TOTAL REPORT GRADE = 42/55

Experiment 3: Equilibria

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Introduction:

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The purpose of this lab is to verify the concepts of buffer, chemical equilibrium, weak acids and bases, pH, indicators, acid dissociation and Le Chatelier's principle. The lab is divided on three parts where each of the concepts is studied with a related level of details. In the first part, "Equilibrium Shift", pH, indicators and chemical equilibrium is studied. It is mainly focused on the effects of adding and consuming reactants on a system at equilibrium. Second experiment is based on the establishment of multiple equilibria using silver ion. It provides details related to the interaction of chemical equilibriums between each other forming an equilibrium system. The goal of the third experiment is a preparation of a buffer solution and the study of the effect of acids and bases on a buffer system that simulates a metabolic reaction. (Refer to the lab manual) Limitations to this lab include the volumes of the compounds added to the test tubes during first and second parts of the experiment. Furthermore, the temperature changes were not recorded, as the compound part of test tubes and beaker were not directly touched or measured with a dedicated equipment.

Materials:

- 0.1 M HCl
- 1 M HCl
- 0.1 M $CuSO_4$
- 0.1 M Na_2S
- 0.01M $AgNO_3$
- 0.1 M Na_2CO_3
- Solid $NaHCO_3$
- 0.85% lactic acid
- NH_4Cl
- 0.1 M KI
- Concentrated ammonia, NH_3
- 6 M HNO_3

Apparatus:

- Test tubes
- Transfer pipets
- LabQuest 2
- pH probe attachment for the LabQuest 2
- Balance
- Graduated cylinder
- Spatula
- Weighing paper
- Stir plate
- Test tube racks
- Stir bar (magnet)
- 150 mL and 250 mL beakers

Procedure:

Equilibrium Shift:

1. Add 1 mL (~20 drops) of 0.1 mol/L $CuSO_4$ to a clean test tube.
2. Dropwise, add concentrated NH_3 solution until change is observed. Do not smell the vapors of ammonia.
3. Dropwise, add 1 mol/L HCl until you observe a change.
4. Repeat steps 1 to 3 recording the secondary observations.

Multiple Equilibria

1. Place 0.5 mL of 0.1 mol/L Na_2CO_3 solution in a test tube.

2. Add 0.5 mL of 0.01 mol/L $AgNO_3$ solution to the test tube.
3. Dropwise, add 6 mol/L HNO_3 until change is observed. Make sure not to spill the acid, as it is very corrosive.
4. When the solution is clear, add 0.1 mol/L HCl dropwise until you see a change.
5. Dropwise, add concentrated NH_3 until you see a change.
6. Repeat steps from 7 to 9 recording the second set of observations.
7. Add 0.1 mol/L KI dropwise until you see a change.
8. Dropwise, add 0.1 mol/L Na_2S until you see a change.

Buffer Solution and Blood pH

1. Place 100 mL of distilled water in a 150 mL beaker. Place the beaker on a stir plate. Insert a stir bar (magnet) in the beaker, turn the stir plate on slowly and create a smooth vortex in the beaker.
2. Gently insert the pH probe into the water such that the magnet does not hit the probe (it will break the glass inside the probe!!). The tip of the probe must be IN the water for the pH reading to be correct. Wait a few minutes.
3. Measure 2.50 g of sodium hydrogen carbonate, $NaHCO_3$ and add it to the beaker. Allow the solution to stir until the salt dissolves completely.
4. Add 30 mL of 0.1 mol/L $HCl_{(aq)}$ to the beaker.
5. Add 10 mL of 0.85% lactic acid to the beaker. While the body normally produces a small amount of lactic acid, this step simulates the formation of a large amount of lactic acid.
6. Increase the stirring rate of the stir plate. Be careful to do not collide stir bar with the pH probe.
7. Add approximately 0.50 g of sodium bicarbonate, $NaHCO_3$ to the beaker.
8. Add another 0.50 g of sodium bicarbonate, $NaHCO_3$ to the beaker.
9. Add a pellet of dry ice.
10. Add 0.40 g of ammonium chloride, NH_4Cl, to the beaker.

Cleaning Up

1. Empty all the transfer pipets and rinse well with warm water.
2. Discard all solutions in the waste container, except for the silver solutions, which should be disposed in the container marked "silver wastes".
3. Clean your bench thoroughly and throw away all paper towels.

Safety Precautions:

- Protection glasses and a lab coat have to be weared during the lab.
- Very cold objects and compounds should not be touched without the proper equipment.
- Toes covering shoes have to be weared during the lab.
- Long hair should be tied during the lab.
- After completing the practical part of the lab, the used material should be disposed according to the orders of the lab personnel.
- Acids and bases, even in dilute solutions, are corrosive. Ensure that if any acid or base is spilled, it should be cleaned up immediately. If any is split on skin or clothing, rinse thoroughly with water for at least 15 minutes.

Observations and Discussions:

Part 1. Equilibrium Shift

$CuSO_4$ solution is transparent blue. It has this color, due to the presence of Cu^{2+} ions in it (a). Addition of concentrated NH_3 turned the solution transparent dark blue. According to the balanced chemical formula of the reaction studied, $[Cu(H_2O)_4]^{2+}$ (aq) + 4 NH_3 (aq) ⇌ $[Cu(NH_3)_4]^{2+}$ (aq) + 4 H_2O (l), change in color can be explained by the formation of $[Cu(NH_3)_4]^{2+}$ (aq) ions. Thus, we can state that the addition of the ammonia changes the color of the light blue copper solution to dark blue (b). Added HCl solution turned the color to cloudy pale blue. The color change observed is a result of formation of NH_4^- ion from the HCl and NH_3. As more reactant was added, the equilibrium shifted towards the products (c). Repetition of steps 2 to 3 showed the studied reaction to be a good example of an equilibrium. Initial reaction between NH_3 and HCl caused the solution turn cloudy pale blue. Addition of extra NH_3 shifted the equilibrium to the reactants and the transparent dark blue expression observed during the second step appeared. Lastly, secondary addition of HCl changed the color back to cloudy pale blue. It clearly demonstrates a reversible reaction (1). Third repetition of the reaction did not yield any precipitate, which can be explained by the too low concentration of HCl added to the solution and a relatively strong buffer created during the previous steps (bonus question).

Part 2. Multiple Equilibria

The Na_2CO_3 solution is transparent (d). Introduction of $AgNO_3$ modified the color to opaque light brown. The reaction observed is outlined in the following equation: $2AgNO_3$ (aq) + Na_2CO_3 (aq) ⇌ Ag_2CO_3 (s) + $2NaNO_3$ (aq). According to the states provided in the lab manual, Ag_2CO_3 is a precipitate, which explains the opaque aspect of the solution. Regarding the light brown, Ag_2CO_3 could not cause this change, as it is a solid precipitate. Thereupon, the color change was caused by $NaNO_3$ product (e). Addition of HNO_3 reversed the color change and returned it to the clear state. This indicates that the studied reaction is reversible. Chemistry wise, it means that the reaction is not stable and its compounds constantly switch between reactants to products and back until an equilibrium will be achieved (f). HCl modified the appearance of the solution turning it opaque white. Presence of HCl creates new interrelated reactions, therefore, creating multiple equilibria (g). Added NH_3 makes the solution clear. This chemical reaction is Ag^+(aq) + $2NH_3$(aq) ⇌ $[Ag(NH_3)_2]^+$(aq).[1] As the positive ion is formed, H^+ should be responsible for the color alteration (h). Repetition of steps from 7 to 9 leaded to the results that differ from the original one. HNO_3 caused the formation of vapor and a temperature increase. Furthermore, it changed the color to slightly opaque white. The formula that can explain this behaviour is $2H^+$(aq) + CO_3^{2-}(aq) ⇌ H_2CO_3(aq) → H_2O(l) + CO_2(g). The observed vapor turned to be the CO_2 gas during this reaction. Furthermore, secondary introduction of NH_3 resulted in an identical change as before. The solution turned transparent clear. The CO_2 release can be explained as a way to reach equilibrium by the reaction. This is further proven as a result of the multiple reaction (2). KI altered the solution to be cloudy white. The chemical formula for the following reaction is Ag^+(aq) + I^-(aq) ⇌ AgI(s). The precipitate outlined in the following reaction is responsible for formation of a cloudiness (i). Na_2S turned the color of a compound into cloudy grey color. In the case of this reagent, Na^+ is a spectator ion. Hence, only S^{2-} ion reacts. Following the provided formula, $2Ag^+$(aq) + S^{2-}(aq) ⇌ Ag_2S(s) (j).

Part 3. Buffers

The starting pH of water was found to be 4.58 pH. Usually, the pH of distilled water equals to 7 pH. The gap between theoretical and practical measurements of pH can be due to the

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storage liquid of a pH probe or due to the impurities in the water (k). $NaHCO_3$ submerged in water resulted in a raise of pH to 8.10. This dissolution can be outlined by the following formula, $NaHCO_3(s) + H_2O(l) \rightleftharpoons NaCO_3^- + H_3O^+$. It is clear that the pH change occurs due to the presence of H_3O^+ ion present in the solution (k). The HCl acid neutralizes the basic solution studied. HCl was added to simulate a behaviour of acids inside the human body that are naturally neutralized.[2] This reaction of neutralization can be outlined by the following formula: $NaHCO_3(s) + HCl \rightarrow CO_2(g) + H_2O(l) + NaCl(aq)$ (m). A chemical reaction of neutralization occurs and it is indicated by the two observations. Solution releases CO_2 and the pH drops from 8.00 to 7.19. The achieved pH is very close to a theoretical pH of water that equals 7.00. Furthermore, the resulting pH is near the normal pH of human blood that equals to 7.35.[3] Additionally, this was an expected result according to the buffer theory outlined in the pre-lab manual. There are two species in this solution and they were identified to be Na^+ and Cl^- (n). Addition of the lactic acid lowered the pH from 7.19 to 7.06. Making connections with a living body, this mimics the release of the lactic acid from the muscles during the anaerobic cellular respiration. The reaction between the lactic acid and distilled water can be outlined by the following reaction: $CH_3CH(OH)COOH(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3CH(OH)COO^-(aq)$ (o). There is a direct relationship between the pH of the solution and the speed of the stirring, because it increases an amount of bonds created in the solution and more CO_2 will be released. This is predicted, as the effusion of CO_2 will lead to the equilibrium shift and, therefore, will raise the pH. The reaction outlining formula: $2H_2O(l) + CO_2(g) \rightleftharpoons H_2CO_3(aq) + H_2O(aq) \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq)$ (p). The raise of pH was recorded after the addition of $NaHCO_3$ from 7.06 to 7.17. It can be derived that the equilibrium constant is larger than one, as it favours the products. This creates a buffer (q). The pH recorded after the addition of another 0.50g $NaHCO_3$ equals to the 7.26. This pH reading considered to be a state of acidosis. Introduction of a reactant should move the equilibrium to the right, producing the acid (r). Submerging of dry ice into the solution did not undergo successfully, as it gets pulled up to the surface of the solution. Dry ice melts and release plenty of CO_2 gas, which forms a white steam. The pH reading changed from 7.26 to 7.30. It can be explained by the fact that formation of CO_2 formed in the body pushes the pH of the blood to the normal level of 7.40.[2] However, CO_2 should cause the equilibrium shift and lower the pH as outlined in the balanced chemical equation: $2H_2O(l) + CO_2(g) \rightleftharpoons H_2CO_3(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H_3O^+$. Theoretically, the pH has to rather decrease, as slightly acidic H_2CO_3 is formed during the dissolution of CO_2 in water. None the less, secondary introduction of the dry ice stabilized the pH reading at 7.30, meaning that the equilibrium was achieved (s). Addition of a NH_4Cl dropped the pH to 7.20. This pH change is low, due to the strong buffer created during the previous steps. Furthermore, the NH_4Cl will not completely ionize in water, since the relatively stronger acid, HCl, is already present in the solution. This formula outlines the dissolution of NH_4Cl in water: $NH_4Cl(aq) + H_2O(l) \rightleftharpoons NH_3Cl^-(aq) + H_3O^+(aq)$ (t).

Conclusion: Conclusion 2/2

Each of the parts went as expected. In the first part, the equilibrium was constantly changed by the addition of acids and bases. Relative observations proved the equilibrium change. In the second part, multiple equilibria was formed through the introduction of new ions and a formation of multiple interrelated reactions. The observations were once again a key to understanding the equilibrium shifts. Lastly, the simulation of a buffer in a blood was created successfully with minor exceptions of a low initial pH of water that can be explained by the impurities in the water and the acidity of the pH probe storage liquid.

References:

1. Flashcard Machine - create, study and share online flash cards. (n.d.). Retrieved from https://www.flashcardmachine.com/transition-metalcomplexescolours.html
2. Boyers, L. (2018, July 31). Acids That Are Important to the Human Body. Retrieved from https://healthyeating.sfgate.com/acids-important-human-body-9207.html
3. Definition of Blood pH. (n.d.). Retrieved from https://www.medicinenet.com/script/main/art.asp?articlekey=10001

Questions to be answered at End of Report Questions 7.5/9

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?
      i. NH_4Cl is a relatively weaker acid compared to the HCl. If HCl was introduced into the bloodstream, it would decrease the pH to dangerously low levels.
   b. Why is NH_4Cl used instead of some other ammonium compound (such as NH_4NO_3 or NH_4I)?
      i. Cl^- ion obtained after the dissolution of NH_4Cl has a very small size and can more easily travel through the body, while other ammonia based acids lack this ability. 0.5
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.
   a. Acidic substances decompose to H^+ ion and X^- ion. For some weaker acidic substances, H^+ can be replaced with a positive polyatomic ion, like NH_4^+. In general, acidic substances increase the amount of positive ions, usually hydrogen, lowering the pH. The formula can be represented in the following form: $2H_2O(l) + H^+(aq) \rightleftharpoons 2H_3O^+$, where the hydrogen can be replaced with a positive ion for the weak acids.
3. The ability of hemoglobin (Hb) to carry oxygen throughout the body as oxyhemoglobin (HbO_2) is dependent on the pH of the blood. What effect would acidosis have on the ability of a patient to transport oxygen?
   a. Based on the given formula, $HbH^+(aq) + O_2(g) \rightleftharpoons HbO_2(aq) + H_2(g)$, introduction of extra H+ ions, as a result of acidosis, would shift the equilibrium to the left, decreasing the binding of hemoglobin to the oxygen. This will cause a lack of oxygen transported through the body.
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.
   a. As HCl will prevent most of the binding between the HbH^+ and O_2 and cause the equilibrium shift to the left, most of the oxygen will be released outside of the solution. Based on this statement, the change observed is mostly due to the decrease of the concentration of O_2 inside the solution.

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?
    a. Dissolved CO_2 in water creates a slightly acidic H_2CO_3. The chemical equilibrium formed can be outlined by the following reaction: $H_2O(l) + CO_2(g) \rightleftharpoons H_2CO_3(aq)$. Decrease of CO_2, that occupies a position of the reagent will lead to the equilibrium shift to the left, decreasing an amount of an acidic H_2CO_3. Therefore, the pH of soda will increase, as one of the acidic compounds was removed from the solution.
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?
    a. Pant, or deep breathing, causes the release of an additional volumes of CO_2 from the bloodstream. Based on the information obtained during the lab, it is possible to state that the pH level will decrease. Furthermore, according to the formula provided, decrease of CO_2 will lead to the equilibrium shift to the left, decreasing an available volume of $CaCO_3$. This will cause a weaker eggshells produced by the chickens.

Raw data:

Equilibrium shift.

1. ~~R.~~ $CuSO_4$ is a blue, transparent liquid.
2. ~~R.~~ Addition of concentrated NH_3 solution turned it dark blue ~~—~~ transparent.
3. ~~R.~~ Addition of HCl turned it pale blue transparent.
4. ~~R.~~ ~~R.~~ Same ~~thing~~ observations appeared, while the steps were repeated.
5. Addition of NH_3 into the previously made ~~solution~~ pale blue ~~so~~ transparent solution with a slight white precipitate.
6. ~~Rep~~ Third repetition did not yeild any precipitate.

Bonus Question: why?

Multiple equilibria

1. Transparent.
2. Yellowish pale transparent.
3. Transparent, produced CO_2 bubbles.
4. After HCl, white non transparent.
5. During the second repetition, white vapour was produced after the addition of HCl.
6. HNO_3

Part 2.

Na_2CO_3	clear
$AgNO_3$	yellow - brown. opaque
HNO_3	clear.
HCl	white opaque
NH_3	clear.
vap (HNO_3	vapor on wall slightly opaque white heated.
HCl	clearer but opaque
NH_3	clear heated. vapor.
KI	white - yellow opaque slightly
Na_2S	light dark opaque slightly.

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Lab III

~~Ph~~ pH of distilled water = 4.50 pH
Buffer solution
Observations:

1. Stirring rate is ~~6.10~~ 8.
2. The solution is transparent
3. pH is ~~8.~~ 8.70 pH.
4. After the addition of HCl, pH dropped to 7.79 pH.
5. After the addition of lactic acid, the pH dropped to 7.06 pH.
6. Bubbles are producing.
~~The stirring~~
7. The ~~stir~~ spinning speed was increased from 8 to 10. Bubbles started to detach stirring bar faster.
8. After the addition of sodium bicarbonate, pH raised to ~~7~~.74 pH. The solution appears to be transparent during all the steps.
9. After the addition of extra 0.50g of sodium bicarbonate pH raised to 7.26 pH.
10. Addition of dry ice resulted in a raise of pH to 7.30 pH. Dry ice ~~was~~ floating on top of the solution.
11. Dry ice melted and disappeared.
12. After the addition of additional CO_2, pH stabilized at 7.30
13. The addition of 0.40g NH_4Cl dropped pH to 7.20 ?.