Experiment Title:"VIOLENT FIRES SOON BURN OUT THEMSELVES SMALL SHOW'RS LAST LONG, BUT SUDDEN STORMS ARE SHORT"

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Procedure:

Starting the Equilibrium Shift Experiment

1. Add approximately 1 mL (~20 drops) of 0.1 mol/L $CuSO_4$ to a clean test tube(a) .

2. Dropwise, add concentrated NH_3 solution (NOTE: Ammonia has a VERY strong smell…do not breathe directly!!!) until you observe a change(b) .

3. Dropwise, add 1 mol/L HCl until you observe a change(c) .

4. What happens if you try to repeat Steps 2-3?

Multiple Equilibria

5. Place 0.5 mL of 0.1 mol/L Na_2CO_3 solution in a test tube(d) .

6. Add 0.5 mL of 0.01 mol/L $AgNO_3$ solution to the test tube(e) .

7. Dropwise, add 6 mol/L HNO_3 (NOTE: This acid is VERY concentrated and thus CORROSIVE!!) until you see a change(f) .

8. When the solution is clear, add 0.1 mol/L HCl dropwise until you see a change(g) .

9. Dropwise, add concentrated NH_3 until you see a change(h) .

10. What happens if you repeat Step 7 and then Step 9?

11. Add 0.1 mol/L KI dropwise until you see a change(i) .

12. Dropwise, add 0.1 mol/L Na_2S until you see a change(j) .

Buffer Solution and Blood pH

13. 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.

14. 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.(k)

15. Measure 2.50 g of sodium hydrogen carbonate, NaHCO3 and add it to the beaker. Allow the solution to stir until the salt dissolves completely(l)

16. Add 30 mL of 0.1 mol/L HCl (aq) to the beaker. (m)(n) Equilibria...23

17. 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(o)

18. The body tries to compensate (or raise the blood pH) when it feels the blood pH is low by increasing the rate of respiration, causing more CO2 (g) to be expired. Simulate this action by increasing the rate at which the solution is stirred (CAREFUL: if the stir bar hits the pH probe, the glass tip will break!!). (p)

19. In a critical clinical situation where the blood pH is low, intravenous sodium bicarbonate may be given after careful calculation of the amount required. Simulate this situation by adding approximately 0.50 g of sodium bicarbonate, NaHCO3 to the beaker. (q)

20. If too much sodium bicarbonate is added, the pH may become too elevated. Simulate this error by adding another 0.50 g of NaHCO3. (r)

21. When the body senses that the pH is elevated, it slows respiration to limit exhalation of CO2. Simulate this condition by adding a pellet of CO2. (s)

22. In extreme cases, if the pH is too elevated, a dilute weak acid may be injected directly into the blood. Simulate this condition by adding approximately 0.40 g of ammonium chloride, NH4Cl, to the beaker.(t) Cleaning Up!

23. Empty all the transfer pipets and rinse well with warm water (draw in water, shake well, and squeeze the water out several times).

24. Discard all solutions in the waste container, except for the silver solutions, which should be disposed in the container marked "silver wastes".

25. Clean your bench thoroughly and dispose of all paper towels in the appropriate waste container (solid wastes).

Discussion

Equilibrium Shift

After adding 20 drops of $CuSO_4$ solution into a clean test tube, the solution was expressed as a light blue colour . Next , two drops of NH_3 were put into the test tube where the ammonium turned the light blue $CuSO_4$ in the test tube into a royal blue. Then, 1 drop of HCl was added into the solution , this caused the color to change from royal blue to a lighter blue (visibly milkier and slightly darker than the initial 20 drops of $CuSO_4$ solution). In step 4, we added only 2 drops of NH3 to set off a reaction, after this drop the solution returned to a royal blue. Finally, 11 drops of HCl were added to the solution which had become a slightly more bright and light blue.

Multiple Equilibria

In step 5, Na_2CO_3 and $AgNO_2$(10 drops) were combined in a test tube yielding a dark yellow-brown colour. Next, in step 7, 1 drop of 6mol/L HNO3 was placed into this solution, Nex nitric acid was added to the new solution it then became clear that the acid has "eaten away" at the solution. In step 8, we added HCl , here the solution reacted to the chloric acid by turning into a milky-white color. Then in step 9,we added 50 drops of NH_3 into this milky-white solution and it then returned to a clear colour. Repeating steps 7-9 the following results occur: after 30 drops of HNO3, no visible results occur except for some precipitate. Then we added HCl, there was no change in color but we did observe little fragments of a precipitate . Then, more NH3 was added, no visible change . In step 11, 15 drops of KI were added to the solution in the test tube, the solution became a yellow-white color and the precipitation still had remained throughout this process .

Buffer Solution and blood pH

The PH of 100 ml of distilled water was 4.57.Immediatley after adding 2.50g of NaHCO3 that was dissolved into the water , the pH level rose and settled to 8.24 , which is a change of 3.67 in pH. Immediately after 30ml of 0.01 mol/L HCl was added to the beaker, the pH dropped to 8.00 which is a 0.24 pH level drop , bubbles were forming.Once the HCl had completely dissolved the PH read 7.40. Next 10ml of 0.85% lactic acid was added, the pH dropped to a pH level of 7.35. After this had completely mixed in , the PH was 7.30.The rate at which the solution was being stirred (the stir rate) had been increased from 6 to 7 and the pH slightly increased to 7.33 which was a 0.03 pH level increase) . Next 0.5g NaHCO3 was added and increased the pH levels to 7.40. When this had been completely mixed in the PH level was at 7.45 If pH level got too high, 0.5g of NaHCO3 would be added again,and they were. This caused an increase in the pH levels to 7.50 . Then a pellet of CO2 (0.85 g of dry ice) was added, and the pH levels increased to 7.55 , here there was steam produced since dry ice is an extremely cold/freezing temperature this caused the solution to now be room temperature. Finally 0.40g of NH4Cl was added to the beaker this decreased the pH level to 7.44

Observations:

Equilibrium Shift

What colour do you observe? What ion causes this colour?

The colour that this CuSO4 ion causes is observed to be a pale blue.The blue colour emitted from CuSO4 is caused by the transition metal changing orbitals.

What changes do you observe? From the chemical reactions provided, what chemistry explains the observations you made? What ion(s) cause(s) the colour?

The blue colour emitted from CuSO4 is caused by the transition metal changing orbitals. Since HCL is an acid and sulfates only react with bases , mixing these will create a highly acid solution of copper sulfate.

What changes do you observe? From the chemical reactions provided, what chemistry explains the observations you made? Have you seen this colour before? If so, how can you explain your observations using the given reactions?

This reaction causes a precipitate of Cu(OH)2 to begin forming. When Ammonia is added this precipitate begins to dissolve , resulting in the presence of $Cu(NH_3)_4{}^2$ which leads to a visible royal blue colour in the test tube.

In Step 4, are you able to see the same changes or are they different? Do you expect them to be the same or not? Why?

In step 4 (repeat steps 2 & 3) the HCl added dropwise lightens the solution, it reacts with the ammonia in the solution and creates ammonium ions , when this happens the equilibrium moves towards the reactants and the concentration of NH_3 decreases. After adding concentrated NH_3 again the solution returns to a royal blue due to the increase in concentration. Repeating these steps will change the direction the equilibrium will favour , when this happens there will be a visible shift in shades of blue due to the changing concentrations

Multiple Equilibria

What do you observe? What ions are causing the colour you observe?

Na_2CO_3 was observed to be clear in colour.

What changes do you observe? From the chemical reactions provided, what chemistry explains the observations you made? What ion(s) cause(s) this colour?

The solution became brown after adding $AgNO_3$ following reaction number 2 from the reaction table. This reaction involves a switching pair relationship between the two substances such that the reaction is written: Na_2CO_3 (aq) + $2AgNO_3$ (aq) -> Ag_2CO_3 (s) + $2NaNO_3$ (aq).

What changes do you observe? Do you think this change is reversible? Why? From the chemical reactions provided, what chemistry explains the observations you made?

6mol/L HNO_3 was added which changed the colour of the solution back to clear. This reaction is written as: Ag_2CO_3(s) + $2NaNO_3$(aq) + HNO_3(aq) → $2AgNO_3$(aq) + CO_2(g) + H_2O(aq) + $NaNO_3$ (aq). This reaction is reversible by the addition of NH_3 as it would react with Ag.

What changes do you observe? From the chemical reactions provided, what chemistry explains the observations you made? What ions are causing the colour you observe?

Unfortunately, adding NH_3 seemed to have no effect as no visible change occurred. Per reaction 5 of the reaction table: Ag(aq) + $2NH_3$ (aq) ⇌ $Ag(NH_3)_2$(aq). Perhaps there was not enough Ag for the NH_3 to react with.

Can you observe the same changes if you repeat Steps 7 and 9?

Upon adding a second round of 6mol/L HNO_3 a small amount of precipitate was successfully formed. The Ag comes together with CO to form Ag_2CO_3 which is the white coloured precipitate we observed. Similarly to the first round of NH_3, we observed no visible change after adding more to the solution.

What changes do you observe? From the chemical reactions provided, what chemistry explains the observations you made?

0.1mol/L KI was added to the solution but resulted in no visible change. Per reaction 7 from the reaction table: Ag(aq) + I(aq) ⇌ AgI (s). Perhaps there were not enough reactants in our solution in order to push the reaction forwards towards the product in order to create a precipitate.

What changes do you observe? From the chemical reactions provided, what chemistry explains the observations you made?

0.1mol/L Na2S made the solution yellow in colour per reaction 8: $2Ag^{+}$ (aq) + S^{2-} (aq) ⇌ Ag_2S (s).

Buffer Solution and blood pH

What is the pH? Is it what you expect the pH of water to be? If not, what does this tell you? Why might this pH be different than the one you expect?

The initial pH of the 100mL sample of distilled water was 4.57, this is due to the fact that the water has come in contact with other substances such as carbon dioxide in the air of the lab room, which make the water sample more acidic, hence having a lower pH value.

What pH do you observe? What reaction could cause this pH to be observed? What ions are responsible for the pH change?

After adding sodium hydrogen carbonate to the water sample the PH was 8.26, the reaction of this change is $NaHCO_3 + H_2O \rightarrow Na_2CO_3 + H_2O + CO_2$ and the ion responsible for the pH change is the sodium ion.

Why do we add the acid? What has been created by the addition of the acid to the solution of hydrogen carbonate ion?

We add HCl acid so that the chloride reacts with the sodium to produce NaCl. Adding the acid to the solution of hydrogen carbonate ion creates this and carbon dioxide as a byproduct (Forster, 2012).

What did you observe when the acid was added to the hydrogen carbonate ion? Why did you observe this? What is the pH? How does it compare to the pH of distilled water? How does it compare to the pH of blood? What species are present in the solution? Is there more of one species than another? Is this what you want based on the theory?

A change in pH was observed as the pH became 7.40 which is fairly close to the pH of distilled water and right in the range of blood's pH. There is one of each species in this reaction which is appropriate in order to achieve equilibrium.

What is the pH of the solution? What condition in the body does this situation simulate? What did you observe? Write a chemical reaction to describe the process that occurred and the observations you made.

Lactic acid lowered the pH to 7.30. This step simulates slow breathing. The chemical formula for lactic acid is written as $C_3H_6O_3$. We observe this because the reaction between the lactic acid and hydrogen carbonate ion alkalinizes the substance (Schmidt And Forsythe, 2000).

What do you observe? Do your observations match the information given to you? What happens to the pH of the solution with vigorous stirring? Is this what you expect? Justify your observations using a chemical reaction and equilibrium.

With vigorous stirring the pH increased to 7.33. Stirring allows for all contents within the solution to be moved around freely and thoroughly "mixed in" with the rest of the solution. This is expected as since CO_2 from the solution escapes the solution becomes less acidic.

What is the pH of the solution? What do you observe? Explain your observations using chemical reactions and equilibrium arguments.

The pH increased to 7.45 after adding $NaHCl_3$.

What is the pH of the solution? What do you observe? Explain your observations using chemical reactions and equilibrium arguments. What is the condition of the body at this point?

Adding 0.50 grams of $NaHCO_3$ pH was measured to be 7.55.

What is the pH of the solution? What do you observe? What happens to the dry ice? Explain your observations using chemical reactions and equilibrium arguments.

We observed vapour after adding a CO_2 pellet and recorded the pH to be 7.50.

What is the pH of the solution? What do you observe? Explain your observations using chemical reactions and equilibrium arguments.

NH4Cl was added and changed the pH to 7.44 giving us a final pH of 7.44.

Conclusion:

The concept of changing equilibria and Le Chatelier's Principle are thoroughly tested through the processes of this lab. As substances are added, various observations can be made about the visible colour of the solution and if precipitates are formed or not by the reaction. These occurrences are evidence that when a substance is added to the solution, a reaction occurs in an attempt for the solution to find equilibria either by turning reactants into products or breaking down products into their reactants until it achieves its balance or equilibria.

Reference(s):

1. "What in the World ISN'T Chemistry", General Chemistry Laboratory Manual, Dr. Rashmi Venkateswaran, 2019
2. "Investigations For The Environmentally Friendly Production Of Na2CO3 And HCl From Exhaust CO2, NaCl and H2O", Journal Of Cleaner Production, Volume 23, Issue 1, Martin Forster, 2012
3. "Sodium Bicarbonate For The Treatment Of Lactic Acids", PubMed, Schmidt And Forsythe, 2000

Answers to Questions 1-6:

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?*

The ammonium ion is used as the source of acid instead of HCl because the ammonium ions consume hydrogen ions, this produces ammonia (constituent of the buffer). Therefore, adding a strong acid will cause a resist in change of pH

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

NH4Cl is used instead of another ammonium compound (such as NH4NO3 or NH4I) since its milder than some of the other ammonium compounds, and it does the desired effect to create a reaction. NH4Cl is the mixture of a weak base and a strong acid, making it acidic. Weaker bases will will only slightly ionize, and a lower concentration of hydrogen ions.

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

Each time an acidic substance was added to the beaker, an observation that the solution had a change in color could be made. This can be explained by different solutions causing a shift in equilibrium which results in concentrations being altered. This may cause different colours to be expressed. Also hydrogen ions lowers the pH of the solution. As well depending on the acidic substance, there was some heat formed and created fog within the test tube. The equilibrium shifting based on the different acidic acids added are examples of physical change, in which, you can tell that a reaction has occurred.

3. *The ability of hemoglobin (Hb) to carry oxygen throughout the body as oxyhemoglobin (HbO2) is dependent on the pH of the blood. What effect would acidosis have on the ability of a patient to transport oxygen?*
HbH+ + O2 ⇌ HbO2 + H+

Hemoglobin (Hb) has the ability to carry oxygen throughout the body via the circulatory system as oxyhemoglobin (HbO2) and is dependent on the pH of the blood. According to Le Chatelier's Principle, there is an equilibrium shift to the left, since hydrogen ions attach to hemoglobin. This will decrease the affinity of the hemoglobin toward the oxygen. Now that there is less oxygen attached to the hemoglobin ,the transportation of oxygen is reduced. There is little to no oxygen reaching the cells because of the low concentration of oxygen in the blood. This can have many negative side effects such as pallor and fatigue.

$HbH^{+} + O2 \rightleftarrows HbO2 + H^{+}$

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.*

Based on our general knowledge about the color of blood and the information in question 3, we can provide an educated answer . The color of the blood will change from red to a dark brown since there is a new equilibrium . When the new equilibrium is adjusted and the reaction occurs backwards ,The oxyhemoglobin will split into oxygen and hemoglobin molecules, hemoglobin proteins are mainly made up of iron in which. The outcome makes the blood into a dark brown color.Blood will change to a darker red when deoxygenated.

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 fresh sample of soda had a pH of 2.92. The soda was placed on a magnetic stirrer to get it to become flat after ,the pH is measured again. The pH of the flat soda is expected to be higher than the pH of the fresh soda. The reason the flat sodas pH will be higher is that there is CO2 within the solution. As the solution is stirred, it comes into contact with the air surrounding it more often, therefore the concentration of carbon dioxide in this solution is higher than that of its surroundings, it will exit the solution which causes the pH to increase.

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

$$2\ H2O + CO2 \rightleftharpoons H2CO3 + H2O \rightleftharpoons H3O+ + HCO3\ - \rightleftharpoons CO3\ -2 + H3O+$$

Panting is fast breathing, which means that the organism releases CO2 from their bodies (the system) much faster than usual. This causes the organism's blood pH to become less acidic meaning an increase in the pH of their blood. The equilibria shifts towards the reactants to achieve equilibria since less product can be created with this new and lesser amount of CO2 within the chicken. This results in a weaker eggshell since less product can be created by the reaction than before , this is mainly due to the depletion of CO_2 which is a main component of the egg shell.

Appendix:

Raw Data:

Table 1. Reactions Being Studied

1	$[Cu(H_2O)_4]^{2+}$ (aq) + 4 NH_3 (aq) $\rightleftharpoons$ $[Cu(NH_3)_4]^{2+}$ (aq) + 4 H_2O (l)
2	2$AgNO_3$ (aq) + Na_2CO_3 (aq) $\rightleftharpoons$ Ag_2CO_3 (s) + 2 $NaNO_3$ (aq)
3	2 H^+ (aq) + CO_3^{2-} (aq) $\rightleftharpoons$ H_2CO_3 (aq) $\rightleftharpoons$ H_2O (l) + CO_2 (g)
4	

	$Ag^{+} (aq) + Cl^{-} (aq) \rightleftharpoons AgCl (s)$
5	$Ag^{+} (aq) + 2NH_3 (aq) \rightleftharpoons [Ag(NH_3)_2]^{+} (aq)$
6	$H^{+} (aq) + NH_3 (aq) \rightleftharpoons NH_4^{+} (aq)$
7	$Ag^{+} (aq) + I^{-} (aq) \rightleftharpoons AgI (s)$
8	$2Ag^{+} (aq) + S^{2-} (aq) \rightleftharpoons Ag_2S (s)$
9	$2 H_2O (l) + CO_2 (g) \rightleftharpoons H_2CO_3 (aq) + H_2O (l) \rightleftharpoons HCO_3^{-} (aq) + H_3O^{+} (aq)$

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Experiment 3 – October 11th 2019

Starting The Experiment
- $CuSO_4$ was light blue (1mL)
- Adding NH_3 darkened the mixture (2drops)
- Adding HCl lightened the mixture (1drop)
- After adding more NH_3, the solution went back to dark blue (2drops)
- After adding more HCl, the solution returned to light blue again (11drops)

Moving On
- 10drops of Na_2CO_3 in test tube was clear in colour
- 10 drops of $AgNO_3$ turned the solution brown
- 6mol/L HNO_3, 1 drop counted, cleared the solution
- NH_3, no visible change occurred (50drops)
- 6mol/L HNO_3, observed a small amount of percipitate formed (30drops)
- More NH_3 added, no visible change
- 0.1mol/L KI, no visible change
- 0.1 mol/L Na_2S made the solution yellow after 15 drops

P.T

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P.T

Buffer Solution And Blood pH
- pH 100mL distilled water: 4.57
- pH immediatley after adding 2.50g of sodium hydrogen: 8.24
- pH after sodium hydrogen settled (dissolved completely): 8.26
- pH immediatley after adding 30mL of 0.1 mol/L HCl: 8.00
- pH after HCl fully mixed into solution: 7.40
- pH immediatley after adding 10mL of 0.85% lactic acid: 7.35
- pH after lactic acid mixed into the solution: 7.30
- Turning up stirrer's rate from 6 to 7 raised the pH to 7.33
- pH immediatley after adding $NaHCl_3$: 7.40
- pH after $NaHCl_3$ mixed into the solution: 7.45
- pH immediatley after adding 0.50 g more of $NaHCO_3$: 7.50
- pH after $NaHCO_3$ settled: 7.55
- pH immediatley after adding CO_2 pellet: 7.53

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- Resting pH after adding CO_2 pellet: 7.50 (including a smokey effect) *vapour
- pH immediatley after adding NH_4Cl: 7.45
- pH after NH_4Cl settled into solution: 7.44
- Final resting pH: 7.44