

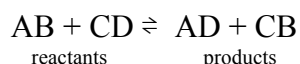
## Laboratory 3: Equilibria

CHM 1311

October 17<sup>th</sup> 2017

## Introduction

Equilibrium is the dynamic state of a reaction when the “rate of reaction of reactants to form products is equalled by the rate of reaction of products to form reactants”. In this stage, products can be easily “converted back into reactants” (Venkateswaran, 2000, p. 15). A reaction in a state of equilibrium can be denoted with a double arrow within the reaction equation:



In this investigation, the behaviour of reactions in equilibrium are studied when different dynamic chemical reactions undergo a manipulation. Manipulations to dynamic chemical reactions can be further understood using Le Chatelier’s principle.

According to Le Chatelier’s principle, when a system in equilibrium is disturbed, the equilibrium position will shift, allowing either the forward or reverse reaction to be favoured in order to compensate for the change (Silberberg et. al, 2016, p. G-10). One can determine the relative equilibrium position by calculating the equilibrium constant K and expressed with the equation,

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

A reaction that favours the forward reaction forms more products, whereas a reaction favouring the reverse reaction forms more reactants. The factors that disturb equilibrium include a change in concentration, a change in temperature or energy input, a change in pressure, and a change pH. This principle is applicable to this investigation as solutions are manipulated by these factors. Le Chatelier’s principle assists in predicting what products or reactants are formed

following the addition of reactants, change in concentration of a given chemical, change of pH and inputting kinetic energy into the system by increasing the stir rate.

In this experiment, the reaction and formation of ions are of the greatest focus in which Le Chatelier's principle can be applied. According to Silberberg et al., "the common-ion effect occurs when a given ion is added to an equilibrium mixture that already contains this ion," and the equilibrium position shifts in the direction that reduces the ion formation (Silberberg et al., 2017, p. 710). In this investigation, given ions are repeatedly dissociated into solutions. Observable changes indicate which ions form and cause the change.

In addition to the common ions, temperature has a role on the shift of the equilibrium position. For an endothermic reaction where more heat is absorbed by the reactants, the system will compensate for this change by shifting the equilibrium position in the forward direction and form more products. In contrast, an exothermic reaction would favour the reverse reaction in which less products are formed to compensate for the change. In this experiment, an exothermic reaction is observed. Knowing this, an observer can determine whether the reaction formed more products or the reaction reversed to favour more reactants.

A complex ion is one that "consists of a central metal ion covalently bonded to two or more anions or molecules, called ligands" (Silberberg et al., 2016, p. 745). A complex ion is formed when a salt dissolves water. In this experiment,  $[\text{Cu}(\text{H}_2\text{O})_4]^{2+}$ ,  $[\text{Cu}(\text{NH}_3)_4]^{2+}$ , and  $[\text{Ag}(\text{NH}_3)_2]^+$  are complex ions being studied. When analyzing observations from the lab, it is important to understand that these ions may be formed by reacting chemicals that are diluted with water. One should also make note that in this experiment, the colour change of a solution is

due to the presence of ion complexes rather than simple ions because the solid metals provide a colour whereas gases are colourless.

According to the Arrhenius theory described in Silberberg et al.'s work, "an acid is a substance that has H in its formula and dissociates in water to yield  $H^+$  or  $H_3O^+$ , and a base is a substance that has OH in its formula and produces  $OH^-$  in water" (Silberberg et al., 2016, p. G-2). Furthermore, the Brønsted-Lowry theory describes an acid as a species that "donate[s]" a proton, whereas a base is a species that "accept[s]" a proton (p. G-3). In this study, the Arrhenius theory was more applicable as the dissociation of  $H^+$  or  $H_3O^+$  and  $OH^-$  ions was more obvious in distinguishing which substances had acidic or basic properties.

Strong acids and bases "dissociate completely into ions" when dissolved in water, whereas weak acids and bases only "dissociate slightly" (Silberberg et al., 2016, p. 668-669). To determine the strength of an acid or base, one would calculate an acid dissociation constant. The strength of an acid can be determined using the acid dissociation constant,  $K_a$ . This constant can be calculated using the equation,

$$K_a = \frac{[A^-(aq)][H_3O^+(aq)]}{[HA(aq)]}$$

when  $[A^-(aq)]$  is the concentration of the acid's conjugate base,  $[H_3O^+(aq)]$  is the concentration of dissociated hydronium ions, and  $[HA(aq)]$  is the concentration of the acid in solution.

Similarly, the strength of a base can be determined using the base dissociation constant,  $K_b$ . This constant can be determined using the equation,

$$K_b = \frac{[B^+(aq)][OH^-(aq)]}{[BOH(aq)]}$$

when  $[B^+ (aq)]$  is the concentration of the base's conjugate acid,  $[OH^- (aq)]$  is the concentration of dissociated hydroxide ions, and  $[BOH (aq)]$  is the concentration of the base in solution.

In addition to acidic and basic species, there are amphoteric species. An amphoteric species "behav[es] both as an acid and a base" (Venkateswaran, 2000, p.17).  $H_2O$  is the most common amphoteric species and abundant in this experiment. One other amphoteric species in this experiment is the  $HCO_3^-$  ion.

The pH of a solution determines whether it is acidic or basic. In solutions with higher concentrations of  $H^+$  ions or  $H_3O^+$  ions, the pH is lower and considered acidic. In solutions with higher concentrations of  $OH^-$  ions, the pH is higher and considered basic. When measuring pH, one would refer to the logarithmic pH scale. A pH value of less than 7 would be considered acidic, whereas a pH value of more than 7 would be considered basic. This scale was referenced in the experiment to observe the changes in pH of solutions when a manipulation was conducted.

Weak acids and weak bases, along with their conjugate bases and conjugate acids respectively are essential in terms of buffer systems. Their characteristic ability to only slightly dissociate in water permits the coexistence of an acid and base in solution, at relatively similar amounts. Their coexistence and the possibility of their dissociated ions recombining to the original complex allows the system to resist changes in pH.

The human blood has its own carbonic acid buffer system in order to maintain a healthy pH. The normal blood pH of the human body is tightly regulated between 7.35 and 7.45. In this experiment, the investigators created a buffer system and conduct manipulations to simulate the blood's buffer system. This simulation is done using the amphoteric  $HCO_3^-$  ion, dissociated in solution from sodium bicarbonate.

This equilibria investigation relied heavily on the changes observed following the manipulation of a given solution. Indicators such as changes in pH, changes in appearance, changes in odour, and changes in temperature inform the investigator of a reaction taking place, and help specify whether it reacts in the forward direction or the reverse. For example, if a chemical is added to a solution and bubbles appear in the beaker, the investigator can conclude that a gas has formed as the product of a chemical reaction.

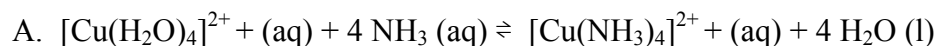
The purpose of this lab is to investigate reactions in solutions in a state of equilibrium. First, the investigators observed the effect of adding and removing a reagent on equilibrium. Next, multiple equilibria is observed with the silver ion. Finally, the investigators prepared a buffer system to stimulate a metabolic reactions within the human body.

A limitation of this experiment is that equilibrium reactions in reality are not as controlled as the reactions seen in this laboratory. In reality, it may not be as obvious as to whether a forward or reverse reaction is being completed. For example, the third part of the experiment is limited because the buffer solution prepared in the laboratory does not exactly mimic the carbonic acid buffer system in the human blood. Often, chemicals were added to the buffer solution in the lab and altered the pH so that it increased or decreased out of a healthy range for the human body. In the true human carbonic acid buffer system, the blood would be capable of compensating for the pH changes so that it would return to a healthy pH. Also, the human body has many other systems that act to regulate the blood pH, which was not accounted for in this experiment.

### Table 1: Raw Data, Observations and Discussion<sup>1</sup>

<sup>1</sup> the letters labeled *A, B, C ...* are to organize the chemical equations. The bullets labeled *1.1, 1.2, 1.3 ...* are used to organize raw data. The letters in brackets (*a*), (*b*), (*c*) ... correspond with the observation questions asked in the experiment pdf.

#### **Part 1: Equilibrium Shift**



- CuSO<sub>4</sub> - pale blue translucent liquid
- NH<sub>3</sub> - clear liquid with strong sour odour
- HCl - clear solution

1.1 Adding 20 drops of CuSO<sub>4</sub> to an empty test tube shows a pale, translucent blue solution. It is odourless and has low viscosity.

1.2 Adding 1 drop of NH<sub>3</sub> to 1.1 causes an instant change in colour from a pale blue to an opaque royal blue colour, still remaining low viscosity liquid. There was a strong sour odour.

1.3 Adding 29 drops of HCl to the products in 1.2 causes the solution to turn a light blue translucent colour. There is no odour.

1.4 Adding 3 drops of NH<sub>3</sub> to the products in 1.3 causes solution to turn royal blue colour with a strong sour smell. This translucent liquid was more dense on the bottom of the test tube.

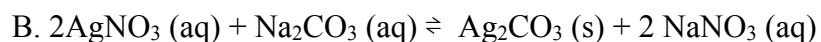
1.5 Adding 38 drops of HCl to the products in 1.4 causes the solution to change colour. There was clear liquid at the bottom of the test tube initially with the separation of royal blue. Overtime, the colours merged into a pale blue translucent liquid with no odour, similar to the appearance of the products in 1.3.

- (a) The CuSO<sub>4</sub> 0.1M solution is a translucent, pale blue colour. The Cu<sup>2+</sup> ions are causing this colour.
- (b) Adding only 1 drop of the concentrated NH<sub>3</sub> clear solution to the 20 drops of pale blue translucent CuSO<sub>4</sub> results in a royal blue opaque solution. This change observed can be explained using the chemical reaction (A) provided above. It is seen that the Cu<sup>2+</sup> ions react with NH<sub>3</sub> to make water and [Cu(NH<sub>3</sub>)<sub>4</sub>]<sup>2+</sup>. Thus, the [Cu(NH<sub>3</sub>)<sub>4</sub>]<sup>2+</sup> ions cause the royal blue colour in the

solution.

- (c) Adding 29 drops of 0.1 M HCl to the CuSO<sub>4</sub> and NH<sub>3</sub> solution changes it to a translucent, light blue coloured solution. This change happens due to the added HCl reacting with unreacted NH<sub>3</sub> particles in the solution. Cl<sup>-</sup> and NH<sub>4</sub><sup>+</sup> ions are produced from that reaction, which are both translucent when dissolved in solution. The change also happened due to equilibrium system shifting left towards the 'reactants' side to produce more NH<sub>3</sub> as a result of adding HCl. When HCl was added to the system, it reacted with NH<sub>3</sub>, thus decreasing the concentration of NH<sub>3</sub> on the reactants side of the chemical reaction. This shifted equilibrium left and not only produced more NH<sub>3</sub>, but more [Cu(H<sub>2</sub>O)<sub>4</sub>]<sup>2+</sup> ions as well. This means that the concentration of pale blue [Cu(H<sub>2</sub>O)<sub>4</sub>]<sup>2+</sup> ions increases while the concentration of royal blue [Cu(NH<sub>3</sub>)<sub>4</sub>]<sup>2+</sup> ions decreased. That explains why the colour changed from opaque royal blue to translucent pale blue when HCl was added to the test tube. This pale blue colour was seen before in the original solution of CuSO<sub>4</sub>. This can be due to the equilibrium shifting left, which makes the overall light blue colour because the solutions consist more of pale blue [Cu(H<sub>2</sub>O)<sub>4</sub>]<sup>2+</sup> as well as clear NH<sub>3</sub>.
1. In step 4, when 3 drops of NH<sub>3</sub> solution and then 38 drops of HCl solution are added to the test tube again, the same changes are observed. It was expected that the changes would be the same because the addition of either of the two solutions will simply cause equilibrium to shift in a certain direction, either towards the products or towards the reactants. When additional NH<sub>3</sub> is added to the system, the equilibrium shifts right towards the products in order to offset the increase in NH<sub>3</sub> concentration, This causes the system to produce more [Cu(NH<sub>3</sub>)<sub>4</sub>]<sup>2+</sup> which causes the solution to appear a royal blue. Consequently, when additional HCl is added, it reacts with the NH<sub>3</sub> and shifts the equilibrium left towards the reactants to offset the decrease concentration of NH<sub>3</sub>. This causes the system to produce more [Cu(H<sub>2</sub>O)<sub>4</sub>]<sup>2+</sup> ions which is why the solution turns pale blue.

## Part 2: Multiple Equilibria



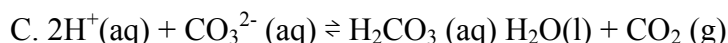
- Na<sub>2</sub>CO<sub>3</sub> - transparent liquid, colourless with low viscosity and no odour.
- AgNO<sub>3</sub> - appears a clear colourless solution

2.1 The Na<sub>2</sub>CO<sub>3</sub> solution is a clear, odourless liquid. The Na<sup>+</sup> and CO<sub>3</sub><sup>2-</sup> ions in the Na<sub>2</sub>CO<sub>3</sub> molecules appear colourless in solutions, which in turn, makes the solution itself colourless.

2.2 Adding 10 drops of clear 0.1M AgNO<sub>3</sub> solution to the 10 drops of 0.1MNa<sub>2</sub>CO<sub>3</sub> solution results in a murky, opaque, beige liquid solution with some precipitate.

(d) The  $\text{Na}_2\text{CO}_3$  solution is a clear, odourless liquid. The  $\text{Na}^+$  and  $\text{CO}_3^{2-}$  ions in the  $\text{Na}_2\text{CO}_3$  molecules appear colourless in solutions, which in turn, makes the solution itself colourless.

(e) Adding 10 drops of clear 0.1M  $\text{AgNO}_3$  solution to the 10 drops of 0.1M  $\text{Na}_2\text{CO}_3$  solution results in a murky, opaque, beige liquid solution with some precipitate. Looking at the chemical reaction (B) above, it is shown that the products of  $\text{AgNO}_3$  and  $\text{Na}_2\text{CO}_3$  are solid  $\text{Ag}_2\text{CO}_3$  in a solution of  $\text{NaNO}_3$ . This explains why the products of this reaction were murky opaque beige with some precipitate, as the  $\text{AgNO}_3$  caused the opacity of the solution due to it being a solid product. Therefore, the  $\text{Na}^+$  and  $\text{NO}_3^-$  ions cause this colour.



- $\text{HNO}_3$  - transparent, colourless liquid with slight odour, low viscosity.

2.3 Adding 2 drops of 6 M  $\text{HNO}_3$  to  $\text{Na}_2\text{CO}_3$ , and  $\text{AgNO}_3$  results in a clear solution.

(f) Adding 2 drops of 6 M  $\text{HNO}_3$  to  $\text{Na}_2\text{CO}_3$  and  $\text{AgNO}_3$  results in a clear solution. This change is reversible because it is an equilibrium shift. For example, the reaction can reverse if the concentration of  $\text{Na}_2\text{CO}_3$  is increased in the system, as this would cause the equilibrium to shift right to produce more  $\text{AgNO}_3$ , forming the precipitate and murky beige colour again. The observations made can be explained with Le Chatelier's Principle. When  $\text{HNO}_3$  was added to the system, it dissolved into its ions,  $\text{H}^+$  and  $\text{NO}_3^-$ , and since an equilibrium system has both reactants and products in it,  $\text{Na}^+$  and  $\text{CO}_3^{2-}$  ions from  $\text{Na}_2\text{CO}_3$  are still present. Thus, the  $\text{H}^+$  ions react with the  $\text{CO}_3^{2-}$  ions in the system which results in the  $\text{Na}_2\text{CO}_3$  concentration in the system to decrease. This change leads to the equilibrium shown in equation B. to shift left towards the reactants in order to produce more  $\text{Na}_2\text{CO}_3$  and at the same time produce more  $\text{AgNO}_3$ . Those two aqueous substances produce colourless solutions.

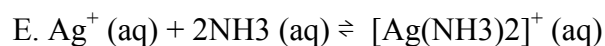


- $\text{HCl}$  - clear solution

2.4 Adding 2 drops of 0.1 M  $\text{HCl}$  to  $\text{HNO}_3$ ,  $\text{Na}_2\text{CO}_3$ , and  $\text{AgNO}_3$  results in a cloudy, translucent, white liquid.

(g) Adding 2 drops of clear 0.1 M  $\text{HCl}$  to  $\text{HNO}_3$ ,  $\text{Na}_2\text{CO}_3$ , and  $\text{AgNO}_3$  results in a cloudy, murky, white liquid. Once the  $\text{HCl}$  was added to the system it dissociated into  $\text{H}^+$  and  $\text{Cl}^-$ . The  $\text{Cl}^-$  ions reacted with the  $\text{Ag}^+$  ions left in the system to produce  $\text{AgCl}$  solid. This is shown as the murky

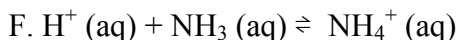
mixture is precipitate. Also, the equilibrium from the equation D. shifts right to produce more AgCl precipitate since the concentration of HCl was being increases.



- NH<sub>3</sub> - clear solution with very strong odour

2.5 Adding 3 drops of concentrated NH<sub>3</sub> to HNO<sub>3</sub>, Na<sub>2</sub>CO<sub>3</sub>, and AgNO<sub>3</sub> results in the solution separating into a clear liquid on top and a cloudy, translucent, white fluid substance on the bottom. Over time, the white substance dissipates into the clear solution. The change provided a slightly sour odour.

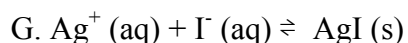
(h) Adding 3 drops of concentrated NH<sub>3</sub> to HNO<sub>3</sub>, Na<sub>2</sub>CO<sub>3</sub>, and AgNO<sub>3</sub> results in the solution separating into a clear liquid on top and a cloudy, translucent, white fluid substance on the bottom. Over time, the white substance dissipates into the clear solution, making the overall product a clear solution. The change provided a slightly sour odour. Referring to equation E. The NH<sub>3</sub> molecules react with the Ag<sup>+</sup> ions present in the system to produce [Ag(NH<sub>3</sub>)<sub>2</sub>]<sup>+</sup> ions. Increasing the concentration of NH<sub>3</sub> made the equilibrium shift right to produce more [Ag(NH<sub>3</sub>)<sub>2</sub>]<sup>+</sup> ions. The NH<sub>3</sub> reacted with the H<sup>+</sup> ions as well to form NH<sub>4</sub><sup>+</sup> ions. When the NH<sub>3</sub> is added to the system and it reacts with the Ag<sup>+</sup> ions present, the concentration of AgNO<sub>3</sub> from equation B. In the system decreases. To supplement the change, Le Chatelier's Principle explains that the equilibrium system has to shift towards the side of the reaction to produce more of the lacking concentration. Thus, the equilibrium in equation B. Shifts left towards the reactant to produce more AgNO<sub>3</sub> which is a transparent clear solution. This explains the observable changes in the experiment after adding NH<sub>3</sub>.



2.6 Adding an additional 5 drops of 6M HNO<sub>3</sub> to NH<sub>3</sub>, HNO<sub>3</sub>, Na<sub>2</sub>CO<sub>3</sub>, and AgNO<sub>3</sub> initially resulted in a separation with clear liquid on top, and a translucent, but cloudy white liquid on the bottom. Over time. The white substance dissipates into the clear solution. An exothermic temperature change was observed and a component of the solution was vapourized.

2.7 Adding an additional 20 drops of concentrated NH<sub>3</sub> to the test tube presented no visible change, however it did produce a sour, pungent odour.

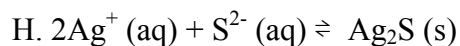
2. Repeating steps 7 and 9 creates the same observable results of producing clear solutions. Repeating step 7 of adding more clear HNO<sub>3</sub> solution to the system, the equilibrium from equation B. will shift left again. This is due to the H<sup>+</sup> ions from the HNO<sub>3</sub> reacting with the CO<sub>3</sub><sup>2-</sup> ions present in the system, which decreases the concentration of Na<sub>2</sub>CO<sub>3</sub> on the left side of equation B. In order for the system to even out the change, the equilibrium will shift towards the reactants on the left to produce more clear solutions of Na<sub>2</sub>CO<sub>3</sub> and AgNO<sub>3</sub>. When repeating step 9 of adding more clear NH<sub>3</sub> solution to the system, the equilibrium of equation B. will again shift left. This is due to the NH<sub>3</sub> reacting with the Ag<sup>+</sup> ions in the system, thus decreasing the concentration of AgNO<sub>3</sub>. To even out this change, the equilibrium will shift left to produce more clear AgNO<sub>3</sub> and Na<sub>2</sub>CO<sub>3</sub> again.



- KI - clear solution

2.8 Adding 3 drops of 0.1 M KI to NH<sub>3</sub>, HNO<sub>3</sub>, Na<sub>2</sub>CO<sub>3</sub>, and AgNO<sub>3</sub> resulted in a separation with a white, cloudy solution on top, and a translucent solution on the bottom of the test tube.

(i) Adding 3 drops of aqueous 0.1 M KI to the system results in a cloudy beige liquid that is more opaque on the top half of the solution in the test tube. When KI is added to the system, it dissociates into K<sup>+</sup> and I<sup>-</sup>; the I<sup>-</sup> ions react with Ag<sup>+</sup> present in the system from equation B. to produce AgI solid. This means that the observations of a cloudy mixture sitting on the top half of the liquid is actually AgI precipitate. As KI concentration increases, the equilibrium in equation G. shifts right to produce more AgI precipitate.



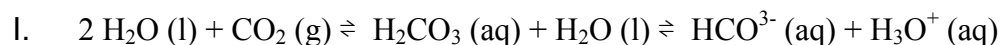
- Na<sub>2</sub>S - clear solution

2.9 Adding 3 drops of 0.1 M Na<sub>2</sub>S to KI, NH<sub>3</sub>, HNO<sub>3</sub>, Na<sub>2</sub>CO<sub>3</sub>, and AgNO<sub>3</sub> resulted in a separation with a translucent brownish-grey solution on top, and a transparent solution on the bottom. Over time, the overall solution became homogenous. The addition released a pungent rotten odour.

(j) Adding 3 drops of 0.1M Na<sub>2</sub>S to the system, a brown solution is produced with a strong rotten odour. When Na<sub>2</sub>S is added to the system, it dissociates into Na<sup>+</sup> and S<sup>2-</sup>. The S<sup>2-</sup> ions react with Ag<sup>+</sup>

ions present from equation B. To produce  $\text{Ag}_2\text{S}$  solid. This means that the brown liquid observed contained precipitate of  $\text{Ag}_2\text{S}$ . As the concentration of  $\text{Na}_2\text{S}$  increases in the system, the equation H. shifts right to produce more  $\text{Ag}_2\text{S}$  solid.

### Part 3: Buffer Solution and Blood pH



- Distilled water - clear, transparent liquid with no odour and low viscosity.
  - Initial pH for 100 mL of distilled water: pH=4.50
- $\text{NaHCO}_3$  - white, powdered, crystallized, finely granulated, low viscosity.
- $\text{NH}_4\text{Cl}$  - white, powdered, crystallized, finely granulated, occasional condensed chunks of powder stuck together.

3.1 Adding 2.50 g of powdered  $\text{NaHCO}_3$  to the beaker allowed the salt to dissolve completely and increase the pH of the solution to 8.26.

3.2 Adding 30 mL of 0.1 M  $\text{HCl}$  to the  $\text{NaHCO}_3$  solution produced several bubbles and decreased the pH to 7.50.

3.3 Adding 10 mL of 0.85% lactic acid reduced the amount of bubbles dissolved in the solution, where many bubbles stuck to the side of the beaker. This change reduced the pH to 7.38.

3.4 Increasing the stirring rate increased the activity of the bubbles, making them spin faster. The pH was slightly increased to 7.43.

3.5 Adding 0.50 g of more powdered  $\text{NaHCO}_3$  slightly increased the pH to 7.60 after the salt was completely dissolved. Very few bubbles were dissolved in the solution.

3.6 Adding an extra 0.50 g of  $\text{NaHCO}_3$  slightly increased the solution's pH to 7.69. There was still a very minimal amount of bubbles dissolved in the solution.

3.7 Adding a rough, hard, translucent, white pellet of  $\text{CO}_2$  to the solution slightly decreased the pH to 7.38. This manipulation allowed the solution to release dense, smoky vapour that escaped the beaker.

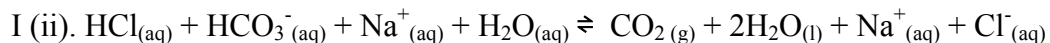
3.8 Adding 0.40 g of powdered  $\text{NH}_4\text{Cl}$  resulted in bubbles dissolved in the solution and a slight

pH increase of 7.42.

(k) The pH of the distilled water was 4.50. A pH of 7.00 was expected for the distilled water, because pure water is considered a neutral substance. This proves that the water is slightly acidic. The water is a lower pH than expected because there may be extra unknown ions dissolved in the water for the purpose of reducing damage to piping.

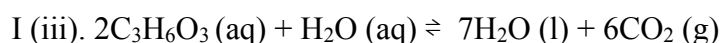


(l) The pH of the sodium bicarbonate solution was 8.26. The dissociation of sodium bicarbonate into hydrogen carbonate ions and sodium ions may have caused this increase in pH.  $\text{HCO}_3^-$  ions are responsible for the increase in pH, because sodium ions do not have any acidic or basic properties. Sodium ions do not dissociate hydrogen ions or hydroxide ions.

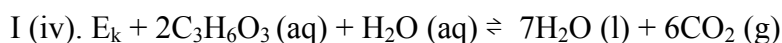


(m) HCl acid is added to the solution in attempt to neutralize it. Adding the acid to the solution of hydrogen carbonate ion created gaseous carbon dioxide.

(n) When HCl acid was added to the hydrogen carbonate ion, several bubbles were formed because the reaction produced  $\text{CO}_2$  gas. The pH of the solution reduced to 7.50. This pH is much higher than the pH of the distilled water and more accurate to the water's predicted pH, however only slightly basic. The solution has a pH slightly higher than that of the normal blood pH range, which would cause alkalosis. In the solution, there are carbon dioxide molecules, water molecules, sodium ions, and chloride ions present. Of all the species, water molecules are twice as more concentrated than the others. Based on the theory on buffer solutions, this is the correct ratio that should be produced.

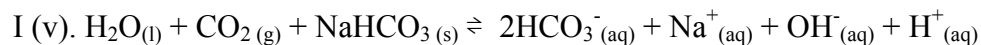


(o) The pH of the solution was reduced to 7.38. This situation was intended to simulate the body's metabolism of lactic acid, and the carbonic acid buffer system's ability to resist the change in pH. Adding 10 mL of 0.85% lactic acid reduced the amount of bubbles spinning in the solution. The majority of bubbles cohered to the beaker wall.

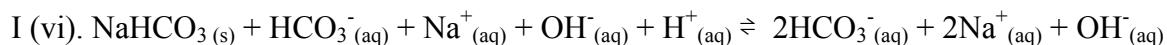


(p) Increasing the stirring rate increased the bubbles' activity by making them spin faster. The pH was slightly increased to 7.43. These observations correspond with the human body's response to hyperventilation. Overall, inputting more energy into the system favours the forward reaction. By increasing the stirring rate, the reactant particles possess more kinetic energy and a greater

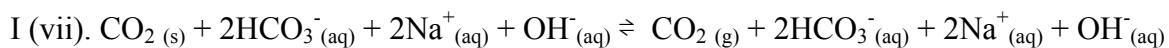
probability of colliding. To maintain equilibrium, the production and expulsion of carbon dioxide should advance with the number of molecular collisions, thus the solution should become more basic.



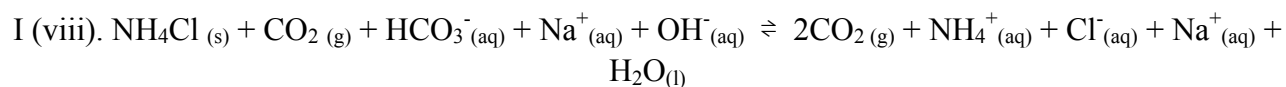
(q) The pH of the solution was increased to 7.60. Adding 0.50 g of more powdered  $\text{NaHCO}_3$  severely reduced the amount of bubbles that were dissolved in solution. The reason for this increase is because dissolving  $\text{NaHCO}_3_{(s)}$  dissociated  $\text{OH}^-$  ions and  $\text{HCO}_3^-$  ions into the solution. Adding this extra reactant was compensated for by increasing the production of these ions.



(r) The pH of the solution was 7.69. Adding an extra 0.50 g of  $\text{NaHCO}_3$  barely changed the appearance of the solution. There was still a very minimal amount of bubbles dissolved in the solution. At this point, the body experiences alkalosis because there are not enough  $\text{HCO}_3^-$  (conjugate acid) ions produced to compensate for the additional  $\text{NaHCO}_3$  (weak base).



(s) The pH of the solution was 7.38. After placing the solid into the beaker, the dry ice immediately vaporized and allowed a dense, smoky vapour to escape from the solution. By adding an extra reagent, the forward reaction was favoured. This produced gaseous  $\text{CO}_2$  to consume the solid  $\text{CO}_2$ .



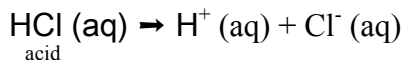
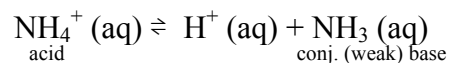
(t) The pH of the solution was 7.42. Adding 0.40 g of powdered  $\text{NH}_4\text{Cl}$  resulted in bubbles being formed and dissolving into the solution. The additional reagent allowed more  $\text{H}^+$  ions to dissociate in the solution, which should have further decreased the pH, however this was not the case. The apparent increase in pH compared to when the  $\text{CO}_2$  pellet was added may be because some of the  $\text{H}^+$  ions bind with  $\text{OH}^-$  ions to form neutral water, and gaseous  $\text{CO}_2$  continued to escape the beaker to relieve the acidity.

## Conclusion

This experiment supports Le Chatelier's Principle and shows how changing the concentration of certain reactant or products can affect the equilibrium of a reaction. Increasing the concentration of a reagent on one side of a chemical equation causes the equilibrium position of that reaction to shift to the side that will consume the excess reagent. Additionally, the lab supports the fact that buffers maintain their pH levels better than distilled water when mixed with an acid or a base.

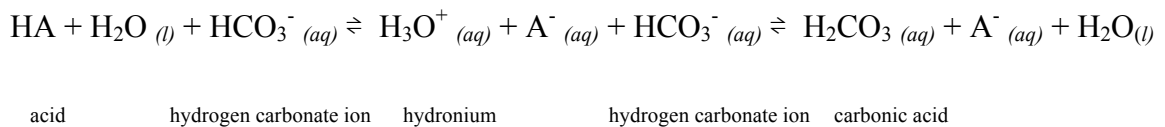
## Questions to be answered at End of Report

1. (a) The ammonium ion is used as the acid source instead of HCl because  $\text{NH}_4^+$  dissociates into  $\text{NH}_3$ , which is a weak base. In order to establish a buffer system, a solution requires a weak base and its conjugate acid for the reaction to be reversible. As a result, the ammonium ion is a more suitable acid than HCl because HCl dissociates chloride ions, which does not have any acidic or basic properties. HCl as an acid would not have a conjugate base.

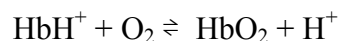


- (b)  $\text{NH}_4\text{Cl}$  is used instead of other ammonium compounds because it is safe to use in this experiment and easily available.

2. Each time an acidic substance was added to the beaker, the pH of the solution would slightly drop. The pH would drop according to the pH scale because the concentration of  $H^+$  or  $H_3O^+$  ions dissociated in solution increases, approaching  $1 \cdot 10^{-1}$  mol/L. (ref. my textbook Ch.16 671-673). However, it would only drop slightly because the presence of  $HCO_3^-$  ions that establish a buffer system. If the solution becomes too acidic, some of the  $H^+$  ions may bind with  $HCO_3^-$  ions to form  $H_2CO_3$  (Mirrakhimov et al., 2017). This process neutralizes some of the  $H^+$  or  $H_3O^+$  ions.



3. Acidosis would excessively decrease the patient's ability to transport oxygen. An elevated concentration of  $H^+$  ions circulating the blood would result in the reverse reaction being favoured. The reaction would shift so that more hemoglobin carries hydrogen ions, and more oxygen is left independent in the lungs. As a result, there is a reduced amount of oxygen being diffused into the blood to be transported by hemoglobin.



4. The reason why adding HCl to the test tube of water and blood made a darker solution of blood can be explained using the chemical equation from question 3. When HCl is added to the system, there is an increase in  $H^+$  ions which shifts the equilibrium of the equation from question 3 left. Shifting the equilibrium left means more oxygen is being produced. When oxygen binds to hemoglobin, the blood is brighter, however when the equilibrium shifts left and there are less  $O_2$  molecules bound to hemoglobin molecules, this causes the blood to be darker.

5. If a fresh sample of soda with a pH of 2.92 was placed on a magnetic stirrer made to go flat, the pH of the flat soda should be higher than the pH of the fresh soda. Looking at the equation provided, taking off the lid of the soda and putting it on the stirrer, the  $\text{CO}_2$  is exiting therefore the system is decreasing in concentration of  $\text{CO}_2$ . The second part of the equation has no equilibrium as it is an open system and so it is not possible to replenish the  $\text{CO}_2$  lost because the reaction will not shift left. As the reaction is shifting right, the  $\text{H}^+$  ions are being used up causing the soda's pH to increase.



6. Since chickens pant in hot weather, the pH of their blood will rise and the strength of the eggshells they produce will weaken. When the chickens pant, it equates to humans hyperventilating, which means that more  $\text{CO}_2$  is lost and there is less  $\text{H}_3\text{O}^+$ . Looking at the equation provided with the question, when there is a decrease in concentration of  $\text{CO}_2$ , the equilibrium shifts left to replenish the  $\text{CO}_2$  and at the same time  $\text{H}_3\text{O}^+$  is being used up. This causes the pH to rise and be more basic. Additionally, the equilibrium shifting left means there is less carbonate molecules and therefore less  $\text{Ca}^+$  ions are binding to the carbonate so the eggshells are weaker.