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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 AC

Table 1. Observations and Discussion

Equilibrium shift

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

The $CuSO_4$ (aq) solution was a clear light blue color and was odorless. The color of this solution is a result of the dissolution of $CuSO_4$ in water to produce the ions Cu^{2+} and SO_4^{2-}. $SO_4^{2-}$is a relatively strong acid, therefore it will almost completely dissolve in water. This reaction forms the compound $[Cu(NH_3)_4]^{2+}$, which gives the light blue color to the solution.

The ammonia NH_3 solution exhibited a potent odor and was colorless. When two drops of this solution were added to the $CuSO_4$(aq) solution the color quickly became dark blue, nor see through or murky, but exhibited a slightly less potent smell. This was caused by the Cu^{2+} ions reacting with the ammonia to form the compound $[Cu(NH_3)_4]^{2+}$ (aq) which gave the dark blue color to the solution.

The HCl(1M) was a colorless solution that exhibited a slightly potent smell. When 31 drops of this acid were added to the aforementioned solution, the color shifts back to light blue. This happened because as HCl is a strong acid and when it reacts with the ammonia it forms $NH_4^+$and Cl^-. This reaction causes the original amount of NH_3 to decrease resulting in a shift to the left on the above equilibrium. The shift accounts for the deficit in NH_3 and makes the equilibrium produce more ammonia, according to La Chatelier's Principle. More reactants are produced including the $[Cu(NH_3)_4]^{2+}$ making the color light blue again.

If these steps were repeated, the color would shift back to dark blue when ammonia is added because the same ions will cause the solution to shift dark blue again. When more HCl drops are added, the color returns to a light blue.

Multiple equilibria

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

The $AgNo_3$ solution (also called silver nitrate) was a clear colourless and odourless solution. The $NaNO_3$ (also called sodium nitrate) was also a clear, colourless and odourless solution. When both were combined in a test tube, the solution became murky, cloudy, beige and odourless. This colour is due to the solution dissociating into the ions Ag^+, NO^{3-},

$2Na^+$ and CO_3^{2-}. The sodium ions and the nitrate ions then fuse together forming a beige coloured aqueous compound.($NaNO_3$)

The HNO_3 solution was clear, colourless and smelt a little bit like plastic, but was practically odourless. When two drops were added to the previous solution, it rapidly became colourless, murky and with a small precipitate. As for the odour, it became slightly potent (like after bite). Adding more HNO_3 into the solution shifted the equilibrium to the left due to the creation of more NO_3^- ions. According to La Chatelier's principle this shift caused the equilibrium to produce more of the reactants to balance out the sudden burst of NO_3^- ions. This also caused the solution to revert to its original compounds: $AgNO_3$and Na_2CO_3 both at an aqueous state.

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

The 0.1 M HCl was a clear, colourless and slightly potent aqueous solution. When 4 drops were added to the aforementioned solution, the solution becomes cloudy, murky, white and the solution becomes odourless. The H^+ ions from the HCl react with the CO_3^{2-} ions from the Na_2CO_3 to form a new compound, H_2CO_3. Which then is dissociated into H_2O (l) and CO_2 (g). Again according to La Chatelier, the solution's equilibrium shifted to balance out the amount of reactants and the amount of products, explaining the murky white colour.

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

The Cl^- ions from the aqueous hydrochloric acid are attracted to the Ag^+ ions from the aqueous $AgNO_3$ to form a solid compound (AgCl). This red-brownish precipitate can be found at the bottom of the test tube containing the cloudy, white solution. The formation of the precipitate proved that it did not dissolve completely in the aqueous solution. When 4 drops of HCl were added, the amount of Cl^- ions increased, this caused a decrease in the amount of Ag^+ ions in the $AgNO_3$ compound. The reaction shifted to produce more $Na_2CO_3$and $AgNO_3$. Their Na^+, CO_3^- and NO_3- ions could have combined with the H^+ and Cl^- ions to produce NaCl(s), H_2CO_3(aq) and HNO_3^- (aq). The NaCl compound could explain the precipitate as it is in a solid state.

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

The clear, colourless and very potent ammonia was added (3 drops) the solution in the test tube and the solution became clear, colourless and smelt very potent (like after bite). Also, a small red-brownish spec remained at the bottom of the tube. This precipitate could be because the reaction wasn't completed, or not enough ammonia was added, and is probably not significant. The Ag^+ ions from the $AgNO_3$ reacted with the NH_3 when the equilibrium caused the reaction to shift because of the Cl^- ions, which are also present in the solution. This caused the amount of $AgNO_3$ ions and the amount of Ag^+ to become more abundant. The product of the Ag^++ NH_3 reaction is $[Ag(NH_3)_2]^+$. This compound causes the reaction to become clear and colourless again.

The excess HNO_3 in the solution reacted with the NH_3 causing a decrease in the amount of NH_3. This caused a shift in the equilibrium as $[Ag(NH_3)_2]^+$and more NH_3 was produced to compensate for a small lack of NH_3. This transfer caused the NH_3 to become in excess causing the reaction to shift again to produce more $[Ag(NH_3)_2]^+$, this is the same compound that causes the colourless solution in the previous step. Therefore, the solution remains colourless.

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

Adding HNO3 (aq) to the clear, colourless and slightly potent solution causes an increase in H^+ ions. The already present ammonia reacts with this ion and shifts the reaction to the right. The NH_4+(ammonium) ion is created this way. As NH_3 is a base that accepts protons, it will force the reaction to produce more products. When HCl (0.1 M) is added to the solution, the number of H^+ ions increased, in return causing in increase in NH_4 ions. This reinforced the equilibrium, but did not cause it to shift. When the addition of HNO_3 is repeated, it helped the equilibrium continue its shift to the right, forming more ammonium and causing the solution to remain colourless. Again, the $[Ag(NH_3)_2]^+$ ion keeps the solution colourless.

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

In this reaction, 2 drops of the clear, colourless and odourless KI (potassium iodine) was added to the clear, colourless and potent solution. This produced a cloudy, light yellow, egg smelling solution. The production of the solid AgI ion caused the light yellow coloration. As the Ag^+ and the I^- ions bind together and are produced in excess; the equilibrium of the solution shifts to the left to balance the solution, after producing an excess in AgI. The solution is cloudy, so it indicates the presence of a precipitate. In this case, the Ag^+ and the I^- ions bind together forming a solid compound AgI, which also causes the light yellow colour of the solution.

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

The addition of one drop of the clear, colourless and odourless Na_2S to the solution causes it to become a murky, dark-grey-brown and smelt slightly potent. The strong reaction between the Ag^{+} and the S^{2-} ions caused the equilibrium to shift to the right. The murkiness and the colour of the solution are caused by a precipitate (Ag_2S) that is not completely dissolved in the aqueous solution.

Preparing and testing of a buffer solution

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

The CH_3COOH (aq) (0.1M) solution is a clear, colourless that smells acidic (vinegar) and has a pH of 4.7. The universal indicator is a murky, red-orange solution that smells like sugar. When both are mixed together, the solution is a clear red colour and smelt like both sugar and vinegar and the pH was found to be acidic (3). As the universal indicators' pH is neutral, it proves that the CH_3COOH compound is acidic.

When the clear, colourless and slightly vinegar smelling $NaCH_3COO$ is added to the solution, the colour becomes reddish-orange and smells sugary. The pH was measured at 4, which is still acidic. The calculated pH is 4.7. This decrease in the acidity is expected as $NaCH_3COO$ is a base.

When clear, colourless, scentless water is added to the universal indicator, the solution stays translucent but becomes, red, smells like sugar and has a pH of 6. When HCl (0.1M) is added to the water based solution, the pH becomes more acidic at a 4 on the scale. When HCL is added to the buffer solution, the pH remains at 4. The calculated pH of the water solution+ HCl is 1.7 and the calculated pH of the buffer solution + HCl is 4.26. The pH of the buffer doesn't change as easily because it consists in a weak acid (CH_3COOH) and a base ($CH_3COO^{–}$) from the $NaCH_3COO$ that counter act the strong HCl acid. The CH_3COOH^{-} ion accepts a proton from the H^{+}(from HCl) preventing the H^{+} ion from making the solution more acidic. This principle shows why the buffer remained practically constant, while the pH of the water solution became increasingly acid when the HCl is added.

When NaOH (0.1 M) was added to H_2O+UI the colour became light orange, and the pH was measured at 12, which is basic. When this base was added to a new buffer solution, the colour became dark purple and the pH was measured at 5. The calculated pH of the water solution+ NaOH is 12.3 and the calculated pH of the buffer solution+ NaOH is 5.23. The huge increase in the water solution's base level indicates the NaOH is a strong base.

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

The $CoCl_2$ is a clear, pink and scentless solution. This colour is due to the dissociation of the compound in water ($CO^{2+)}$. When hydrochloric acid is added, the colour of the solution becomes purple and smells slightly potent. The excess of Cl^- from the HCl causes the reaction to shift to the right, producing more $CoCl_4^{2+}$ ions. This ion causes the purple colour. When more water is added to the solution, the colour becomes pink and doesn't really smell anything. The excess water causes the equilibrium to shift back to the left. H_2O and $CoCl_4^{2-}$ are consumed to produce Cl^- and $[Co(H_2O)_6]^{2+}$ causing the pink colour.

The effect of temperature: when the solution is placed in boiling water, the original pink colour at room temperature becomes blue. The heat, which is on the reactant side causes the equilibrium to shift the right to balance the heat on both sides of the equation. This shift causes more $CoCl_4^{2+}$ to be produced giving the blue colour to the solution.

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

$CuBr_2$(s) is made of shiny dark grey specs that smell metallic. When this solid is dissolved in water, it becomes a clear, vibrant green colour and is odourless. The $[Cu(H2O)_4]^{2+}$ compound was created when the ions Cu^{2+} and H_2O reacted together and it gives the solution its bright green colour. The reaction between $[Cu(H2O)_4]^{2+}$ and the Br^- ions produced the aqueous solution $[CuBr_4]^{2+}$ makes the solution turn a brownish, slightly green colour. When more water is added, the equilibrium shifts to the left producing $[Cu(H2O)_4]^{2+}$ and Br^- ions which gives the green colour to the solution.

The KBr is composed of a lot of white specs and couple of brown ones and does not have a particular scent. When mixed with water, the solution is colourless and smells slightly acidic. This shows that the KBr dissociates into its components (K^+ and Br^-) but does not create a compound with the H_2O. When CuBr is added to this solution, the colour becomes a vibrant green and it smells like salt. The increase in the Br^- ions forces a shift in the equilibrium to the right, which creates $[CuBr_4]^{2-}$. As seen previously, this ion causes the green colour.

When placed in boiling water, the solution becomes kaki(brown-green) and is murky. This proves that the heat was part of the reactants and caused a greater amount of products $[CuBr_4]^{2-}$and water, as some of the original water evaporated (due to the heat). Also, this caused a decrease in the amount of $[Cu(H2O)_4]^{2+}$ being produced, added with the brownish colour of the $[CuBr_4]^{2+}$ ion, the colour of the solution will become kaki.

Calculations:

pH of water: 7

pH of buffer: $CH_3COOH/CH_3COO^-Na^+$

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

10 dr $CH_3COOH(aq)$= 0.1mol/L
$[CH_3COOH]$= (0.1 mol/l)(10dr/20dr)
= 0.05 mol/L

10dr CH_3COO^- (aq)= 0.1mol/L
$[CH_3COO^-]$= (0.1 mol/L) (10dr/20dr)
= 0.05 mol/L

$Ka(CH_3COOH)= 1.8x10^{-5}$

pH= pKa+ log ($[A^-]/[HA]$)
=-log ($1.8x10^{-5}$)+ log ((0.05mol/L)(0.05 mol/L))
= 4.7 pH of CH_3COOH

pH of H_2O + HCl:

Vi of HCl= 5dr
Ci = 0.1 mol/L

vi of H_2O= 20 dr
$V_{final(solution)}$= 25 dr

Ci Vi = Cf Vf

(o.1 mol/L) (5 dr)= C_{final} (25 dr)

C_{final}= 5/25 (0.1 mol/L)
= 0.02 mol/L

pH= - log [HCl]
= - log (0.02)
= 1.70 H_2O+HCl

pH H_2O+ NaOH:

NaOH= 0.1 mol/L
V_{NaOH}= 5 dr
V_{total}= 25 dr

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

pOH= -log $[OH^-]$
= - log0.02
= 1.7

pH=14.00-pOH
= 14.00- 1.7
= 12.3 pH of NaOH+ H_2O

pH of buffer+ HCl:

$CH_3COO^-(aq)$+ HCl(aq)--------> $CH_3COOH(aq)$ + $Cl^-(aq)$

V_{total}= 25dr

CH_3COO^-= 10 dr at 0.10 mol/L

(0.10mol/L) (10dr/25dr)= 0.04 mol/L

HCl= 5 dr at 0.10 mol/L

(o.10 mol/L) (5dr/25dr)= 0.02 mol/L

CH_3COOH= 10 dr at 0.10 mol/L

(0.10mol/L) (10dr/25dr)= 0.04 mol/L

CH_3COO^- +	HCl-------->	CH_3COOH + Cl^-
0.04mol/L	0.02mol/L	0.04mol/L
-0.02	-0.02	+0.02
0.02mol/L	0.00 mol/L	0.06mol/L

pH= -pKa+ log([A^-]/[HA])

= - log ($1.8x10^{-5}$)+ log(0.02M/0.06M)

= 4.74- 0.28

= 4.26 pH of buffer + HCl

pH of buffer+ NaOH:

NaOH(aq)+ $CH_3COOH(aq)$----> $CH_3COO^-(aq)$ +Na^+ (aq)+ $H_2O(l)$

V_{total}= 25 dr

CH_3COOH= 10 dr at 0.10 mol/L

(0.10mol/L) (10dr/25dr)= 0.04 mol/L

CH_3OOH^-= 10 dr at 0.10 mol/L

(0.10mol/L) (10dr/25dr) = 0.04 mol/L

NaOH= 5 dr at o.10 mol/L

(0.10 mol/L) (5dr/25dr)= 0.02 mol/L

NaOH(aq)+ $CH_3COOH(aq)$----> $CH_3COO^-(aq)$ +Na^+ (aq)+ $H_2O(l)$

0.02 mol/L	0.04 mol/L	0.04 mol/L
-0.02	-0.02	+0.02
0 mol/L	0.02 mol/L	0.06 mol/L

Kw= (Ka)(Kb)

$1.0x10^{-14}$= $1.8x10^{-5}$ Kb

Kb= ($1.0x10^{-14}$)/($1.8x10^{-5}$)

= $5.6x10^{-10}$

pOH= -Kb+ log [HB]/[B^+]

= -log $5.6x10^{-10}$ + log(0.02M/0.06M)

= 9.25- 0.48

= 8.77

Ph= 14.00- POH
= 14.00- 8.77
= 5.23pH of buffer+ NaOH

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

La Chatelier's principle was not describe in the observations, so it will be covered here. Also named equilibrium law, it can be used to predict changes in chemical reactions. An equilibrium will change or shift when it is subjected to change. These changes include, temperature, volume, concentration and pressure. the system adjusts to the changes and forms a new equilibrium.

Sources of error: As the procedures did not always specify the exact amount of drops that were needed, we might have added too little or too many drops for certain steps. This can slightly change the colour, murkiness or odour of the solution observed. The results would not change radically, but as this was a qualitative lab it might alter our results. Also because this was a qualitative lab, we might not have observed the same colour or smelt the same thing as other people. For example, we might have a purple solution in our observations, but for other people that same solution might have seemed blue.

Conclusion:

In this experiment, we learnt that adding an acid or a base can shift the equilibrium. It shifts according to which side of the equation is favoured, the right or the left. The experiment on multiple equilibria showed us that the equilibrium will shift continually when different reactions are taking place in the same solution. The buffer section showed us the resilience of a buffer when adding an acid or a base compared to a water-based solution. Also, I was found that the common ion effect will shift an equilibrium to the right to increase the amount of product. Lastly, when heat is added to a solution, the equilibrium will favour the reactants or the products. It will favour the side that exhibits the less heat, shifting the equilibrium from the side that has excess heat.

Audrey Chamberland

CuSO4 observations

colour: before reaction = light ~~clear~~ transparent blue
odor: before reaction = odorless
appearance: before reaction = see through

Exp. 1 **NH3 observations**

colour: ~~to clear~~ colourless
odour: potent (corrosive), smells like after ~~at~~ bike
appearance: seethrough

CuSO4 + NH3 observations added 2 drops of NH3

c = dark blue
o = slightly smells like the NH3, but semi odourless
a = between see through and murky

HCl lead observations ~~[illegible]~~

c = ~~[illegible]~~ colourless
o = ~~odourless~~ slightly potent
a = see through

CuSO4 + NH3 + HCl 31 drops

c = very light blue almost colourless
o = odourless
a = see through

Na2 CO3

Exp. 2 c = colourless
o = odourless
a = see through

AgNO3 (bottle was wrapped in tin foil)

c = colourless
o = odourless
a = ~~murky~~ see through

Na2CO3 + AgNO3

c = beige
o = odourless
a = cloudy murky

HNO_3

C= colourless

O= ~~[illegible]~~ smells a little bit like plastic but practically odourless

a= see through

$NaCO_3 + AgNO_3 + HNO_3$ (+2 drops added)

C= colourless

O= slightly potent, kind of like after bite

a= because of precipitate it's slightly cloudy

HCl 0.1 mol/L

C= colourless

O= slightly potent

a= see through

$NaCO_3 + AgNO_3 + HNO_3 + HCl$ (+4 drops added)

C= white + red-brownish precipitite at the bottom

O= odourless

a= cloudy, murky, precipitate at the bottom

NH_3 (+3 drops added)

C= colourless

O= potent, smells like after bite

a= see through

$NaCO_3 + AgNO_3 + HNO_3 + HCl + NH_3$

C= colourless

O= very potent, smells like after bite

a= see through, very little red-brown precipitaz at the bottom

KI 0.1 mol/L

C= colourless

O= odourless

a= see through

$NaCO_3 + AgNO_3 + HNO_3 + HCl + NH_3 + KI$ (+2 drops added)

C= pale yellow

O= eggs, strong, potent

a= murky, cloudy

Na_2S

C= colourless

O= odourless

a= see through

$NaNO_3 + BaNO_3 + HNO_3 + HCl + NH_3 + KI + Na_2S$ (+1 drop added)
c= dark, gray, brown
o= potent (slightly)
a= murky, cloudy

Exp 3 Buffer Solution

CH_3COOH
c= colourless
o= acidic, smells like vinegar
a= see through

Universal indicator
c= red-orange
o= smells like candy, sugary
a= murky

CH_3COOH + UI
c= red-pink
o= mix between sugary and vinegar (acidic)
a= see through
pH = 3 (acidic) for both wells

$NaCH_3COO$
c= colourless
o= slightly smells like vinegar
a= see through

CH_3COOH + UI + $NaCH_3COO$
c= orange-red
o= sugary
a= see through
pH = 4 (acidic) for both wells

H_2O + UI
c= red-pink
o= sugary
a= transparent
pH = 6

H_2O + UI + HCl 0.1 mol/L

c = turned a little bit redder than previous pH = 3

O = more acidic than sugary

a = transparent

a) The solution with H_2O is slightly more acidic than the buffer solution

CH_3COOH + UI + $NaCH_3COO$ + HCl 0.1 mol/L

C = turned little bit redder than previous solution

O = more acidic than sugary pH = 4

a = transparent

NaOH /

C = colourless

O = odourless

a = transparent

H_2O + UI + NaOH 0.1 mol/L

c = light orange

O = sugary

a = transparent

pH = 12 (basic)

CH_3COOH + UI + $NaCH_3COO$ + NaOH 0.1 mol/L

c = purple (dark)

O = sugary

a = see through

pH = 5 (slightly acidic)

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exp. A

common ion effect

$CoCl_2$ 1 mol/L

c = pink

O = odourless

a = clear (see through)

12 mol/L HCl

c = ~~[illegible]~~ colourless

O = acidic (corrosive)

a = transparent

$CoCl_2$ + HCl (+ 6 drops added)

C = purple blue

O = slightly acidic

a = transparent

$CoCl_2$ + HCl + H_2O (+ 6 drops H_2O added)

c = pink

O = something slightly basic, doesn't smell much

a = transparent

$CuBr_2$

c= dark gray

o= metallic (slightly)

a= shiny, granulated, specs

$CuBr_2 + H_2O$ (+5 drops added)

c= brown, looks like soy sauce slightly green

o= slightly metallic

a= murky and dark

$CuBr_2 + H_2O$ (+10 drops)

c= green (vibrant)

o= odourless

a= clear

KBr

c= white with brown specs

o= odourless

a= granulated

$KBr + H_2O$

c= colourless

o= slightly acidic

a= transparent

$CuBr_2 + H_2O + KBr$

c= vibrant green

o= salty

a= clear

+heat = kaki brown green, murky

$CoCl_2$ + heat = transparent

c= it turns blue, while the solution was pink and transparent at
a= room temperature

o= odourless