

Experiment 3 - Equilibria

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Introduction:

Most chemical reactions will not go to completion. Instead, they reach a dynamic state called the *equilibrium*, where the rate of reaction of reactants forming products is equal to the rate of reaction of products forming reactants. The excess on one side of the reaction is consumed by the other. For example, when writing out an aqueous reaction, it's denoted by two arrows facing in opposite directions:

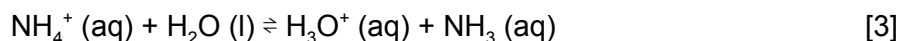


At this state, either the formation of product or reactant is favoured. Reactions usually favour the products, meaning that most of the reactants are converted into products. However, if the reactants are favoured the products will be converted to reactants. The concentrations of products and reactants can be used to determine an *equilibrium constant*, K , which expresses which side of the reaction will be favoured. A higher equilibrium constant expresses a product-favoured reaction while a lower constant expresses reactant-favoured. Using variables from equation [1], the equilibrium constant is:

$$K = \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad [2]$$

This experiment explores the outcome of imposing change on a dynamic state by altering the concentration of acids or bases. *Le Chatelier's Principle* states that the reaction would move in a direction that minimizes the effect on the system. Therefore, increasing the concentration of a reactant in equation [1] would cause the reaction to move in the direction of the products.

In aqueous reactions at equilibrium, an acid will produce and conjugate base while a base will produce a conjugate acid. A species will not react with its conjugate form. When working with acids and bases in an aqueous reaction that favours one side, the species, acid or base, is considered *strong* when it completely dissociates in water and *weak* when it does not. To determine whether the species is *strong* or *weak*, a similar equation to equation [2] is used to express the *acid/base dissociation constant*. For example, reacting an ammonium ion with water [3] will express the *acid dissociation constant* shown in equation [4].



$$K_a = \frac{[\text{H}_3\text{O}^+ (\text{aq})] [\text{NH}_3 (\text{aq})]}{[\text{NH}_4^+ (\text{aq})]} \quad [4]$$

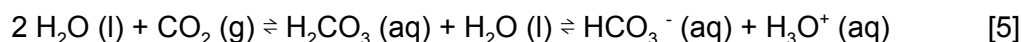
In equation [4], the concentrations are substituted as approximations of the *activity* of the substance. Since liquid water has an activity of 1, H_2O is not present in the equation since it will not affect the result. The same method would be used to find the *base dissociation constant*,

K_b .

This experiment also requires the use of indicators to determine the acidity of a solution. An indicator is a complex organic compound, usually a weak acid that will dissociate differently depending on the acidity of the solution. An indicator produces a colour dependant on the undissociated and dissociated species being used.

Buffers and the Human Body

In order to function properly, the human body must remain at a pH level in between 7.35 and 7.45. Levels much lower than 7.35 means the body falls into acidosis and levels much higher than 7.45 means the body goes into alkalosis. Both situations can lead to death. This experiment explores the carbonate/carbonic chemical buffer system in the blood of humans that works with the respiratory system in maintaining homeostasis. For this buffer system to take place, there must be a considerable amount of both the acid and its conjugate base. In blood, carbonic acid, H_2CO_3 , is a weak acid that is in equilibrium with the hydrogen carbonate (bicarbonate) ion, HCO_3^- . The bicarbonate ion in the blood buffer is significantly high with a ratio of approximately 20:1. This is because there are more metabolic acids produced then there are metabolic bases. Carbonic acid is not a stable acid and breaks down often producing water and carbon dioxide. The equilibrium for this buffer system is expressed as the following:



This high concentration of bicarbonate ions neutralize the metabolic acids making it feasible to have a smaller amount of carbonic acids in the buffer. Since carbon dioxide is being formed in the reaction, the respiratory system is applied in this buffer system and can act to neutralize the pH as well. However, if someone hyperventilates, meaning they are breathing too quickly and exhaling too much carbon dioxide, the rest of the system would have to accommodate for the lost CO_2 by depleting the amount H_3O^+ in the blood causing alkalosis. In equation [5], the equilibria would shift to the left. If someone were to hypoventilate, meaning they are breathing too slowly and not exhaling enough CO_2 , there would be an excess of H_3O^+ consuming CO_2 in the blood causing acidosis. In the equation [5], the equilibria would shift to the right. In part 3 of this experiment, a buffer solution will be prepared mimics a metabolic reaction.

Materials:

<i>Chemicals</i>	<i>Equipment</i>
0.1 M and 1 M HCl 0.1 M CuSO ₄ 0.1 M Na ₂ S 0.01M AgNO ₃ 0.1 M Na ₂ CO ₃ NaHCO ₃ 0.85% lactic acid NH ₄ Cl 0.1 M KI NH ₃ concentrated 6 M HNO ₃	test tubes transfer pipets Spatula stir plate pH probe magnet test tube racks 150, 250 mL beakers

Procedure:

Part 1 - Equilibrium Shift

1. Add approximately 1mL (~20 drops) of 1 M CuSO₄ to a test tube.
2. Add concentrated NH₃ by drops until change occurs.
3. Add 1 M of HCL by drops until change occurs.
4. Repeat steps 2-3.

Part 2 - Multiple Equilibria

1. Add 0.5mL of 1 M NaCO₃ in empty test tube.
2. Add 0.5mL of 0.01 M AgNO₃ to tube.
3. Add 6 M HNO₃ by drops until solution becomes clear.
4. Add 0.1 M HCl by drops until change occurs.
5. Add concentrated NH₃ until change occurs.
6. Repeat step 2 and step 5.
7. Add 0.1 M KI by drops until change occurs.
8. Add 0.1 M Na₂S until change occurs.

Part 3 - Buffer Solution and Blood pH

1. Place 100mL of distilled water in 150mL beaker with stir bar (magnet) then place on stir plate. Turn stir plate on low until a slight vortex occurs in water.
2. Insert pH probe in water keeping distance away from stir bar and wait a few minutes.
3. Add 2.50g of NaHCO₃ into beaker and wait until dissolved.
4. Add 30mL of 0.1 M HCl into beaker.
5. Add 10mL of 0.85% lactic acid into beaker.
6. Simulate increase respiration rate by increasing stir speed. (Caution that stir bar does not bounce and hit pH probe)
7. Add approximately 0.50g of NaHCO₃ to beaker and wait to dissolve.
8. Add another 0.50g of NaHCO₃ to beaker and wait to dissolve.
9. Add a pellet of CO₂ and dissolve.
10. Add approximately 0.40g of NH₄Cl to beaker.

Observations/Results:

Data Table 1. Equilibrium Shift

$[\text{Cu}(\text{H}_2\text{O})_4]^{2+}(\text{aq}) + 4\text{NH}_3(\text{aq}) \rightleftharpoons [\text{Cu}(\text{NH}_3)_4]^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	
Step(s)	Observation(s)
1-2	When CuSO ₄ was in the test tube (clear blue liquid with a strong odour), and drops of NH ₃ were added, the liquid turned to a layered pale blue solution. Light blue at the bottom and dark blue at the top.
3	When 1M of HCl was added to the solution the light blue liquid turned into a solution with three layers: dark blue, clear in the middle and then dark blue once again.
4	There was no change in colour when NH ₃ solution was added was added to the liquid already found in the test tube. But, when HCl was added to the test tube once again, it was noted that the solution became a paler blue with a more prominent ombre. The colour is more or less equal to the blue in Step 1.

Date Table 2. Multiple Equilibria

$2\text{AgNO}_3(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightleftharpoons \text{Ag}_2\text{CO}_3(\text{s}) + 2\text{NaNO}_3(\text{aq})$	
Steps(s)	Observation(s)
1-2	Na_2CO_3 is a clear liquid when it was added alone in the test tube. When AgNO_3 was added to the solution it turned to a white-cloudy solution.
3	2 drops of HNO_3 were added to the solution which made it become clear.
4	Afterwards, 6 HCl drops were added to the solution and it became cloudy once again.
5	6 drops of NH_3 drops were added to the solution also, except there was no change in colour. There was no change with the addition of the solution NH_3 since it isn't strong enough to react with any other solution.
6	HNO_3 and NH_3 were both added to the solution and there was no change.
7	13 drops of KI solution was added to the the test tube and it created a layer between the cloudy and clear part of the solution.
8	15 drops of Na_2S was also added to the solution, it created a reaction which caused the clear and cloudy layers in the test tube to turn a dark grey. The drops of Na_2S caused a reaction with the silver present in the solution. When the sulfide and silver reacted together, it caused an equilibrium shift causing AgI ions to dissolve but solid silver ions were still present.

Data table 3. Buffer Solution

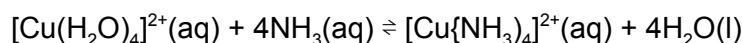
$2\text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g}) \rightleftharpoons \text{H}_2\text{CO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HCO}_3^-(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$	
Steps (s)	Observation(s)
1-2	The final measurement of the distilled waters pH with the stir stick in it was 4.68.
3	When 2.51g of NaHCO_3 (white crystallized powder) was added in the water, the pH changed to 8.20 after a few minutes.
4	The addition of 30 mL of HCl which had a concentration of 0.1M (clear liquid) caused the solution to decrease to a pH of 7.39.
5	After adding 10 mL of 0.85% lactic acid the pH was 7.25. When the lactic acid was being dissociated in the solution, it caused a formation of gas (bubbles) of CO_2 to form and rise to the top.
6	The increased speed of the stir stick caused the pH of the solution to increase to 7.31. Also, there were also bubbles forming.
7	As 0.50g of NaHCO_3 , a crystallized powder, was added the pH went up to 7.43. It started to dissociate almost immediately. Although, it did take some time for all of the NaHCO_3 to completely dissolve in the solution.
8	Adding another 0.50g of NaHCO_3 in the solution caused the pH to increase to 7.53. This time it took longer for the crystallized powder to dissociate completely.
9	When a pellet of CO_2 was added in the solution, the pH decreased to 7.34 finding a more neutral state. As soon it was added, it sank to the bottom of the beaker and spun around because of the stirring stick. There were cloudy bubbles forming that floated and bursted at the top of the beaker. There were also fumes coming from the solution. When the pellet became smaller, it started to float to the top, since it was losing mass and then it finally dissociated after 4 minutes.
10	The addition of NH_4Cl in the beaker caused the solution to only increase its pH by 0.2, so up to 7.36. The NH_4Cl is a white powder substance that's added in the solution

Discussion:

Part 1 - Equilibrium Shift

In this experiment, it was important to observe the changes in the system's equilibrium and on how it reacts with another solution being added to an already existing one that was found in the test tube.

During the part of the "Equilibrium Shift", the following equation was used in order to determine the systems chemical balance by adding $\text{NH}_3(\text{aq})$ to $\text{Cu}(\text{H}_2\text{O})_4$ creating $[\text{Cu}(\text{NH}_3)_4]$.



When the new solution was added, the system favoured the right side of the equation, in other words the products. The $[\text{Cu}(\text{NH}_3)_4](\text{aq})$ cause the light blue colour to form in the test tube.

Next, HCl was added to the already existing solution, causing a tri-layer of colors: dark blue (on the top), transparent-blue (in the middle) and dark blue (on the bottom) of the test tube. The reason for this is because, since HCl is a strong acid it will ionize completely. With that being said, it will form H^+ and Cl^- ions that will be found at the bottom of the test tube and cause a dark blue appearance. Le Châtelier's principle explains that a system will favour a reaction that will cause the most elements to form. So, the system will favour the left side of the equation (the reactants) and then forming more $\text{Cu}(\text{NH}_3)_4(\text{aq})$. That element will cause the transparent-blue liquid in the tri-layer of the test tube.

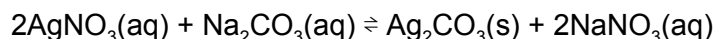
When more NH_3 was to be added in the solution, it created a larger volume of liquid that had a transparent-blue colour. The reason is that the system will favour the products being produced, so creating more $\text{Cu}(\text{NH}_3)_4(\text{aq})$ that has the physical effect of a light blue colour.

Adding more HCl to the solution, clearly demonstrated a thicker layer of a darker blue on the outer layers of the test tube, since the HCl will ionize completely. Which means that more ions of H^+ and Cl^- will be found in the extremities of the test tube. The colour was very similar to the colours found in step three when the first few drops of HCl were added.

Finally, the more NH_3 or HCl you added throughout the experiment, would hardly cause a reaction since the system has already found its equilibrium.

Part 2 - Multiple Equilibria

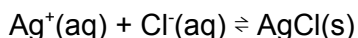
The second part of the laboratory included an experiment that has multiple equilibria, it first started off with the chemical equation:



The first step was to add Na_2CO_3 in the test tube followed by AgNO_3 . Together they reacted and formed a precipitate of Ag_2CO_3 . By adding a sufficient amount of both solutions together, the system favoured the products which caused the formation of a solid.

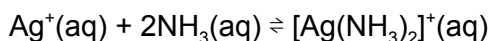
Afterwards, NO_3^- was added to the solution. The reaction caused the precipitate to dissolve and the solution to become clear once again. The excess of NO_3^- created a new equilibrium by favouring the left side (reactants). With that consequent, it used the Ag_2CO_3 and NaNO_3 to create more reactants, in other words removing the precipitate followed by the cloudiness of the solution.

The next step was to add HCl to the already existing solution in the test tube, the following chemical equation demonstrates and explains the reaction that happened:



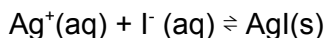
Since the solution returned into an opaque cloudy-white colour, there was a formation of another precipitate. Although, the precipitate was formed by Ag^+ ions (from the compound AgNO_3) and Cl^- ions reacting together. So, the system favoured the products of this reaction.

When NH_3 was combined with the already existing solution, it created a reaction that dissolved the $\text{AgCl}(\text{s})$ causing the Ag^+ ions and NH_3 ions react together and form $\text{Ag}(\text{NH}_3)_2^+$, as is shown in the chemical equation:



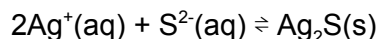
So, the reaction favoured the right side of the equation, forming more products than reactants. Fortunately, the solution is reversible. If there was a lack of Ag^+ in the solution, the system would shift towards the left causing it to favour the reactants and producing more Ag^+ ions.

KI was the next solution to be included in the solution. Since it is an aqueous solution, it will dissociate into K^+ and I^- ions in the solution. With the new ions present in the solution — especially I^- ions will bond with the Ag^+ ions in the solution, forming another precipitate, described in the reaction:



So the system favoured the right side of the equation once again and forming the precipitate AgI . The silver ions didn't react with the Cl^- ions, since it's more unstable and not as strong as I^- , and it will take less energy to form with it compared to chloride.

The last step of this experiment was to add Na_2S until a change was observed. The chemical equation for this step is:



Therefore, when it was added, the system created a new equilibrium that caused the dissociation of the solution Na_2S ; that created Na^{+2} ions and S^{-2} ions. The S^{-2} ions reacted well with the Ag^+ , since there was an excess already found in the remaining solution. The system attempted to decrease those numbers by forming the Ag_2S precipitate and favouring the products of the reaction.

Part 3 - Buffer Solution

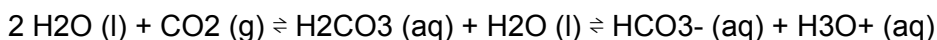
The final part of laboratory included an experiment that adding different solutions to distilled water in order to determine and understand a buffer solution.

The first step was to add distilled water in a beaker and measure the pH which was 4.68. Although, the pH was expected to be around 7.0, since it is considered a neutral element if it is coming from a tap. The pH did indicate that the water that was in the beaker was more acidic than normal. An error would be not cleaning the beaker before use, since there could have been possible solutions still in it causing the change in pH.

Next was to add a white crystalline powder of NaHCO_3 in the beaker with the distilled water and it raised the pH to 8.20. When the system was introduced to a new solution that could dissociate in which it could give or receive an H^+ (either making it more acidic or basic), the system attempted to become more neutral. So, it dissolved the NaHCO_3 into Na^+ and H_2CO_3 , making it more basic. Although, the elevated pH would be dangerous for a human body, so an attempt to balance the pH once again is to add HCl .

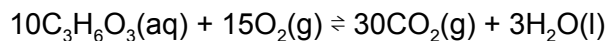
To continue, HCl (a strong acid), was used since it can ionize completely, causing H^+ to bond with CO_3 and having an extra H^+ from the H_2CO_3 to come together and lower the OH^- ions that cause a more basic solution. With that being said, the pH then lowered to 7.39, which is a comfortable pH level for a human body that includes the circulating blood.

The chemical equation for this reaction is:

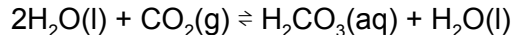


Aqueous and gas species are present in this reaction. Although, aqueous species are more present than gases, since there were only a small amount of gases formed with the addition of HCl . To add, the mol ratio of aqueous solutions is much greater than the ratio of the gas moles used in this experiment.

The experiment required an addition of 10 mL of 0.85% of lactic acid which then decreased the pH to 7.25. The condition that this reaction stimulates is the body producing lactic acid when there's a lack of O_2 being distributed to the muscles. A large amount of $C_3H_6O_3$ would cause muscle cramps and pains. The following chemical equation demonstrates this reaction:

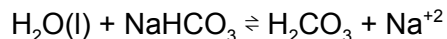


As the experiment progresses, the next step is to increase the speed of the stir stick the beaker. By doing so, the pH level increases to 7.31. The information that was given was that the body would use the respiratory system to try and increase pH levels by increasing the respiration rate in order to get rid of the CO_2 present in the body. Therefore, the obtained results were expected. The chemical equation to demonstrate this chemical reaction is:



So the increase in CO_2 will cause the system to favour the products and then creating more H_2CO_3 that is more basic, finally lowering the pH level.

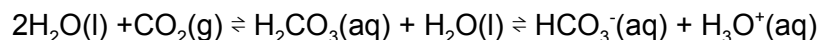
Moreover, 0.50g of $NaHCO_3$ was also added in the solution. The reaction caused the pH to increase to 7.43. In the equation:



The system realized that there was a change in the H^+ ions, so in order to minimize the effects, it favoured the products. With that being said, it produced more H_2CO_3 creating a more basic pH afterwards. Not long after, another 0.50g of $NaHCO_3$ is added to the solution, which caused the pH levels to increase to 7.53. Using the same chemical equation as the one above, the system once again recognized a change in its equilibrium causing it to favour the products and raising its pH.

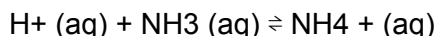
At this point, if the pH in the body was this elevated, it would create another form of reactions to lower it as explained in the next step.

In this part of the experiment was to stimulate a large amount of CO_2 being kept in the body in order to lower the pH. To create this effect, a pellet of CO_2 was placed in the beaker with the solution causing the pH to decrease to 7.34. It can be determined by the chemical equation:



When the pH in the body is too elevated, the body will favour the reactants on the far left side of the equation that will cause it to produce more CO_2 , thus lowering the pH levels back to a neutral state.

Finally, when NH_4Cl was added to the solution, the pH level did not differ too much from the last measurement, since it only went up to 7.36. The chemical equation used to describe this reaction is:



If it was an extreme case, where the solution's pH was too elevated, the injection would favour the reactants, since it has the strong acid which would cause a decrease in the pH. To further explain, NH_4^+ will dissociate into H^+ and its conjugate base which is NH_3 . Since the system was close to its equilibrium, it favoured the reactants in order to use the NH_3 base to raise the pH by 0.2 which created the last measurement.

Conclusion:

In part one of the experiment, the reaction went from favouring one side of the reaction to the other depending on what was added until it reached its equilibrium and would not change anymore.

In part two of the experiment, the silver ion was able to establish multiple equilibria in the solution as different chemicals were associating and dissociating.

In part three of the experiment, the buffer system was set up and the pH began in the neutral threshold at 7.39. As the lactic acid was added and pH decreased to 7.25 simulating acidosis, the stirring rod mimicking the increased respiratory rate was able to neutralize it back to 7.31. The addition of NaHCO_3 mimicking alkalosis brought the pH up to 7.53 while the addition of the CO_2 pellet simulated slowing down the respiratory rate which again neutralized the solution back to 7.34.

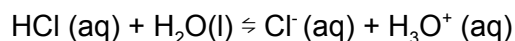
Questions:

1. A. The ammonium ion is used as the acid source instead of HCl to simulate a different metabolic acid being produced by a muscle or an organ. The acid would transfer to the blood and the buffer would neutralize it. Also, when NH_4Cl dissolves in the water, the Cl^- ion will not affect the pH in a remarkable matter. In a buffer solution, it is important to not use a strong acid such as HCl since it will ionize completely. So, it was important to use NH_4Cl , since NH_4^+ has a conjugate base, which is NH_3 . Finally, it would make the reaction reversible if the system's equilibrium wasn't balance correctly.

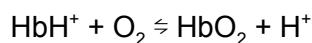
B. NH_4I wasn't used in the experiment since, it forms a strong acid (HI) and a strong base (KOH), to add I^- ions doesn't have any acidic or basic properties. So that means, that this solution wouldn't cause a change in the buffer solution⁵. Next, NH_4NO_3 wasn't used in the

buffer solution experiment since it would cause a solution that involves a conjugate base (NH₃) and a strong acid (HNO₃), which would cause a strong acidic reaction. Thus, it was not needed in this experiment since the purpose was to lower the pH in the buffer solution and not increase it.

2. The addition of acidic substances lowered the pH value of the solution. Since a basic solution has a higher concentration of hydroxyl ions compared to hydronium ions, an increase in the concentration of hydronium ions in that solution would make it less basic. Thus, the pH value decreases.



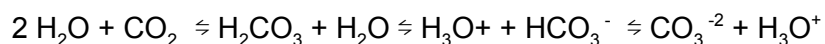
3. There would be an increase in the H⁺ ions causing a shift to the left in the equilibrium by Le Chatelier's Principle. This would reduce the transport of oxygen throughout the body causing fatigue and a headache.



4. The addition of HCl in the solution on the right that contained blood shows the deoxygenation of the blood as well as an increase in the acidity of the solution. Similarly to question 3, the equilibrium would be forced to the left to consume the additional H⁺ ions preventing the oxygen from binding to the hemoglobin.

5. The pH of the flat soda would be higher than that of the fresh soda. It would be the same process as the respiratory system exhaling CO₂ to release some of the acidity.

6. As the chickens pant to keep cool, they exhale CO₂. This causes the equilibrium to shift to the left to as per Le Chatelier's Principle. The shift to the left leads to a decrease in the concentration of carbonate dissolved in the blood. The decrease in carbonate means there is less of it for the calcium ions to bind to in the process of making their shells. This results in much thinner and weaker shells.



References:

Clark, Roy W., et al. "From Titration Data to Buffer Capacities: A Computer Experiment for the Chemistry Lab or Lecture." *Réseau De Bibliothèques - Library Network*, Middle Tennessee State University, Murfreesboro, pubs.acs.org.proxy.bib.uottawa.ca/doi/pdf/10.1021/ed072p746.

1: Unknown. "Le Châtelier's Principle: Buffers."
lincolnparkhs.org/ourpages/auto/2010/11/21/60435055/Chapter%2020-%20Experiment%20-%20LeChatelier_s%20Principle%20-%20Buffers.pdf.

2: Unknown. "Hydrolysis and Acid-Base Properties of Salts." *15.10*, The McGraw-Hill Companies, 2001, www.mhhe.com/physsci/chemistry/chang7/ssg/chap15_10sg.html.

Date 2017-10-17
 Experiment #3 - Equilibria *Anthony G.*

Part 1 - Equilibrium Shift

1. Add approx. 1ml (~20 drops) of 1ml/L CuSO_4 to test tube
2. Add concentrated NH_3 by drops.
3. Add 1 ml/L HCl drops.
4. Repeat steps 2-3.

Part 2 - Multiple Equilibria

1. 0.5ml of 1ml/L Na_2CO_3 in test tube.
2. Add 0.5ml of 0.1ml/L AgNO_3 to same tube.
3. Purposely add 6 ml/L HNO_3 until change.
4. When solution clear, add 0.1ml/L HCl drops until change.
5. By drops, add concentrated NH_3 until change.
6. Repeat step 3 and then step 5.
7. Add 0.1ml/L KI drops until change.
8. Add 0.1ml/L Na_2S drops until change.

Part 3 - Buffer Solution and Buffer pH

1. 100ml of distilled water in 150ml beaker. Place on stir plate. Insert stir bar and turn on slow.

Date

2. Insert pH probe in water away from stir bar. Wait a few minutes.
3. Add 2.50g of NaHCO_3 into beaker. Wait until dissolved.
4. Add 30ml of 0.1ml/L HCl to beaker.
5. Add 10ml of 0.85% lactic acid.
6. Simulate increased respiration by increasing stir speed.
7. Add approx. 0.50g of NaHCO_3 .
8. Add another 0.50g of NaHCO_3 .
9. Add a pellet of CO_2 .
10. Add approx. 0.40g of NH_4Cl .

Clean up properly

Observations

Part 1

- 1 - pale blue
- 2 - goes to darker blue (inky)
- 3 - No change (crystal blue)
- 4 - No change
- 5 - Pale, ~~dark~~ amber ^{yellow} at bottom

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Date

Part 2

- 1-2 - Clear to cloudy
- 2-3 - 2 drops → Clear
- 3-4 - 6 drops → back to cloudy (thick)
- 4-5 - 6 drops → back to clear
- 6 - no change
- 7 - Cloudy layer on top
- 8 - Cloudy layer turns dark grey (blackish)

Part 3

pH water: 4.66
 2.51g NaHCO_3 ↑

pH $\text{H}_2\text{O} + \text{Na}_2\text{CO}_3$: 8.283

pH $\text{H}_2\text{O} + \text{Na}_2\text{CO}_3 + 0.1\text{ml/L HCl}$: 7.39

" + 0.85% lactic acid: 7.25

increased speed: 7.31

+ 0.50g NaHCO_3 : 7.43

+ 0.50g NaHCO_3 : 7.53

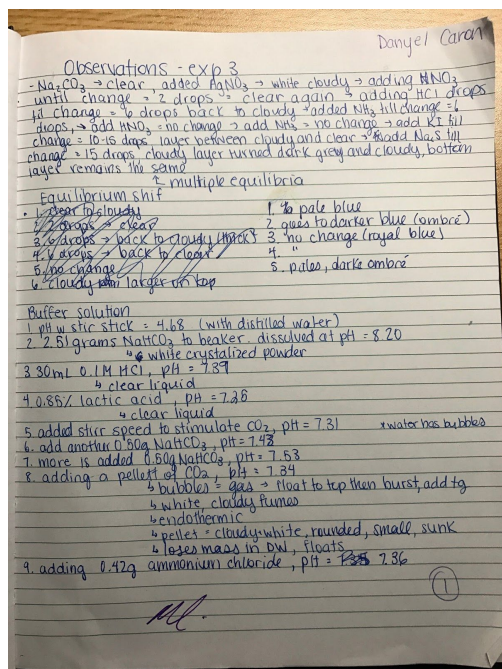
+ CO_2 : 7.34

+ 0.40g NH_4Cl : 7.16

Pond Big cloudy ^{grey} at bottom
 Bubbles ^{to} ~~from~~ coming out

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Danyel Caron



COURSE: CHM 1311 Sec. D7

TA Name: Michael Coates

YOUR NAME (PRINT): Anthony Rankin

SIGNATURE: _____ A.R _____

CONFIDENTIAL PEER EVALUATION FORM FOR EXPERIMENT __3__

In assessing the work of your fellow team members, consider the following aspects:

- Quality of work with others
- Contribution to the work as a whole when asked to correct
- Ability to get along
- Improvements

Team member name	Comments	Grade
Danyel Caron		A

**A – Excellent (5) B: Great (4) C: Good (3) D: Fair(2)
F: Poor (1)**

Note: Do not evaluate yourself on this form