Experiment 3: Chemical Equilibria

Introduction:

Equilibrium describes a state where products and reactants have finally reached a balance within the system. The forward and reverse reactions of a chemical reaction are equal so the reaction will appear to stop. Equilibrium reactions can favour the forwards or backwards reaction, which gives different concentrations of reactions or products. The experiments display equilibrium shifts through the addition or removal of species in reactions at equilibrium. Le Chatelier's Principle can effectively used to describe the change in a system at equilibrium. It states: a reaction will always move in the direction that minimizes the effect of any change imposed on a system of equilibrium. The changes could be caused by a variety of factors such as pressure, volume, temperature, and concentrations. The equilibrium constant (K) can be found from the following equation:

$$K_c = \frac{[\,C\,]^c[\,D\,]^d}{[\,A\,]^a[\,B\,]^b}$$

In terms of acid-base equilibria, it involves reversible reactions which means it will eventually achieve equilibrium. Acids donate protons and Bases accept protons. We can measure the equilibrium constants of acids and bases and determine if they are strong (favoring products) or if they are weak (favoring reactants). All acid-base equilibria takes place in water (aqueous solution). This is important because it is amphoteric and can therefore serve as an acid or base. Every equilibrium reaction is composed on an acid, a base, a conjugate acid, and a conjugate base. The strength of acids and bases can be measured by its pH, which lies on a scale between 0 - 14. It measures the concentration of hydrogen ions in a specific solution. A neutral solution (such as water), typically has a pH of 7. A pH value lower than 7 means the solution is acidic and a pH value higher than 7 means the solution is basic.

Buffers play a large role in the concept of acid-base equilibria. Buffers work by either combining a large volume of a weak acid with its conjugate base, or a large quantity of weak base with its conjugate acid. These combinations help resist large changes in pH when there is a strong base or acid added. When there is a strong base added, the acid already present helps to neutralize the hydroxide ions. Consequently, when there is a weak base added, the base already present helps to neutralize the hydronium ions. With this, the pH level is barely altered because they balance each other out. A perfect example of a successful buffer system is the maintenance of blood pH levels in our bodies. It is critical that our blood level maintains a pH level between 7.35-7.45. (Schwalfenberg, 2012) The system uses a carbonic acid buffer, a phosphate buffer, and a plasma buffer in the blood.

In our experiment, several different acids and bases are used in real-life scenarios. Due to our body's fluctuation of pH, it may be necessary to raise or lower pH levels to maintain homeostasis. If our bodies' pH levels drop too low, it is brought back to the ideal level by hyperventilation, allowing more oxygen to enter our blood and more carbon dioxide to be released. This causes the blood to be more alkaline. If the pH level goes over 7.45, hypoventilation is used to reduce the amount of oxygen in our bodies and increase the amount of carbon dioxide. This causes our bodies' pH to become more acidic. When pH levels in our body are raised to the extent where our body alone is not capable of bringing it back down (too basic), a weak acid can be used to bring it back to normal levels. Contrarily, when pH levels in our body are way too low (too acidic), a weak base can be used to bring it back to normal levels.

Works Cited

Schwalfenberg, G. K. (2012, October). The Alkaline Diet. Retrieved October 10, 2018, from https://www.ncbi.nlm.nih.gov/pmc/articles/PMC3195546/

Qualitative Observations:

Table 1: Equilibrium Shifts Observations and Discussion

<table>
<tr><th>Amount</th><th>Chemical (in order of addition to the reaction)</th><th>Observation of Chemical</th><th>Visible Observation of Reaction</th></tr>
<tr><td>20 drops</td><td>$CuSO_{4\ (aq)}$</td><td>Light blue, transparent, clear, liquid solution</td><td rowspan="2">- Dark blue, cloudy, white
- Cloudy precipitate at bottom of test tube</td></tr>
<tr><td>2 drops</td><td>$NH_{3\ (aq)}$</td><td>Transparent, clear, liquid solution</td></tr>
<tr><td>30 drops</td><td>$HCl_{(aq)}$</td><td>Clear, transparent, liquid solution</td><td>- loss of precipitate at bottom of test tube
- cloudy, light blue/teal gradient after mixing</td></tr>
<tr><td>1 drop</td><td>$NH_{3\ (aq)}$</td><td>Transparent, clear, liquid solution</td><td>- forms 2 layers: dark blue colour on top and light blue on bottom
- cloudy solution
- when mixed, becomes all a solid medium blue colour</td></tr>
<tr><td>20 drops</td><td>$HCl_{(aq)}$</td><td>Clear, transparent, liquid solution</td><td>- clear/transparent, gradient-like, layer on top
- cloudy precipitate between clear layer and dark blue layer</td></tr>
</table>

Discussion:

$CuSO_4$ and NH_3 are mixed together and a single displacement reaction is created between the Cu and the NH_3. It can be expressed by the equation:

$$[Cu(H_2O)_4]^{2+} (aq) + 4NH_3 (aq) \rightleftharpoons [Cu(NH_3)_4]^{2+} (aq) + 4H2O (l)$$

The original light blue colour from the $CuSO_4$ is from the Cu+ ions it contains. When the reaction with the NH_3 occurs, the solution turns into a darker blue colour due to the shift of its equilibrium (to the right), creating $[Cu(NH_3)_4]^{2+}$. Adding HCl increases the concentration of H+ ions in the solution and neutralizes the base in the solution (NH_3). This causes the reaction to shift to the left and therefore become light blue and then clear as the excess H+ ions are used up. This reaction is shown in the following equation:

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

There is a higher concentration of NH_4^+ ions (which is clear) and a lower concentration of the $[Cu(NH_3)_4]^{2+}$ (which is responsible for the blue colour).

As these steps are repeated, the same changes took place due to the same change in equilibrium. However, it took a larger amount of the compounds because a higher concentration is now necessary for visible equilibrium shifts to take place. A higher concentration of [Cu(NH3)4]2- is indicated by a deeper blue colour and it increases as more NH3 is added to the solution. On the other hand, a higher concentration of NH4+ ions is indicated by a clear colour and it increases as more HCl is added because of the presence of H+ ions.

Table 2: Multiple Equilibria Observations and Discussion

Amount	Chemical (in order of addition to the reaction)	Observation of Chemical	Visible Observation of Reaction
10 drops	$Na_2CO_{3\ (aq)}$	Clear, transparent, liquid solution	- cloudy, rusty colour (brown/orange)
10 drops	$AgNO_{3\ (aq)}$	Clear, transparent, liquid solution	
3 drops	$HNO_{3\ (aq)}$	Clear, transparent, liquid solution	- clear, transparent, liquid
4 drops	$HCl_{(aq)}$	Clear, transparent, liquid solution	- Cloudy, white - slightly darker/rusty colour at bottom of test tube
3 drops	$NH_{3\ (aq)}$	Clear, transparent, liquid solution	- gas produced - solution is cloudier and less rusty in colour - 3 layers form: a clear layer and a

			cloudy layer on top and below it
2 drops	HNO_3 (aq)	Clear, transparent, liquid solution	- clear, transparent - more gas is formed - condensation along sides of test tube - cloudy precipitate formed at bottom of test tube - after a couple of seconds, whole solution becomes cloudy and gas disappears
2 drops	KI (aq)	Clear, transparent, liquid solution	- gradient: more opaque yellow on bottom, translucent yellow on top - after a couple of seconds, solution turns completely opaque
1 drop	Na_2S (aq)	Clear, transparent, liquid solution	- all one solid colour: dark/muddy brown - opaque and cloudy -white precipitate on bottom of test tube - little brown solid specs along the sides of the test tube where the chemical touched - after a couple of seconds, precipitate disappears

Discussion: When Na_2CO_3 is mixed with $AgNO_3$, a double displacement reaction occurs between the Na and Ag. This can be shown with the equation:

$$2AgNO3\ (aq) + Na2CO3\ (aq) \rightleftharpoons Ag2CO3\ (s) + 2\ NaNO3\ (aq)$$

A muddy brown Ag_2CO_3 precipitate is formed after the double displacement reaction takes place. By adding HNO_3, the solution becomes clear, gas is formed, and the precipitate disappears. The following reaction is shifted to the right:

$$2H+\ (aq) + CO3^{2-}\ (aq) \rightleftharpoons H2CO3\ (aq) \rightarrow H2O(l) + CO2\ (g)$$

The above reaction also created CO2. This change is likely not reversible due to the gas in the atmosphere which removed the CO2. From this loss of CO2, it cannot be converted back into H2CO3, which makes this a non-reversible reaction.

After adding HCl, the solution a white cloudy colour which indicates that the precipitate is still present but there was a colour change. An ionic reaction occurs and a AgCl(s) precipitate is formed:

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

The precipitate is formed due to the dissociation of the strong acid (HCl), creating ions of H^+ and Cl^- in the solution.

The solution becomes clear again following the addition of NH_3. This is because the the NH_3 causes the AgCl(s) to dissociate. The equilibrium shifts to the right and produces $[Ag(NH_3)]^{2+}$ as well as removes the precipitate. The condensation along the sides of the test tube and the heat produced indicates that some NH_3 gas is formed. This can be shown in the following equation:

$$Ag^+ (aq) + 2NH_3\ (aq) \rightleftharpoons [Ag(NH3)]^{2+}\ (aq)$$

The re-addition of HNO_3 again caused the solution to display a cloudy white appearance with some gas being formed. This reaction can be shown by the following equation:

$$H+\ (aq) + NH3\ (aq) \rightleftharpoons NH4+\ (aq)$$

The HNO_3 performed the same as before, forming CO2 gas. However, the cloudiness of the solution indicates that more precipitate has formed. The NH_3 formed with the remaining H+ ions to make NH_4^+.

Before the addition of KI, the test tube felt slightly warm (due to the NH_3). After the addition of KI, the test tube showed a slightly pale yellow-tinted cloudy colour. This reaction can be shown in the following equation:

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

The pale yellow colour change is due to the formation of the precipitate AgI(s). The K+ ions do not have a qualitative effect on the solution.

After having added Na_2S into the solution, it developed a dark muddy brown appearance. The following equation demonstrates a displacement of the iodide with sulfur, which forms $Ag_2S(s)$:

2Ag+ (aq) + S2- (aq) ⇌ Ag2S (s)

The Ag_2S(s) is the precipitate seen (solid dark brown specs) and is responsible for the dark muddy brown colour that was observed. This is because the equilibrium shifts to the right and favors the product (Ag_2S).

Table 3: Buffers Observations and Discussion

Amount	Chemical (in order of addition to the reaction)	Observation of the Chemical	pH Level (after addition to the beaker)	Visible Observation of Reaction
100mL	$H_2O_{(l)}$	Clear, transparent, liquid solution	5.08	NA
2.53g	$NaHCO_{3\ (s)}$	Solid, granulated white substance	8.10	No change
30mL	$HCl_{\ (aq)}$	Clear, transparent, liquid solution	7.18	No change
10mL	$C_3H_6O_{3\ (aq)}$	Clear, transparent, liquid solution	7.00	-bubbles formed around sides on the beaker
Unknown	$CO_{2\ (g)}$	Colourless gas	7.06	No change
0.50g	$NaHCO_{3\ (s)}$	Solid, granulated white substance	7.15	No change
0.50	$NaHCO_{3\ (s)}$	Solid, granulated white substance	7.28	No change
1 pellet	$CO_{2\ (s)}$	Solid, crystal-like, white substance	7.24	-gas formed on top of beaker -cold air
0.40g	$NH_4Cl_{\ (aq)}$	White, granulated, solid substance	7.30	No change

Discussion:

The initial pH of the distilled water is expected to be neutral (7) but it was read as more acidic, with a pH of 5.08, This is different than the expected value because distilled water tends to have very little H+ ions and OH- ions, which can cause an interference with the pH.

The pH then changed from 5.08 to 8.10 after the addition of NaHCO3 because it dissociates into a strong base, NaOH, and a weak acid, H2CO3. The strong base makes the solution more basic. Next, HCl is added to lower the pH, by increasing the concentration of H+ ions and increasing the acidity. The reaction can be expressed as:

$$H_2CO_3\ (aq) + H_2O\ (l) \rightleftharpoons HCO_3^-\ (aq) + H_3O^+\ (aq)$$

HCl dissociates into H+ and Cl- which again, increases the number of H+ ions. This makes the pH level decrease to 7.18. The system used up the H+ ions and then split up H2CO3 into 2 ions; HCO3- and H3O+. The hydronium ions cause the solution to be more acidic. Although after this step the pH is close to neutral water, just slightly more basic, it is still too acidic for human blood. This means is in a state of acidosis.

$$2\ H_2O\ (l) + CO_2\ (g) \rightleftharpoons H_2CO_3\ (aq) + H_2O\ (l) \rightleftharpoons HCO_3^-\ (aq) + H_3O^+\ (aq)$$

The lactic acid being introduced causes a decrease in pH (7.00) which causes H2CO3 to be used up. This shifts the equilibrium to the left which explains the release of CO2 (the bubbles). This is an example of the muscles in the human body when metabolic acids are transferred to the blood. With a release of carbon dioxide, the amount of carbonic acid decreases which causes an increase in hydronium ions.

Next, speed was increased to mimic the body hyperventilating. When the speed of the stirrer increased, a small vortex forms which causes CO2 to form and be released. This has an effect on the pH, an increase to 7.06. This occurs because the system looks for a way to replenish the CO2 which decreases the concentration of hydronium ions.

NaHCO3 is added and acts like a base. It increased the pH level to 7.13 by dissociating into NaOH and H2CO3. This caused the hydronium concentration to decrease, making the concentration more basic.

Then, the pH increased more to 7.28 for the same reasons. By adding more NaHCO3, the solution becomes more and more basic.

By adding more CO2(s), the pH of the solution decreased to 7.24. This shifted the reaction to the right and the release of CO2(g) (gas released on top that was observed) causes the formation of H2CO3. It dissociates and decreases the pH by increasing the concentration of hydronium ions, making it more acidic.

Finally, the addition of NH4Cl should make the pH decrease to about 7.20. Typically, it will dissociate into HCl (strong acid) and NH3 (weak base). Since there is a strong acid, the concentration of H+ ions increases, making the solution more acidic and decreasing the overall pH level. But, in the lab the pH of our solution increased to 7.30 after adding the ammonium chloride. This could have occured because not all of the substances could have fully dissolved since there were visible pieces present at the bottom of the beaker which could throw off the pH level.

Questions:

1. In one step, ammonium chloride (NH4Cl) was added to lower the blood pH. The ammonium ion is what acts as the acid. The chloride ion does not have any acid/base properties.
   a) Why is the ammonium ion used as the acid source instead of HCl?

It is used as the acid source instead because when NH_4Cl separates, it dissociates into HCl and NH3.When the NH_4Cl is used, HCl is still used in the solution since it becomes the strong acid. It then further dissociates into H+ cations and Cl- anions. With an increase of H+ ions, the pH decreases, but the Cl- has no acidic or base properties.

   b) Why is NH4Cl used instead of some other ammonium compound (such as NH4NO3 or NH4I)

These other ammonium compounds form ionic bonds with the other species in the solutions. For example, in part 2 iodide formed a precipitate and the bottom. NO3 also formed a precipitate when displaced with CO3.

2. What observation did you make each time an acidic substance was added to the beaker? Write a general reaction OR use equilibrium arguments to explain this.

The addition of an acidic substance to the beaker caused the pH to decrease. This happened by increasing the concentration of H^+ or H_3O^+ ions.

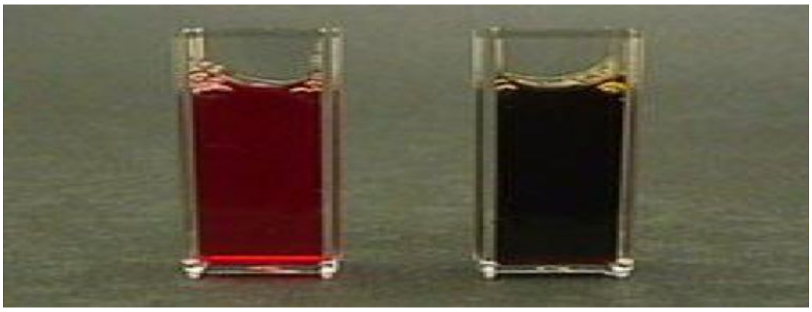
$$\text{Acid substance (aq)} + H_2O\ (l)\ \text{à Conjugate base} + H_3O^+$$

3. The ability of hemoglobin (Hb) to carry oxygen throughout the body as oxyhemoglobin (HbO2) is dependent on the pH of the blood. What effect would acidosis have on the ability of a patient to transport oxygen?

$$HbH^+ + O_2\ \ HbO_2 + H^+\ \square$$

An excess of H+ ions in the blood would cause more oxygen to bind to the ions, which creates H2O instead of displacing the H+ ions from HbH+. This would make the patient's ability to transport oxygen much lower because the oxygen would be binded to these ions.

4. The solution on the left was made by dissolving several drops of blood in some water. The solution on the right was made the same way except that a small amount of HCl was also added to this tube. Based on your general knowledge about the color of blood and the information in question 3, propose an explanation for what happened.



The solution on the right has blood in a state of acidosis. This is because HCl, the strong acid, mixed with the solution which caused the pH to decrease. This happened because the concentration of H+ ions increased. The excess H+ ions will have an affect on the body's ability to carry oxygen which explains the darker colour of the solution on the right.

5. A fresh sample of soda had a pH of 2.92. The soda was placed on a magnetic stirrer and made to go flat. The pH is measured again. Should the pH of the flat soda be higher, lower or the same as the pH of the fresh soda?

The pH of the can will not change if the can is closed because the stirring will not release any CO2 into the atmosphere. However, the magnetic stirrer could cause a gas build up in the can and CO2 can be removed from the solution. This would cause the pH to increase because the depletion of H3O+ ions in the soda. This is similar to how hyperventilating in the body functions, making the solution more basic in order to make up for lost CO2.

6. The bicarbonate/carbonic acid buffer is also present in chickens. However, chickens also combine the carbonate in their blood with calcium ions to make calcium carbonate for their eggshells. Since chickens do not sweat, they pant in hot weather. What effect would this have on the pH of their blood and the strength of the eggshells they produce?

2 H2O + CO2 H2CO3 + H2O H3O+ + HCO3- CO3-2 + H3O+

Chickens require CO32- for their egg shells, and this causes a lot of H3O+ will form in their blood. Since chickens cannot sweat it out, they will pant a large amount in order to remove the excess CO2 in their bodies. When this happens, their body will begin to try and replenish the loss of carbon dioxide which will create a depletion of H3O+. This maintains the chicken's blood pH level at the ideal level.

Conclusion:

This lab allowed us to view how the addition and removal of products and reactants can cause visible shifts in equilibrium such as colour change. The body's chemical buffer system that maintains an ideal blood pH can be replicated through the experiment and demonstrates the effectiveness of the system.