

Equilibria Lab

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Introduction:

Equilibrium is defined as being a balance. In chemistry, equilibriums are a very important phenomenon in which are used to determine the outcome of a reaction. When considering reactions, arrows are often used to show in which direction it will tend to fall. Realistically all reactions typically don't proceed all the way to break down the initial reactants or even go in one static direction. Equilibrium in chemistry is a dynamic state in which defines a point in reaction where the concentrations of reactants and products are equal and will not change. It is the stable state of a reaction, which will not change unless acted on by outside forces, such as changing pressure or temperature. An equilibrium is a dynamic state meaning that the reaction doesn't just stop. At this given point, the reaction rates at which reactants are turned into products equals the rate at which products are turned into reactants, giving it the characteristic of being a dynamic state.

Each reaction is given a constant in which defines the equilibrium in which it lies. This constant, known as the equilibrium constant is described using the letter K and explains the ratio of concentrations between products and reactants in a reaction. For any given reaction of the form $aA + bB \rightleftharpoons cC + dD$, where the lower case letter describes molar ratios and upper case letters the chemicals being reacted, the K value is equal to $\frac{[C]^c[D]^d}{[A]^a[B]^b}$, where the square brackets are equal to the individual concentrations of the substances at equilibrium. The relative value of K can tell a lot about how much product and reactant are present at equilibrium. "If K is a large number, it means that the equilibrium concentration of the products is large. In this case, the reaction as written will proceed to the right (resulting in an increase in the concentration of products). If K is a small number, it means that the equilibrium concentration of the reactants is large. In this case, the reaction as written will proceed to the left (resulting in an increase in the concentration of reactants)" (Libretexts).

Given this K value, if a reactant or product was added into solution, the direction in which the added substance will break down can be found. The principle that describes this relationship between added substance and concentration shift is known as Le Chatelier's Principle. This principle states that "if a system in chemical equilibrium is subjected to a disturbance it tends to change in a way that opposes this disturbance" (Collins).

Although K does describe how a reaction will tend to fall and the concentrations of the individuals substances involved at equilibrium, it is also reliant on temperature. This means for the exact same reaction of given substances at different temperatures, the reaction will have a different K value. Changes in heat are always dependant on if a reaction is exothermic or endothermic. An increase in heat will favour the endothermic reaction, where-as a decrease in temperature will favour the exothermic reaction. The change in the values of K at different temperatures can be described using The Van't Hoff Equation; $\ln(k_1/k_2) = (\Delta_r H^\circ / R)(1/T_1 - 1/T_2)$, where $k_1 = \text{constant at } T_1$.

When looking at acids and bases, equilibrium is very important. Strong acids and bases are determined based on how well they dissociate in solution to H^+ / OH^- ions and their respective salt ion, meaning that the equilibrium in which they fall describes a lot about their strength. When an acid for example dissociates completely or very closely to in solution, it is known to be very strong, if it only slightly dissociates, weak.

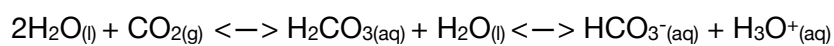
Using a similar method as finding the K value for equilibrium, the strength of acids and bases can be found. The K value of acids and bases describes the extent to which they dissociate in a solution and can be found in a very similar way to that of the K value. Given the general equation of acid dissociation in solution; $HA_{(g \text{ or } L)} + H_2O \rightleftharpoons A^-(aq) + H_3O^+$, the K value, which for an acid is known as the acid dissociation constant (K_a) is $= \frac{[H^+][A^-]}{[HA]}$. This value is very

similar to that of K , where it describes the extent to which an acid breaks down and the ratio of concentrations of each individual substance in the equation in which is being studied. Similarly, the base dissociation constant can be found the same way, just replacing the $[H^+]$ with $[OH^-]$.

The SI unit used to calculate the strength of a particular acid or base is pH and is a scale from 0-14 (but very strong acids or bases can often fall out of this scale), where 0 is the value for a strong acid, 7 a neutral solution and 14 a strong base. To find this value, the K_a value can be used, by substituting it into the equation $pH = -\log[H^+]$. Therefore, when the concentration of hydrogen ions in solution are high (meaning the acid dissociates very efficiently and K_a is large), the substance will have a very low pH, again showing that it is a strong base.

Within the human body, many of these concepts are at work. The ideal pH of the body lies between 7.35 and 7.45, which is a very tight margin. When the pH of the body falls below or above this, the equilibrium in which everything lies to function properly in the body is disrupted. This can cause significant damage to the body, or even death in more severe changes. When the pH falls below this normal zone for body function, it is considered to be in acidosis, which ranges from 7.35 to 6.8. When the body climbs above this normal zone, it can be considered to be in alkalosis, which ranges from 7.45 to 7.8. Anything below or above these two extremes can cause death. In order to keep the body in this very tight zones when dealing with multiple changes in environments and other factors such as sickness or diseases etc, the human body uses a system of buffers, which help maintain the proper pH in which the body functions.

The body naturally produces more acids in metabolic processes than it does bases, meaning the buffer system is more used to stabilize using bases than it does acids. An example of this is carbonic acid (H_2CO_3) and the bicarbonate ion (CO_3^{2-}) in the blood.



Where the concentration of HCO_3^- in the blood is much larger than the concentration of H_2CO_3 due to the fact that it is in equilibrium with the acids around it. The carbon dioxide in this reaction gets released through the respiratory system, causing the reaction to shift to the left to create more.

In the body, many chemical properties come into play when trying to maintain homeostasis. Buffer systems in the body maintain adequate pH in order for everything to work the way it needs in order for it to survive. The example above is one of many of the small reactions that lie in equilibrium in order to maintain pH and is joined by many others to maintain as close a pH to 7.40 as possible in order to keep everything running smoothly. In this lab, a few more buffer systems will be looked at in order to experimentally determine how they are able to maintain proper pH given multiple different situations within oneself.

Procedure:

See lab manual (Venkateswaran)

Data and observations/ results:

	Corresponding step in procedure	Observations
A)	1)	CuSO ₄ had a slightly blue colour when it was added to the test tube. This colour was because of the copper that is present in the solution, which is naturally blue
B)	2)	One drop (0.05mL) of NH ₃ caused liquid to instantaneously go dark blue, the ion that causes this is the Cu(NH ₃) ₄ ²⁺ ion which has a dark blue colour
C)	3)	4 drops of HCl caused the solution to turn brighter blue, very similar to the original blue colour of CuSO ₄ . This occurred because the HCl caused the Cu to bind to the SO ₄ molecules again, creating a quantity of CuSO ₄ similar to that first put in the test tube to be created. (Cu(NH ₃) ₄)SO ₄ + 4HCl = CuSO ₄ + 4NH ₄ Cl

Figure 1.0-

	Corresponding step in procedure	Observations
D)	5)	Na ₂ CO ₃ was a colourless transparent substance. Both ions Na ⁺ and CO ₃ ⁻ are colourless substances and therefore the entirety of the aqueous solution has no colour associated with it
E)	6)	The addition of AgNO ₃ caused a reaction to occur resulting in a milky brown liquid and the slight formation of a precipitate. This colour is associated with the Ag ₂ CO ₃ which is known to have a yellow or "greyish" colour to it, as well as the AgOH that would be found in the solution which naturally has a brown tint to it
F)	7)	One drop of HNO ₃ caused the solution to reverse back to the colourless solutions that were observed individually before they were mixed and able to react. The HNO ₃ caused the equilibrium of the reaction to shift towards the reactants, causing the more AgNO ₃ and Na ₂ CO ₃ to be produced, making the entirety of the substance clear again
G)	8)	One drop of HCl caused the substance to turn thick and milky. This caused AgCl to be formed along with other substances, AgCl being the main culprit in the milky like appearance as it is a white opaque substance
H)	9)	Two drops of NH ₃ caused the substance to turn back to its colourless transparent state. This is because it caused the reaction to revert back to its initial states, keeping the sodium carbonate and silver nitrate from reacting.

I)	11)	
J)	12)	

	Corresponding step in procedure	Observations
K)	14)	The pH of the water that was initially recorded was 4.82. This is significantly lower than the pH of water which should be 7. This occurred due to the fact that the pH meter sits in an acidic solution when it isn't being used, and acid still left on the reader made it seem as though the water was acidic
L)	15)	The pH of the solution was recorded at 8.12. The solution became basic when the NaHCO_3 was added because the H^+ found in the water reacted with the HCO_3^- in NaHCO_3 leaving a higher concentration of OH^- in the water than H^+ , giving the solution a basic pH
M)	16)	The acid is added to lower the pH of the water back down to its natural pH. When HCl is added, it completely dissociates in the water due to being a strong acid meaning it raises the abundance of H^+ in the solution. The addition of H^+ is similar to that which was lost due to the addition of NaHCO_3 and therefore brings the pH back down to around that of water (7.18 in this case). As a byproduct of this reaction, CO_2 was created which were observed as little gas bubbles being formed in the solution
N)	16)	The pH was 7.18 which is very close to that in which is what the blood should be. There will be more HCO_3^- in this solution than that of HCl as it dissociates completely where HCO_3^- doesn't. This is what causes the pH to be slightly more basic than 7.00. There will be slight amounts of HCl found, as well as CO_2 , Cl^- and H_2O as products
O)	17)	The addition of lactic acid caused the pH to drop to 7.06. This would be considered Acidosis in the body and is created from the muscles. This is a naturally occurring acid in the body and causes blood pH to rise due to its nature of breaking down the buffer HCO_3^- in the blood, causing the acidity of the blood to rise. $\text{C}_3\text{H}_6\text{O}_3(\text{aq}) + \text{HCO}_3^-(\text{aq}) \rightleftharpoons \text{H}_2\text{O} + \text{CO}_2$
P)	18)	When the rate at which the solution is stirred is sped up, it increases the rate at which the equation from the previous step is done, meaning more and more CO_2 is created and bubbled up out of the solution. This didn't change the pH very significantly

Q)	19)	0.50g of NaHCO ₃ caused the pH to rise up to 7.19. This occurred because the Na ⁺ and HCO ₃ ⁻ were split into their individual ions, where Na doesn't react and the HCO ₃ caused the overall pH of the solution to rise do to bind to the H ⁺ ions in the solution and leaving a higher concentration of OH ⁻ in the solution
R)	20)	0.5g of NaHCO ₃ were added once more to the solution which had the same impact on the solution as the observation from the last step. The pH raised to 7.26, which would (if it were in the body) cause the solution to be in the normal zone for body pH.
S)	21)	When the dry ice was added to the solution, most of it evaporated out of the beaker and onto the table or into the environment but some of it mixed in with the solution (could be seen as small gas bubbles inside of the beaker floating around). It didn't change the pH of the solution but should have affected the amount of H ₃ O ⁺ in the solution, causing it to deplete and cause alkalosis.
T)	22)	NH ₄ Cl caused the pH to drop by 0.02 (very slightly) to 7.24 $\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{NH}_3$ meaning that the ammonia in NH ₄ Cl binds with water to form H ₃ O ⁺ in the blood, which will lower the pH of the solution. Ammonia is a weak base which is why it doesn't lower the pH significantly.

Questions:

1.a) The ammonium ion is used as the acid source instead of HCl, as it is a weak acid meaning that its conjugate base is stronger and has a larger effect in the bloodstream. HCl is also a very strong and vigorous acid in which completely dissociates in the blood at once and can hence raise the pH very quickly. Due to ability of NH₄⁺ to dissociate much slower and not all at once (still have some NH₄⁺ at equilibrium) it will have a longer affect of managing blood pH over a period of time. If the H⁺ ions that it creates in its reaction become out of equilibrium, it will slightly shift over to balance out again, where as HCl is an all at once type of reaction.

b) Ammonium chloride is used in the body for multiple different reasons. One of these is that chloride is a very widely used ion in our bodies (to produce HCl for example) and is a much smaller atom than that of Iodine or NO₃. The size makes a significant difference due to the fact that these ions and elements have to travel through very small capillaries and spaces in the blood stream which are sometimes very small. Chlorine is a bit smaller which allows it to travel through the bloodstream much quicker, safer and efficiently.

2. Every time an acid was added to the beaker, the pH would go down because of their tendency to release H⁺ ions in a solution. When there is a higher concentration of H₃O⁺ in a solution than there is of OH⁻, then the pH will be lower (the exact value dependant on the exact ratio). Depending on the strength of the acid that was added, the pH would change accordingly. Some acids which are stronger will dissociate almost completely in a solution to their individual ions where as some won't dissociate as strongly meaning they won't give out as much H⁺ ions, causing less H₃O⁺ in a solution. An example of a strong acid would be HCl which breaks down in a solution to just its individual H⁺ and Cl⁻ ions: $\text{HCl} \rightleftharpoons \text{H}^+ + \text{Cl}^-$. As Cl⁻ is

negligible as a base, the H^+ reacts with the H_2O to create a much larger quantity of H_3O^+ , causing the pH to rise $H^+ + H_2O \rightleftharpoons H_3O^+$.

3. The Normal pH for the human body is 7.35, meaning that it is slightly basic. When the body falls into acidosis, which is stated as an acidic state below the pH of 7.35, there is much more H^+ present in the blood, meaning that the reaction between oxygen and hemoglobin: $HbH^+ + O_2 \rightleftharpoons HbO_2 + H^+$, falls to the right in trying to reach its equilibrium. This means that there will be a smaller amount of HbH^+ that binds to oxygen to become HbO_2 . When this occurs, oxygen can't be transported as efficiently as it should at normal body pH, and makes it difficult for the body to acquire the O_2 in which it needs to function correctly.

4. When HCl was added to the solution of the left, it completely dissociated into its individual H^+ and Cl^- ions which would cause the solution to become much more acidic. In the body (which is what the blood is used to being in to function properly) the pH should be in between 7.35 and 7.45, where as adding the HCl would cause it to go into acidosis. When acidosis occurs in the body, it causes the equilibrium of the equation from Q3 to revert backwards and towards the reactants. The abundance of H^+ causes the $HbHO_2$ (which is what gives the blood its red colour) to revert back to HbH^+ and O_2 , which causes the substance to lose its red colour and become more abundant in HbH^+ and individual O_2 substances (which can also be lost into the atmosphere, in the case of oxygen as it is naturally occurrent as a gas).

5. One of the reasons (along with many others) as to why soda is so acidic is due to how much carbon dioxide it contains. When the carbon dioxide in the soda reacts with the water it causes a proton (hydrogen ion) and HCO_3^- to be created. When there are high amounts of CO_2 reacting with water, it will cause more H^+ to be released, meaning the solution will become more acidic. If the pop is left time to go flat and for the CO_2 to escape from the bottle and the solution, there won't be as much H^+ being created. This means that the excess H_2O that isn't being used in that reaction will fall into its naturally occurring state of creating both OH^- and H^+ in equal proportions, causing more OH^- to be present in the drink and less H^+ , essentially causing the pH to go up slightly. CO_2 is not the only substance present in pop that causes it to be acidic, but is one of the reasons why, and for that reason if a lot of the CO_2 that is present in the solution right when the bottle is opened is let to escape, then the pH will slightly go up.

6. When chickens pant due to heat, they are hyperventilating, meaning that the CO_2 present in their bodies will be exhaled at a higher rate. This will cause the equation $2H_2O + CO_2 \rightleftharpoons H_2CO_3 + H_2O \rightleftharpoons H_3O^+ + HCO_3^- \rightleftharpoons CO_3^- + H_3O^+$ to shift towards the left and less CO_3^- and H_3O^+ will be created. This will cause the body to fall into alkalosis and also give less CO_3^- for the calcium to bind to, meaning the shells won't be as strong as if it were cooler out. If this occurred for long periods of time, the shells would become much weaker and the higher pH would be very hard on the body. In a cool climate or situation, the chickens would release much less CO_2 and cause the equilibrium to shift in the other direction, essentially causing acidosis and stronger egg shells due to the increase in CO_3^- in their bodies.

Conclusion:

In conclusion, the experiment proved that the buffer system of the blood does work efficiently. When small amounts of acids were added and the pH rised, the small given amounts of base that were added were able to correct the pH and bring it back to the normal values (around 7.40) without significantly changing the entire chemistry of the system.

- when 2 drops of ammonia were added, the substance cleared back up to a clear transparent liquid.

↳ if steps 7 and 9 are re-done, the same things will occur - it will reverse back to the equilibrium.

- when KI was added - there was a light precipitate formed (a lemon-yellow colour)

↳ when Na_2S was added, the solution turned a thick charcoal colour. (1 drop of each was added)

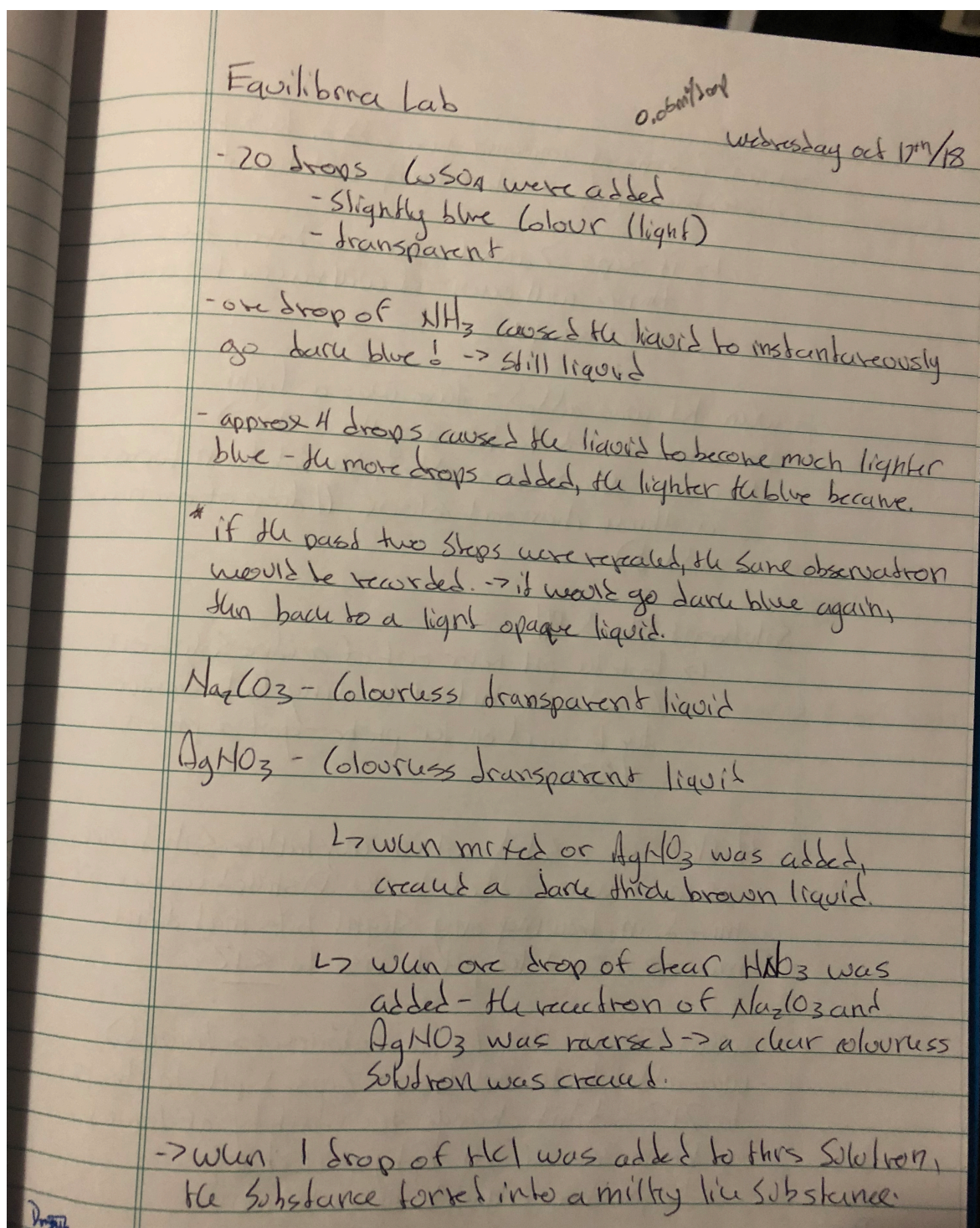
Solubility Blood pH

↳ took the pH meter out of acidic solution, caused our initial reading to be acidic due to acid on the probe - got a pH of 4.82

- NaHCO_3 is a solid which crystalline substance. when it was added to the distilled water, it mixed in, leaving only slight bits that didn't mix. → caused the pH to rise to 8.12

- when 30mL HCl (clear colourless liquid) was added to the solution - caused pH to drop to 7.18 - didn't change the identity of liquid on the meter at all (physical changes)
↳ produced gas bubbles to form.

Dan



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- lactic acid is a clear colourless substance
↳ when added to the solution, the pH dropped to 7.06.

- 0.50g of NaHCO_3 was added which caused pH to 7.19

- another 0.50g of NaHCO_3 was added pH raised to 7.26

- when dry ice was added, pH didn't change - the dry ice evaporated - then some mixing with the solution - didn't significantly change pH (CO_2) caused no change in pH.

- when NH_4Cl was added, the pH change (went down) by 0.02.

TYPE TO ENTER A CAPTION.

Citations:

Definition of 'Le Chatelier's principle'. (n.d.). Retrieved from <https://www.collinsdictionary.com/dictionary/english/le-chateliers-principle>

Libretexts. (2017, September 24). 15.4: The Magnitude of an Equilibrium Constant. Retrieved from [https://chem.libretexts.org/Textbook_Maps/General_Chemistry/Map:_General_Chemistry_\(Petrucci_et_al.\)/15:_Principles_of_Chemical_Equilibrium/15.4:_The_Magnitude_of_an_Equilibrium_Constant](https://chem.libretexts.org/Textbook_Maps/General_Chemistry/Map:_General_Chemistry_(Petrucci_et_al.)/15:_Principles_of_Chemical_Equilibrium/15.4:_The_Magnitude_of_an_Equilibrium_Constant)