Introduction:

We originally looked at chemical equations as being non-reversible and only a one way reaction, as implied by a single arrow from reactants to products. Realistically, however, reactions actually reach a point where the rate of reaction of reactants is equal to the rate of the reaction of products becoming reactants. This dynamic state is called equilibrium. The reaction does not come to a stop, it continues going back and forth, converting reactants to products and products to reactants at a constant rate. The amount of reactants being formed and products being formed varies according to different factors. Using the equilibrium constant, K, we can figure out which side of the reaction is favoured in an aqueous reaction such as:

$$aA + bB \rightleftharpoons cC + dD \quad + \text{heat} \qquad [1]$$

The equilibrium constant expression can be written as :

$$K= [C]^c [D]^d / [A]^a[B]^b \qquad [2]$$

If the product side of the reaction is favoured we can see that the equilibrium constant would be large and vice versa if the reaction is reactant-favoured.

A reaction will always try to remain in equilibrium. As stated in *Le Chatelier's Principle*, a reaction will always try to minimize the effect of a change or disturbance on a system at equilibrium. There are a number of factors that have an effect on a system at equilibrium. What would happen if we added more reactant or product? If more reactant (B) is added to the equation [1] above, the equilibrium will shift to the product side to compensate for the excess B. Another factor is temperature. The shift of the equilibrium depends on whether or not the reaction is endothermic or exothermic. For example, equation [1] is an exothermic equation which means an increase in temperature will cause a shift to the left in order to consume the heat being added. Addition of a catalyst or inert gas will not have any effect on equilibrium.

After discussing how equilibrium can favour a products or reactants side, an important thing to know is how strong or weak a species is in a reaction. A strong species is defined as being able to dissociate completely. An example is a reaction between HCL and NaOH. Both the acid and the base are termed strong as they dissociate completely. A weak acid/base is a compound that does not dissociate completely when placed in water. Weak acids and bases have an equilibrium which strongly favours the left side of the equation. K_a is the equilibrium constant for an acid, also called the acid dissociation constant. Water can act as an acid or a base. Autoprotolysis is a process where a water molecule donates a proton to another and this leads us to the expression for the equilibrium constant of water which is $K_w = 1x10^{-14}$. This kind of species is classified as amphoteric. K_b is the base dissociation constant. We can express the relationship between these constants as:

$$K_w = K_a \cdot K_b = 1x10^{-14} \qquad [3]$$

The concentration of hydrogen and hydroxyl ions must be equal in pure water and have a value of $1x10^{-7}$ mol/L. If there is a greater concentration of hydrogen, a solution will be acidic. If there is a greater concentration of hydroxyl ions, a solution will be basic. The product of the concentration of hydrogen and concentration of hydroxyl ions will always equal $1x10^{-14}$.

The pH scale is defined logarithmically with the expression:

$$pH = -\log[H^+(aq)]$$

In general,

$$pX = -\log X$$

To determine the pH of a solution we use pH meters or indicators. Indicators are usually weak acids that dissociate to different extents, shown through change in colour, depending on the acidity of whatever solution it is placed in.

Our bodies function ideally at a pH between 7.35 and 7.45. A disturbance to this natural pH level can be harmful to the body. Our bodies have a buffer system in the blood that prevents alkalosis (pH higher than 7.45) and acidosis (pH lower than 7.35). The chemical buffer is made up of 3 separate buffers but this experiment will be focusing on the carbonate/ carbonic acid buffer. To create a buffer, there needs to be a sufficient amount of both the acid and its conjugate base in the solution. In the blood carbonic acid (weak) is at equilibrium with hydrogen carbonate ion. This buffer contains a sufficient amount of the hydrogen carbonate ion as there are more acids produced in the body than there are bases. The reaction for this buffer is expressed as:

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

The metabolic acids that are produced are neutralized by the high concentration of the bicarbonate ion in the buffer.

Procedure & Materials:

Refer to lab website

Observations:

Part 1-

a) After adding $CuSO_4$, the test tube contained a light blue solution caused by the copper ion.
b) After adding NH_3 (a clear solution), the solution in the test tube turned into a dark blue colour due to the ammonia.
c) After adding HCl, the solution slowly returns to the original light blue colour

1. Adding more NH_3 caused the solution to turn back to the darker blue colour and then adding more HCl caused the solution to go back to the light blue colour

Discussion:

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

Adding the initial amount of NH_3 to the test tube, according to the equation of the reaction above, meant that the reaction took place and product was formed. This product had a darker blue colour. After adding HCl, it added more product and according to Chatelier's principle, equilibrium must shift to compensate, therefore, the solution turned a lighter blue colour after it shifted to the left side. When repeating the steps, equilibrium had continued to shift to compensate for the addition of reactants and products.

Part 2-

a) After adding Na_2CO_3, solution in test tube was clear

b) After adding $AgNO_3$, solution turned into a light brown/beige colour. The equation responsible for this reaction is $2AgNO_3$ (aq) + Na_2CO_3 (aq) $\rightleftharpoons$ Ag_2CO_3 (s) + 2 $NaNO_3$ (aq). The colour is a result of the precipitate Ag_2CO_3.

c) After adding HNO_3, the solution became clear again. $2H^+$ (aq) + CO_3^{2-} (aq) $\rightleftharpoons$ H_2CO_3(aq) H_2O(l) + CO_2 (g) is the equation responsible. The addition of HNO_3 reacts with Ag_2CO_3 to form H_2CO_3 which is a product and will cause the equilibrium to shift to the left. This reaction is reversible as it reaches a state of equilibrium.

d) After adding HCl, the solution became cloudy and opaque. Ag^+ (aq) + Cl^- (aq) $\rightleftharpoons$ AgCl(s) is the equation that explains this observation. The cloudiness was a result of the precipitate AgCl.

e) After adding NH_3, solution slowly turned clear again. Ag^+ (aq) + $2NH_3$ (aq) $\rightleftharpoons$ $[Ag(NH_3)_2]^+$ (aq) is the equation responsible. The added NH_3 reacted with the Ag^+ ions to form a clear solution.

2. After adding more HNO_3, the test tube let out steam and test tube heated up. Adding more NH_3 caused the test tube to cool down. The same changes were not observed. H^+ (aq) + NH_3 (aq) $\rightleftharpoons$ NH_4^+ (aq). This is the chemical equation responsible for the reaction. The added HNO_3 reacted with the NH_3 already in the test tube from the previous reaction to form NH_4. The heat observed shows that

this reaction was exothermic. When more NH_3 is added this caused a reverse reaction due to Chatelier's Principle to compensate, hence the temperature decrease.

f) Adding KI caused the solution to become yellowish and opaque. Ag^+ (aq) + I- (aq) ⇌ AgI(s) is the equation responsible for the reaction. The I ions reacted with the Ag already in the test tube to create the precipitate shown in the equation (AgI).

g) After adding Na_2S, the solution turned dark brown/black. $2Ag^+$ (aq) + S_2^- (aq) ⇌ Ag_2S(s). The black colour was due to the precipitate formed in the equation.

Part 3-

a) The pH of the water alone was 5.5 which is not to be expected as water should be neutral and have a pH of around 7. This means the water in the beaker was more on the acidic side. Pure water is supposed to have a pH of 7 however tap water is not pure water as it also contains other ions in it.

b) After adding $NaHCO_2$, the pH went up to 8.21. The carbonate reacted with the water in this equation: $2H_2O$ (l) + CO_2 (g) ⇌ H_2CO_3 (aq) + H_2O (l) ⇌ HCO_3- (aq) + H_3O+ (aq). The ions responsible for making the pH more basic is the carbonate.

c) After adding HCl the pH lowered to 7.21. This reverses the effect of the addition of $NaHCO_2$. This created a buffer solution.

d) After adding lactic acid the pH decreased to 7.09. The solution in the beaker is more neutral compared to the original pH of the distilled water. In comparison to the pH of blood it is still a little on the acidic side but not too far from the pH level required in the blood. $2H_2O$ (l) + CO_2 (g) ⇌ H_2CO_3 (aq) + H_2O (l) ⇌ HCO_3- (aq) + H_3O+ (aq). There is more acid than base in this solution as the pH is more acidic

e) Increasing the stir rate raised the pH to 7.13. This simulates the condition acidosis where the body increases the rate of respiration by expiring more CO_2 when the pH level of the body is too low. H_2CO_3 (aq) ⇌ H_2O (l) + CO_2 (g).

f) After adding $NaHCO_3$, the pH level increased to 7.33. With vigorous stirring the pH level would increase more.

g) After adding the H_2CO_3 increased the pH to 7.70 which is much more basic. since more reactant is added which shifts the equilibrium into producing more products as shown in the equation mentioned above. This causes an increase in the pH of the solution unlike what was expected in the lab report procedure.

h) After adding the dry ice, the solution released large amounts of steam and became cloudy at the top of the beaker. The pH decreased to 7.03. Equilibrium shifted as more CO_2 was added and had to compensate by going through with the reverse reaction.

i) After adding, NH_4Cl, the steam being released decreased significantly. The pH increased to 7.10.

<u>Questions:</u>

1. a) In order to maintain a pH like the one in our blood we need to use compounds that will not fully dissociate and significantly change the pH inside our bodies. This is why we need to use weak acids and weak bases in order to change the pH slightly. HCl is a very strong acid and will completely dissociate, releasing may hydrogen ions causing the solution to become too acidic. NH_4Cl is a weak acid which releases less protons than HCl would thus minimally altering the pH to get it to the perfect pH level.

   b) The point of this experiment is to mimic the buffer system that is actually happening in our bodies to maintain our blood pH level. Our bodies use up chemicals that are already present in the system, therefore we use NH_4Cl because it is already present in the system. NH_4Cl is also the conjugate acid in the reaction between HCl and NH_4NO_3.

2. Every time we added an acid to the solution in the beaker the pH would decrease to less than 7. Just as measured in LabQuest 2, the pH indicator implied the solution was acidic as well with the lighter blue colour. When adding these acids to our solution we keep in mind there are also basic compounds within the solution. This means the acids and the bases will react and can create a neutralization reaction where water and a salt are formed. It is all depending on the concentrations of the acids and bases. If the concentrations were the same then a simple neutralization reaction would occur. However if the concentration of the acid was less than the concentration of the base then the acid would be fully used up with an excess base leaving the pH to be slightly basic. Vice versa if the concentration of the acid was greater than the concentration of the base.

3. There are many metabolic processes that happen inside our bodies and most of these metabolic reactions result in the formation of acidic compounds. This will lower the pH level and increase the concentration of H^+ in the body which, according to Le Chatelier's Principle and the equation given, shift the equilibrium to the left. What is occurring is acidosis which is when the blood pH level is lower than 7.35. This shift results in reduction in the supply of oxygen to the cells as it is favouring the reactants side rather than the oxyhemoglobin. The lack of oxygen in cells causes harm and stress to the human body.

4. Oxygenated blood has a bright red colour as the concentration of oxyhemoglobin is high. Deoxygenated blood has a dark red colour as the concentration of oxyhemoglobin is low. The

test tube on the left has a brighter red colour which indicates a high concentration of oxyhemoglobin. The blood was added into the test tube on the right and the hemoglobin reacted with the oxygen and created more of the oxyhemoglobin. The addition of HCl to a solution like this (such as the right test tube) will create an increase in hydrogen protons as HCl is a strong acid and dissociates completely. The addition of this hydrogen caused the equilibrium to shift to the reactants side as according to Le Chatelier's Principle, resulting in a darker red colour as the concentration of oxyhemoglobin was much lower.

5. The purpose of the magnetic stirrer is to make the soda go flat. If the soda were to be made flat, the pH should increase from its original level. Stirring the solution should cause the pH to increase similarly to when our bodies increase the rate of respiration to expire more CO_2. Stirring the soda will cause the solution to expire more CO_2 and will result in a flat soda.

6. As chickens pant in hot climate, this means their rate of respiration increases. They are exhaling more CO_2 causing a disruption to the bicarbonate/carbonate buffer system. This shifts equilibrium to the right in order to replenish the large decrease of CO_2. This means there is a decrease in the concentration of H_3O^+ therefore, increasing the pH level. This shift in equilibrium also causes a larger production of carbonate which chickens pair with calcium ions to strengthen their eggshells. This means panting in hot weather will increase the pH level of their blood and strengthen their eggshells.

Conclusion:

Equilibrium is when the rate of the reaction of the reactants is equal to the rate of reaction of the products. Through experimentation we have seen how a disturbance to an equilibrium system is compensated with a reverse reaction through Le Chatelier's Principle. In part 1, $CuSO_4$ reacted with the NH_3 and by adding more product or reactant, the equilibrium shift was observed through colour change. In part 2, it was observed how multiple reactions can continuously occur as we added Na_2CO_3 ,$AgNO_3$, HNO_3, HCl, KI, and Na_2S after one another and observed how the equilibrium shifted, again, through colour change of the solution. The reverse reaction can occur continuously. Finally in part 3 we mimicked the bicarbonate/carbonate buffer system that is used in the human body to maintain blood pH level. This buffer system is kept at equilibrium to prevent acidosis or alkalosis and keeps the pH level at the perfect level of 7.35-7.45. The system was successful.

RAW DATA:

Part 1 :

$CuSO_4$ + NH_3 (concentrated)

$CuSO_4$ → Light
when NH_3 was added
it became light blue

+ HCl (0.1 mol/L) = returns
to its original color

If you repeat step 2-3 the
colour will change colors from
light blue to dark blue

NH_3 was a clear solution

The colour observed was
due to the pH
↳ the light blue indicated
acidity while dark blue
indicated that it was
basic

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Part 2:

Na_2CO_3 + $AgNO_3$

↓ ↓

clear solution clear solution

① When $AgNO_3$ is added the color of the solution turned into a paige / ~~lightish brown~~ color

② + HNO_3 → Clear solution

After the second drop the solution became clear

③ +HCl → clear solution

solution became slightly opaque / cloudy after the 3rd drop. After the 4th drop the solution was completely opaque

④ + NH_3 → clear solution

After adding several drops the solution became opaque @ the bottom & a clear layer was seen at the top. After 1 drop it became completely clear

⑤ + HNO_3 → clear solution

The test tube started to steam after a couple drops & it became hot indicating that the reaction is exothermic

⑥ + NH_3 → clear solution

Test tube became cold indicating an endothermic reaction, bottom was slightly opaque

* note that each step is reversing the effect of the previous step

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and the solution become completely clear

(7) + KI → clear solution

The ~~solution~~ test tube was cloudy & the solution become cloudy on top & the bottom layer remained clear. After a couple more drops it become yellowish opaque

(8) + Na_2S → clear solution

The solution turned blackish brown at the top and it remained yellowish opague at the bottom of the test tube. After adding two more drops & slightly agitating 3/4 of the solution become blackish brown (grey) but the bottom of the solution remained yellowish

Lab # 3

Lab Procedure:
refer to bright space → no modifications were made to the procedure online

Part 3:
observations:
→ when measuring the salt the weight of the paper it is placed on is not considered
→ initial pH of the water is about 5.50
→ After the salt dissolved the pH became 8.21 (basic
→ After adding HCl the pH became 7.21
→ After adding lactic acid the pH became 7.09
→ when increasing the rate the pH also increased to 7.13 after a couple seconds
→ it continues to slowly increase
→ Adding sodium bicarbonate

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increased the pH ~~[illegible]~~ to
~~[illegible]~~ 7.33
→ Adding another 0.5g of
Sodium bicarbonate it (pH)
~~went up to 7.71~~
→ After adding ~~dry~~ ice (CO_2)
the solution was releasing
large amounts of steam & bubbling
and formed a layer of steam
on top (cloudy)
→ the pH of the
solution decreased to
7.03
→ After adding ammonium
chloride the amount of
steam being released decreased
significantly (it was more like
puffs of smoke)
→ eventually the solution stopped
releasing this smoke and
returned to a pH of
7.10