	0.06	…

pH=4.74+log[0.06/0.02]
=5.21 pH for buffer with the added salt.

6) $H_3O^+ + NaOH \rightarrow H_2O + Na^+$

C1V1=C2V2
C=[0.1](0.00025)/0.00125
$[OH^-]$=0.02 mol/L

pOH=-log[0.02]
=1.70
pH=14-pOH
pH=14-1.70
pH=12.30 of the solution of water with the added base.

Conclusion:

In conclusion, by conducting multiple experiments using various equilibria, n the reactions conducted, it was determined that the temperature of the solution can affect the equilibrium. If a common ion is added to the solution, the equilibrium will shift towards the side that is lacking the added common ion. Buffers are solutions that are at constant equilibrium. Since the pH was not significantly changed, the adding of different solvents does not effect the equilibrium. However, non buffer solutions when solvents are added, will have the equilibrium shift towards the sides that are lacking the common ion from the solvent. It was also noticed that a solution can have multiple equilibria. Finally, the big picture of the lab was that the equilibrium can be shifted to favour the products or the reactants using different factors, such as heat, addition of common ions, or the amount of a reactant or product that was initially used meaning too much product shifts equilibrium towards the reactants. The factors experienced in this lab are not the only factors that can effect the equilibrium. Things such as pressure and volume can also effect the direction in which the equilibrium will shift.

Bibliography:

Clark, Jim . " Complex Metal Ions- Ligand Exchange Reactions". Chemguide.co.uk. 2003. Modified 2014. Web. Date of Access. November 19th, 2015.